



Inorganic Chemistry 1

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I. Inorganic Chemistry 1

by Chrispin Kowenje

II. Prerequisite Courses or Knowledge

Introduction to Chemistry 1: Matter and measurements, Structure and periodicity of the atom, Molecules and compounds, and Chemical reactions.

- Computer literacy (microsoft word processor and excel).

III. Time

120 hours

- Unit 1. Periodic Table of elements (14 hours)
- Unit 2. Atomic Structures and trends in s and p Block elements (30 hours)
- Unit 3. General Properties of s and p Block elements (40 hours)
- Unit 4. Occurrence, abundance and Extraction of the s and p block elements (16 hours)
- Unit 5. Economic uses of s and p Block Elements (20 hours)

IV. Material

In general, the learner needs to have;

- Course recommended text books
- Computer with internet connection
- CD-ROMs
- Access to e-Library
- Scientific calculator

V. Module Rationale

Introductory Chemistry 1 undepinned the concepts of matter, measurements, structure and periodicity of the atom, Molecules and compounds, and their chemical reactions. This study builds on the aforementioned concepts by being the foundation stone for better understanding of chemicals and the various ways of classifying the elements in Chemistry. With the ever expanding chemistry knowledge, here we show how best to predict and understand the simple applications of both chemical principles and chemical compounds and the use of the elements of the periodic table.



VI. Content

6.1 Overview

The course commences with a critical look at the Periodic table of elements and why they are classified as such. The table makes it easier to comprehend, early enough, the concepts and facts of chemistry of elements. The chemistry of the s- and p-block (or main block) elements and their compounds which then follows provides a systematic treatment of these elements and their compounds based on their common properties.

In unit 1, which utilises the concept of electronic configuration of the elements, you will review the classification of elements into metals, non-metals and metalloids as well as into s, p, d, and f blocks of the periodic table. This will be followed by a study of the atomic structures and general trends across a period and down a group in unit 2. Unit 3 examines general properties of s- and p- block elements. Unit 4 examines the abundance, occurrence and the extractions of the elements.

In recent years industrial research in the chemistry of main block elements has flourished; many new technological developments, such as Si-Ge transistors in computers and VCRs, have been achieved using these elements or their compounds. Examples of the economic uses are given in unit 5 and we hope that this module will stimulate your interest in inorganic chemistry, a field in which trained personnel are still scarce in both academic institutions and industrial laboratories.

6.2 Outline

120 hours

Unit 1. Periodic Table of elements (14 hours)

- Introduction and history to Periodicity in chemical elements
- Terminologies used in the periodic table of elements
- Classification of periodic table elements
- Electronic configuration of s-Block elements
- Electronic configuration of p-Block elements
- Electronic quantum numbers and electronic orbitals
- s-, p-, d-, and f-block elements
- Metals, non-metals and metalloids

**Unit 2. Atomic Structures and trends in s and p Block elements (30 hours)**

- Atomic size across and down the table
- Melting and boiling points of s and p block elements
- Electronegativities
- Ionization Energy trends
- Electron affinities
- Nuclear binding energies
- Effective nuclear charge

Unit 3. General Properties of s and p Block Elements (40 hours)

- Valency and formula of s-, and p-block elements and their compounds
- Naming of simple s-, and p-block compounds
- General characteristics of metals, non-metals and metalloids
- Physical Properties of S block elements
- Physical and chemical Properties of s-Block elements
- Physical and chemical Properties of p-Block elements

Unit 4. Occurrence, abundance and Extraction of s and p block elements (16 hours)

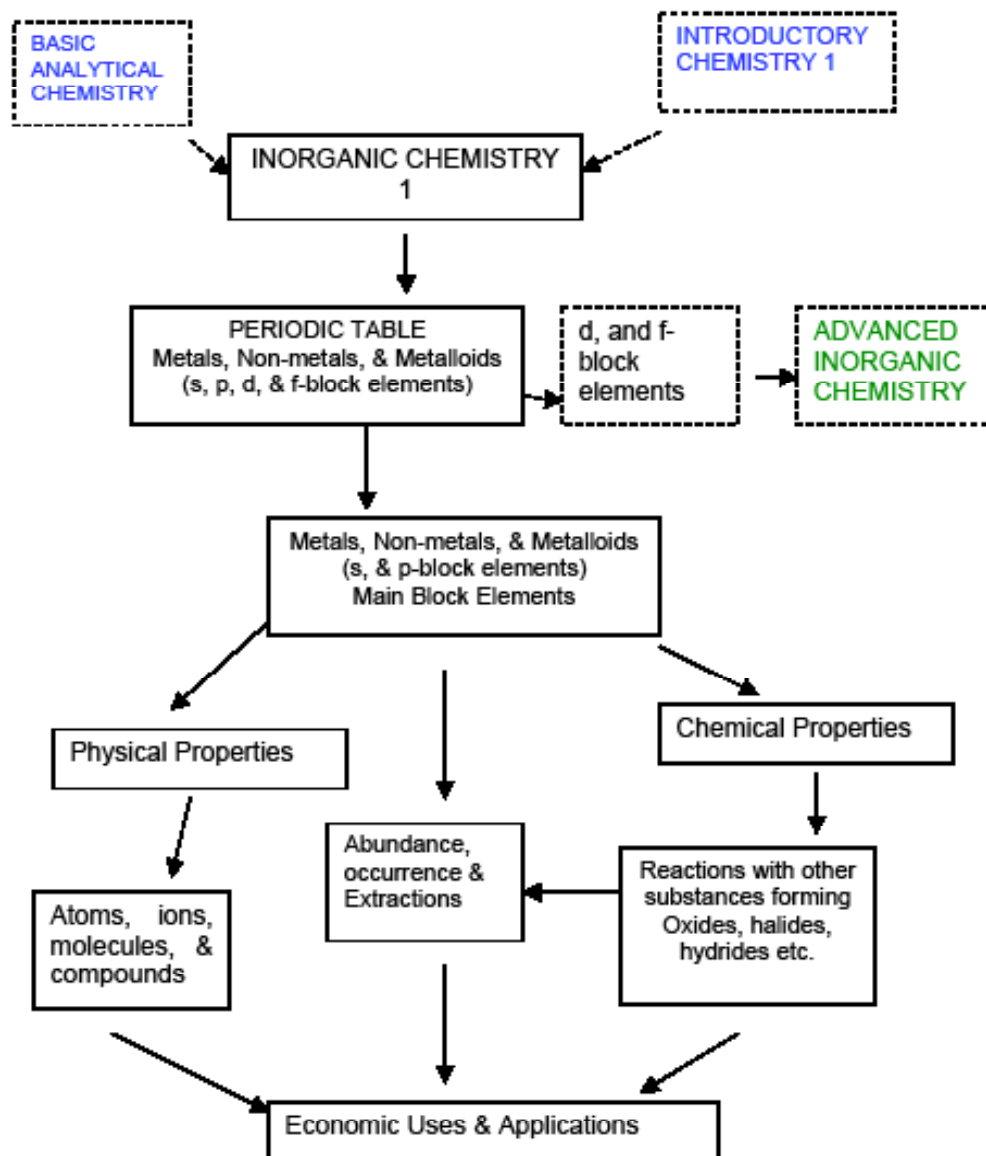
- Occurrence, abundance and extraction of group 1 elements
- Occurrence, abundance and extraction of group 2 elements
- Occurrence, abundance and extraction of group 13 - 16 elements
- Occurrence, abundance and extraction of group 17 elements
- Occurrence, abundance and extraction of group 18 elements

Unit 5. Economic uses of s and p Block Elements (20 hours)

- Economic use of selected group 1 elements
- Economic use of selected group 2 elements
- Economic uses of selected group 13 – 16 elements
- Economic uses of selected group 17 elements
- Economic uses of selected group 18 elements



The Flow of the Module: A Learner's Progression through the units.





VII. General Objective(s)

By the end of this module you are expected to be able to:

- Understand the placements of elements in the periodic table
- Classify elements into s, p, d and f blocks of the periodic table
- Describe and explain the periodicity in the physical properties of the s- and p-block elements
- Describe and explain the general features in the chemical properties of the s- and p-block elements
- Describe, explain and predict the physical and chemical properties of the compounds of s- and p-block elements
- Describe and even predict the methods used for extraction of the s- and p-block elements
- State and predict the economic uses of the s- and p-block elements

VIII. Specific Learning Objectives (Instructional Objectives)

Unit	Learning objective(s) At the end of this unit the learner should be able to:
1. Periodic Table of elements	<ul style="list-style-type: none"> - Describe and predic the position of an element in the periodic table by use of the atomic numbers. - Use different labelling sytems for the periodic table - Use the periodic table to classify the elements according to IUPAC system. - Identify and locate s-block elements in the periodic table according to their electron configuration - Identify and locate p-block elements in the periodic table according to their electron configuration - Classify the elements into s-, p-, d-, and f-block elements also as metals, non-metals and metalloids according to loca-tion in the periodic table



2. Atomic Structures of s and p Block elements

- Describe and explain the periodicity in atomic and ionic radii of elements of the periodic table.
 - Explain and describe how the trends in melting and boiling points are related to the electronic structures of the s and p block elements.
 - Describe, explain and predict the periodicity in electronegativity of the elements across the table.
 - Describe, explain and predict the periodicity in ionization energy of the elements across the table
 - Describe, explain and predict the periodicity in electron affinity of the elements across the table.
 - Describe, explain and predict the periodicity in nuclear binding energies of the elements across the table.
 - Calculate the nuclear binding energy of any element in the periodic table.
 - Calculate the effective nuclear charge of any element in the periodic table.
-

3. General Properties of s and p Block Elements

- Examine the valencies of the s-, and p-block elements.
 - Predict and state the common oxidation states for s-, and p-block elements.
 - Predict the formula and the name of simple compounds of s- and p-block elements.
 - Explain the stability of common oxidation state for s and p block elements.
 - Predict and explain the trend in the chemical and physical properties of s-block elements.
 - Predict and explain the trend in the chemical and physical properties of p-block elements.
 - Explain, give examples of the anomalous behaviour of the first (period 2) elements.
 - Describe what Inert pair effect is.
 - State and give example of diagonal relationships in the periodic table.
-



4. Occurrence, abundance and Extraction of Elements

- Appreciate the natural abundances of the s and p block elements.
 - Define terminologies associate with extraction of elements.
 - Observe the natural abundances table of the s and p-block elements.
 - Describe and explain the general methods used for extracting s block (metals) elements.
 - Describe and explain the general methods used for extracting groups 13, 14, 15, and 16 elements.
 - Describe and explain the general methods used for extracting groups 17 elements.
 - Describe and explain the general methods used for extracting group 18 elements.
 - Carry out a research project based on the extraction of one s-block element and two p-block elements and write a report.
-

5. Economic uses of s and p Block elements

- Appreciate the economic uses of some selected s block elements.
 - Appreciate the economic use of some selected p block elements.
 - Predict economic use(s) of some hypothetical element(s) of the s or p sections of the periodic table.
-



IX. Teaching and Learning Activities

9.1 Pre-assessment

Title of Pre-assessment : Pre-assessment test for Inorganic Chemistry 1.

Rationale : The test below evaluates your preparedness level for the Inorganic Chemistry 1 course. It is in no way intended to find your mistakes, but to bring out the bests you can recall, as per now.

Questions

1. An anion is,
 - a) an atom that has gained electrons
 - b) an atom that has lost electrons
 - c) a charged molecule
 - d) a charged element
2. Define an element
 - a) an atom that takes part in chemical reaction
 - b) a pure substance that consists of only one type of atom
 - c) different kinds of atoms of similar chemicals
 - d) an atom that can never be divided into simpler particles
3. What is an atomic number of an atom?
 - a) Number of protons
 - b) Number of electrons
 - c) Number of protons and neutrons
 - d) Number of electrons and protons
4. ${}_1\text{H}^1$, ${}_1\text{H}^2$, and ${}_1\text{H}^3$ are all
 - a) isotopes
 - b) allotrops
 - c) group members
 - d) anisotropes



5. chemical changes are those that
 - a) take place very fast
 - b) produce a new substance
 - c) where electrons are involved
 - d) where heat is applied

6. What is an atomic orbital
 - a) spot where electron is in an atom
 - b) a wave for an allowed energy state for an electron in atom
 - c) A wave for an allowed energy state for an electron in an atom or molecule
 - d) Sub-shell of an atom.

7. What are allotropes?
 - a) members of the same atom but different neutrons
 - b) Different forms of the same element
 - c) Different forms of the same atom
 - d) Different structures of the same element

8. what is octet rule?.
 - a) When electrons are 8 in an orbital
 - b) During bond formation, when atoms attain 8 electrons in valence shell electronic configuration.
 - c) when elements of group 8 (noble gases) are stable
 - d) having 8 valence electrons

9. Define a chemical bond
 - a) Shortest distance connecting two atoms in a molecule
 - b) attractive forces that hold two or more atoms together in a molecule
 - c) the two electrons that attract each other and hold atoms together
 - d) Shortest distance of approach between two atoms

10. Melting point of a substance is?
 - a) The temperature at which the solid and liquid phases are in equilibrium
 - b) the temperature where all solid has turned to liquid
 - c) the temperature when water is ice
 - d) is 0°C for Ice.



11. Electronic configuration

- a) the arrangement of electrons in atomic orbitals
- b) the presence of electrons in orbitals
- c) the shape of electronic orbitals such as s, p, and d-orbitals
- d) the act of studying electrons in an atom

12. Isoelectronic species are

- a) Atoms having the same amount of current in them
- b) Atoms or ions that have the same electronic arrangements
- c) Atoms that have the same number of electrons
- d) Atoms that have lost all their electrons

13. What is an electrochemical series?

- a) series for atoms of same group,
- b) The list of elements/compounds when arranged in order of their standard reduction potentials, from most oxidizing to most reducing.
- c) The list of atoms when arranged in order of their standard reduction potentials, from most oxidizing to most reducing.
- d) members of same period in periodic table.

14. The correct ranking of bonds in order of greatest to least bond strength is;

- a) covalent, Van der waals, hydrogen,
- b) Van der waals, hydrogen, covalent,
- c) covalent, hydrogen, Van der waals,
- d) hydrogen, Van der waals, covalent

15. Covalent bonds are formed by,

- a) metals and non-metals,
- b) sharing of electrons,
- c) transfer of electrons,
- d) non-metals alone

16. A chemical compound is always,

- a) chemically neutral,
- b) electrically neutral,
- c) both positively and negatively charged,
- d) either positively or negatively charged.



17. Define a catalyst

- a) a substance that increases the rate of a reaction
- b) a substance that increases the rate of a reaction but is itself not consumed
- c) a substance that produces the desired product faster in a chemical reaction
- d) a substance that changes a rate of a reaction.

18. Excited state of an atom is when?,

- a) one or more electrons are not in the lowest possible energy level
- b) when an electron leaves the atom
- c) when an atom acquires more energy
- d) when an atom has more electrons than it needs

19. Define an atomic mass unit (AMU) according to the IUPAC.

- a) 1 amu = 1/12 the mass of Carbon 12
- b) 1 amu = 1/16 the mass of Oxygen 16
- c) 1 amu = 1/1 the mass of hydrogen
- d) 1 amu = the maximum number of protons in an atom.

20. Anode

- a) tip of a bulb element
- b) electrode of an electrochemical cell where reduction occurs
- c) terminal of a dry cell
- d) electrode of an electrochemical cell where oxidation occurs



Answer Key

- 1.A
- 2.B
- 3.A
- 4.A
- 5.B
- 6.C
- 7.B
- 8.B
- 9.B
- 10.A
- 11.A
- 12.B
- 13.B
- 14.C
- 15.B
- 16.B
- 17.B
- 18.A
- 19.A
- 20. D



Pedagogical Comment For Learners

As a learner, the pre-assessment test evaluates your present level of chemistry knowledge as a link to that knowledge you are to acquire in this Inorganic chemistry 1. Your test score should help in identifying your competence and indicate areas where you need special emphasis on. The basics of understanding Inorganic chemistry 1 lies in appreciating the effects of electronic configurations and their concomitant interactions in atoms, ions, molecules, compounds etc as they direct the periodicity of the elemental properties. A learner who scores 40 percent or less in the pre-assessment test is likely to encounter difficulties comprehending the contents of this module and is, therefore, advised to review Introductory Chemistry 1, which is a prerequisite to this course. However, your performance index is not in anyway intended to make you be discouraged or be complacent; it is for you to appreciate how much effort you need to put in this work, be ready to make that extra mile.

Key Concepts

Atomic number. Is the number of protons in the nucleus or the number of electrons in an atom.

An ion. A charged atom or molecule. An **ion** is positive (cation) if it has lost electrons or negative (anion) if it has gained electrons

Isotopes. One member of a (chemical-element) family of atomic species which has two or more nuclides with the same number of protons (Z) but a different number of neutrons (N). Because the atomic mass is determined by the sum of the number of protons and neutrons contained in the nucleus, isotopes differ in mass. Since they contain the same number of protons (and hence electrons), isotopes have the same chemical properties.

Allotropes. One or more forms of an elementary substance. Examples are Graphite and diamond are both allotropes of carbon. O_2 and ozone, O_3 , are allotropes of oxygen.

Electronic ground state. This is electronic configuration of an atom with the lowest energy orbitals all occupied according to Hund's rule.

Isoelectronic series. A series for atoms or ions that have the same electronic arrangements/configuration.

Electroneutrality. The principle expresses the fact that all pure substances carry a net charge of zero. That is the overall charge in a molecule like $[Na^+Cl^-]^0$ is zero.

Chemical change. A change that results in the formation of a new substance, such as the burning of wood.

Catalyst. Anything/substance which creates a situation in which change can occur at a faster rate.



Atomic mass unit. An atomic mass unit (symbolized AMU or amu) is defined as precisely $1/12$ the mass of an atom of carbon-12. The carbon-12 (C-12) atom has six protons and six neutrons.

Chemical bond. An attractive force that holds atoms together to form molecules or Electrical interaction between electrons of one atom and the positive nucleus of another atom that result in the binding of atoms together in a stable unit.

Alloy. Is a homogeneous mixture of two or more elements, at least one of which is a metal, and where the resulting material has metallic properties. The resulting substance usually has different properties (sometimes substantially different) from those of its components.

Base. A substance that ionises in water to form hydroxide ions and a cation (there are more fundamental definitions of the term).

Bond polarity. The extent to which the bonding electron pairs between the two atoms is displaced towards one of the atoms.



X. Learning Activities

Learning activity # 1

Title of Learning Activity : **PERIODIC TABLE OF ELEMENTS**

At the end of this Unit, the learner should be able to;

1. Describe and predict the position of an element in the periodic table by use of the atomic numbers.
2. Classify the elements into s-, p-, d-, and f-block elements also as metals, non-metals and metalloids according to location in the periodic table.
4. Use the periodic table to classify the elements according to IUPAC system.
5. Use different labelling systems for the periodic table.

Summary of the learning activity

Being the first topic covered in this module, activity 1 includes the historical development of the Periodic table of elements by arranging the elements in horizontal rows according to their atomic weights. Identification of columns (groups) and rows (periods) of the periodic table and mark the metallic, non-metallic, and metalloids regions of the table and also as s-, p-, d-, and f-block elements portions are subsequently discussed. In addition, the elements will then be classified according to IUPAC system and lastly, different numbering for the modern periodic table will be discussed. At the end of each topic, relevant worked examples and excersises will follow to to aid you in development of conceptual and quantitative problem solving skills.

List of Required Readings

Text books

1. Alan G. Sharpe; *Inorganic Chemistry, 3rd Edition*. Longman Singapore Publisher, 1992.
2. Catherine E. Housecroft and Alan G. Sharpe; *Inorganic Chemistry*. Prentice-Hall International, USA. 2000.
3. J. D. Lee, *Concise Inorganic Chemistry, 4th edition*. Chapman & Hall, New York. USA. 1993.
4. Thomas R. Gilbert, Rein V. Kirss, and Geoffrey Davies; *Chemistry, The science in context*. W.W. Norton and company NY, USA. 2004.

**List of relevant useful links:**

1. <http://www.dayah.com/periodic/>
2. http://en.wikipedia.org/wiki/Periodic_table
3. <http://www.chem1.com/acad/webtext/atoms/atpt-6.html#ORG>
4. <http://chemistry.about.com/od/elementgroups/a/metals.htm>
5. <http://chem.lapeer.org/Chem1Docs/Mendeleev.html>
6. http://en.wikipedia.org/wiki/Electron_configuration

List of relevant MULTIMEDIA resources:

- Computer with internet connecting facility to access relevant links and free source resources.
- Multi-media resources such as CD players, VCD etc.
- CD-ROM for this module for compulsory reading and demonstrations.

Learning activities**Introduction and historical aspects of Periodic table:**

The periodic table of the chemical elements is a tabular method of displaying the chemical elements. Although precursors to this table exist, its invention is generally credited to a Russian chemist Dmitri Mendeleev in 1869. Mendeleev invented the table to illustrate recurring («periodic») trends in the properties of the elements. The layout of the table has been refined and extended over time, as new elements have been discovered, and new theoretical models have been developed to explain chemical behavior. The periodic table is now ubiquitous within the academic discipline of chemistry, providing an extremely useful framework to classify, systematize and compare all the many different forms of chemical behavior. The current, as of October 2006, standard table contains 117 elements (while element 118 has been synthesized, element 117 has not). Ninety-two elements are found naturally on Earth, and the rest are synthetic elements that have been produced artificially in particle accelerators.



Structure of the periodic table

Some definitions:

1. A **group** is a vertical column in the periodic table of the elements.
2. A **period** is a horizontal row in the periodic table of the elements.

See the Figure 1.1 below. Also available at: <http://en.wikipedia.org/wiki/Image:800px-PTable.png>.

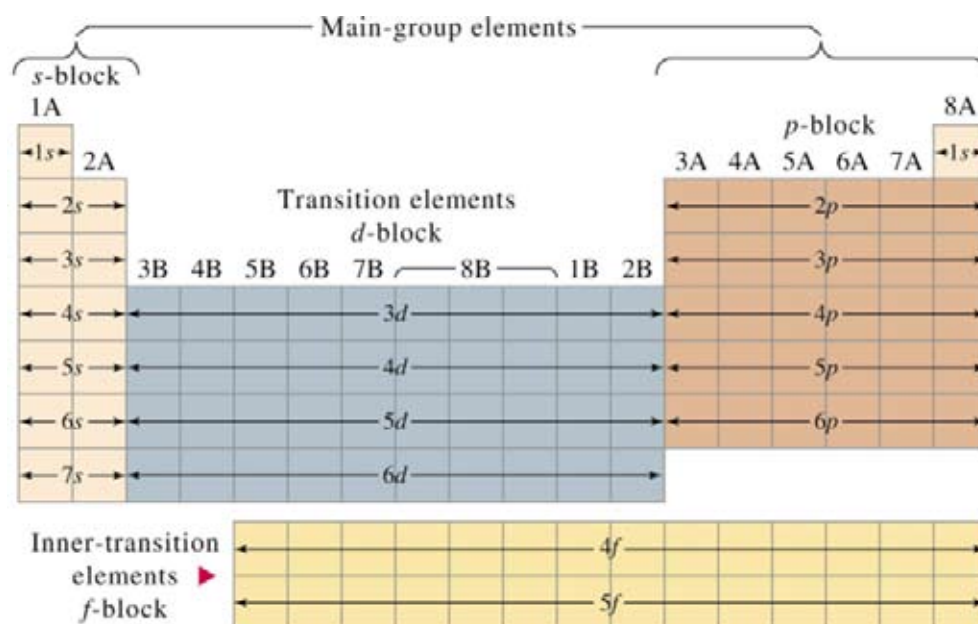


Figure 1.1: Periodic table of Elements showing the outermost shells and s, p, d, and f-block regions.

There are more than one way of designation for the groups in the periodic table. **Table 1.1** below compares labelling for the rest of available Periodic tables to that of **Figure 1.1**.



Table 1.1; The different ways of labellings used in modern periodic table.

Arabic + Alphabets	1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
Arabic No.	1	2	1	2	3	4	5	6	7	8	9	10	3	4	5	6	7	8
Roman No	IA	IIA	IIIB	IVB	VB	VIB	VII B	VIII	VIII	VIII	IB	IB	IIIA	IVA	VA	VIA	VIIA	VIIIA
IUPAC	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18

Key. In Arabic numbers (second row), in Blue are for main block while in black for transition block elements.

N.B. Since the last category is the IUPAC recommended, it will be used for all subsequent discussion.

Recommended name for Groups of the periodic table

Group Number	Recommended name
1	Alkali metals
2	Alkali earth metals
15	Pnicogens
16	Chalcogens
17	Halogens
18	Noble gases / inert gases

Electronic Configuration

Definition. Electronic configuration refers to the way in which electrons are arranged in the atomic orbitals.

Currently it is more fashionable to look at a Periodic table as determined by the number and the arrangement of the electrons of the elements. The primary determinant of an element's chemical properties is its electron configuration, particularly the valence shell (outer most) electrons. In addition, the type of orbital in which the atom's outermost electrons reside determines the «block» to which it belongs. The number of valence shell electrons determines the family, or group, to which the element belongs.

Exercise 1: Arrange the following orbitals; s, p, d, and f in terms of their energy levels (closeness to the nucleus) starting with the lowest (one nearest to the nucleus).

- A) s, p, d, f; B) p, s, f, d; C) d, f, p, s; D) f, d, p, s.

Solution is (A).



Electronic configuration of any element is intimately related to the position of the element in the periodic table. The chemical properties of an atom are largely determined by the arrangement of the electrons in its outermost «valence» shell (although other factors, such as atomic radius, atomic mass, and increased accessibility of additional electronic states also contribute to the chemistry of the elements as atomic size increases) therefore elements in the same table group are chemically similar because they contain the same number of «valence» electrons.

Summary of the quantum numbers

The state of an electron in an atom is given by four quantum numbers Three of these are integers and are properties of the atomic orbital in which it sits.

number	denoted	allowed values	represents
principal quantum number	n	integer, 1 or more	Partly the overall energy of the orbital, and by extension its general distance from the nucleus. In short, the energy level it is in. (1+)
azimuthal quantum number	l	integer, 0 to $n-1$	The orbital's angular momentum, also seen as the number of nodes in the density plot. Otherwise known as its orbital. (s=0, p=1...)
magnetic quantum number	m	integer, $-l$ to $+l$, including zero.	Determines energy shift of an atomic orbital due to external magnetic field (Zeeman effect). Indicates spatial orientation.
spin quantum number	m_s	$+\frac{1}{2}$ or $-\frac{1}{2}$ (sometimes called «up» and «down»)	Spin is an intrinsic property of the electron and independent of the other numbers. s and l in part determine the electron's magnetic dipole moment.

N.B. According to Pauli Exclusion Principle; No two electrons in one atom can have the same set of these four quantum numbers.



Shells and subshells

Shells and subshells (also called energy levels and sublevels) are defined by the quantum numbers, not by the distance of its electrons from the nucleus, or even their overall energy. In larger atoms, shells above the second shell overlap i.e. the restriction is no longer valid (see Aufbau principle).

States with the same value of n are related, and said to lie within the same *electron shell*. States with the same value of n and also l are said to lie within the same *electron subshell*, and those electrons having the same n and l are called equivalent electrons. If the states also share the same value of m , they are said to lie in the same atomic orbital. Because electrons have only two possible spin states, an atomic orbital cannot contain more than two electrons (Pauli exclusion principle).

A subshell can contain up to $4l + 2$ electrons; a shell can contain up to $2n^2$ electrons; where n equals the shell number.

Worked example

Here is the electron configuration for a filled fifth shell:

Shell	Subshell	Orbitals	No. of orbitals	Max No. of electrons
$n = 5$	$l = 0$	$m = 0$	1 type s orbital	2
	$l = 1$	$m = -1, 0, +1$	3 type p orbitals	6
	$l = 2$	$m = -2, -1, 0, +1, +2$	5 type d orbitals	10
	$l = 3$	$m = -3, -2, -1, 0, +1, +2, +3$	7 type f orbitals	14
	$l = 4$	$m = -4, -3, -2, -1, 0, +1, +2, +3, +4$	9 type g orbitals	18

Aufbau principle

Aufbau is a German word meaning to fill-up. It states, electrons enter into states in order of the states' increasing energy; i.e., the first electron goes into the lowest-energy state, the second into the next lowest, and so on. The order in which the states are filled is as follows:

The order is; $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 5d \approx 4f < 6p < 7s < 6d \approx 5f$.

The orbital labels **s**, **p**, **d**, and **f** originate from a now-discredited system of categorizing spectral lines as *sharp*, *principal*, *diffuse*, and *fundamental*, based on their observed fine structure. When the first four types of orbitals were described, they were associated with these spectral line types, but there were no other names. The designation **g** was derived by following alphabetical order. Shells with more than five subshells are theoretically permissible, but this covers all discovered elements. Some call the s and p orbitals *spherical* and *peripheral*.



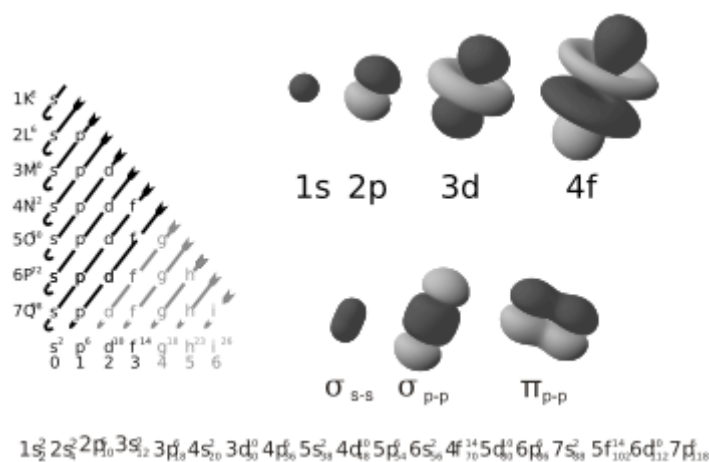
N.B. The energies of different orbitals are close together for higher values of n (1, 2, 3,..) and their relative energies change significantly upon ionization.

Hunds rule

In general the hund's rule of maximum multiplicity states that every orbital in a subshell is singly occupied with one electron before any one orbital is doubly occupied, and all electrons in singly occupied orbitals have the same spin.

Pauli exclusion principle

Simply stated: No two electrons in the same atom can be in the same quantum state. This means that no two electrons can have the same set of quantum states of: 1) energy, 2) angular momentum magnitude, 3) angular momentum orientation, and 4) orientation of intrinsic spin.



The order of increasing energy of the subshells can be constructed by going through downward-leftward diagonals of the table above (also see the diagram at the top of the page), going from the topmost diagonals to the bottom. The first (topmost) diagonal goes through 1s; the second diagonal goes through 2s; the third goes through 2p and 3s; and so on. This explains the ordering of the blocks in the periodic table.



Notation and the Filling of orbitals

You can think of an atom as a very bizarre house (like an inverted pyramid!) - with the nucleus living on the ground floor, and then various rooms (orbitals) on the higher floors occupied by the electrons. On the first floor there is only 1 room (the 1s orbital); on the second floor there are 4 rooms (the 2s, 2p_x, 2p_y and 2p_z orbitals); on the third floor there are 9 rooms (one 3s orbital, three 3p orbitals and five 3d orbitals); and so on. But the rooms aren't very big . . . *Each orbital can only hold 2 electrons.*

In the notation, a subshell is written in the form nx^y , where n is the shell number (i.e. 1, 2, 3..), x is the subshell label (i.e. s, p, d, f..) and y is the number of electrons in the subshell. An atom's subshells are written in order of increasing energy – in other words, the sequence in which they are filled. For instance, ground-state Lithium has two electrons in the 1s subshell and one in the (higher-energy) 2s subshell, so its ground-state configuration is written $1s^2 2s^1$. Phosphorus (atomic number 15), is as follows: $1s^2 2s^2 2p^6 3s^2 3p^3$.

For atoms with many electrons, this notation can become lengthy and so the noble gas notation is used. It is often abbreviated by noting that the first few subshells are identical to those of one or another noble gas. Phosphorus, for instance, differs from neon ($1s^2 2s^2 2p^6$) only by the presence of a third shell. Thus, the electron configuration of neon is pulled out, and phosphorus is written as follows: $[\text{Ne}]3s^2 3p^3$.

An even simpler version is simply to quote the number of electrons in each shell, e.g. (again for phosphorus): 2-8-5.

Valence and core electrons

The electronic configuration of the outermost (valence) electrons is significant. These electrons determine the chemical properties of the element. Electrons that occupy lower energy quantum levels are called core electrons. That of oxygen is $1s^2 2s^2 2p^4$. The core electrons of oxygen are those in 1s atomic orbital; the six electrons with $n = 2$ are the valence electrons.

Exercise 1

Write the electronic configuration for the following elements. In brackets are their Atomic numbers.

- Be (4), Na (11), Rb (37).
- B (5), N (7), P(15).
- Sc (21), Co (27).
- He (2), Ne(10), Ar(18).



Solution

- a) **Be, 4:** $1s^2 2s^2$; **Na, 11:** $1s^2 2s^2 2p^6 3s^1$; **Rb, 37:** $[\text{Kr}] 5s^1$
 b) **B, 5:** $1s^2 2s^2 2p^1$; **N, 7:** $1s^2 2s^2 2p^3$; **P, 15:** $1s^2 2s^2 2p^6 3s^2 3p^3 \equiv [\text{Ne}] 3s^2 3p^3$
 c) **Sc, 21:** $[\text{Ar}] 4s^2 3d^1$, **Co, 27:** $[\text{Ar}] 4s^2 3d^7$
 d) **He, 2:** $1s^2$, **Ne, 10:** $1s^2 2s^2 2p^6$, **Ar, 18:** $1s^2 2s^2 2p^6 3s^2 3p^6$

Relation to the structure of the periodic table

Electron configuration is intimately related to the structure of the periodic table. The chemical properties of an atom are largely determined by the arrangement of the electrons in its outermost «valence» shell (although other factors, such as atomic radius, atomic mass, and increased accessibility of additional electronic states also contribute to the chemistry of the elements as atomic size increases) therefore elements in the same table group are chemically similar because they contain the same number of «valence» electrons.

Those with ns^x are in group x , with np^x are in group $(12+x)$, with nd^x are in group $(2+x)$.

The total number of electron shells an atom has determines the period to which it belongs.

Element	Atomic No (Z)	Electronic configuration	Group	Period
Na	11	$1s^2 2s^2 2p^6 3s^1$;	1	3
Be	4	$1s^2 2s^2$;	2	2
Co	27	$[\text{Ar}] 4s^2 3d^7$	$(2+7) = 9$	4
Ar,	18:	$1s^2 2s^2 2p^6 3s^2 3p^6$	$(12+6) = 18$	3

Tally the findings in the table with the positions of the elements in **Table 1.1**.



Segmentation of the Periodic Table

As s-, p-, d-, and f-blocks of the Periodic table

Because of the importance of the outermost shell, the different regions of the periodic table are sometimes referred to as **periodic table blocks**, named according to the sub-shell in which the «last» electron resides, e.g. the *s*-block, the *p*-block, the *d*-block, etc.

Excise What commonality do you notice with the answers for exercise 1.1 with regards to the outershell electronic configurations of sections a, b, and c?

Answer: For; a) outermost orbitals have s^x configuration (s^1 or s^2), b) outermost orbitals have p^x configuration (p^1 or p^3), c) outermost orbitals have d^x configuration (d^1 or d^7), d) outermost orbitals have filled *s* and *p* (maximum allowed number of electrons) orbitals and are termed noble gases.

Thus elements in group 1 of the Periodic Table all have an outer electronic structure of ns^1 (where *n* is a number between 2 and 7). All group 2 elements have an outer electronic structure of ns^2 . Elements in (a) above are in groups 1 and 2 and are described as **s-block** elements.

The elements of groups 13 to 18 of the Periodic Table all have an outer electronic structure of np^x (where *n* is a number between 2 and 7 and *x* is a number between 1 and 6). All these elements in (b) above are in **p-block** region of the periodic table.

The in (c) above are *d*-block elements and are elements in which the last electron to be added to the atom is in a *d* orbital. They fall in groups 3 to 12 of the periodic table. The first series of these contains the elements from scandium to zinc they are also called transition elements or transition metals. The terms «transition element» and «**d-block** element» don't quite have the same meaning, but it doesn't matter in the present context.

While **f-block** elements are elements in which the last electron to be added to the atom is in a *f* orbital. They are normally presented as a pull-out from the main periodic table at the bottom (see figure 1.1)

Metals, Non-metals and metalloids

Metals fall into groups in the periodic table determined by similar arrangements of their orbital electrons and a consequent similarity in chemical properties. Groups of similar metals include the alkali metals and the alkaline-earth metals (*s*-block), transition metals (*d*-block) and the rare-earth metals (the lanthanide and actinide series of Group 3 see *f*-block). Most metals other than the alkali metals and the alkaline earth metals are called transition metals (see transition elements). Similarly metallic properties do overlap into the *p*-block elements bordering the *d*-block (later on diagonal relationship).



Non-metals are found in the upper right hand side of the periodic table in the p-block region. While metalloids are in the intersection between the metals and the non-metals and they roughly form a right hand leaning diagonal from top to bottom across the periodic table (see http://en.wikipedia.org/wiki/Periodic_table)

The six metalloids are B, Si, Ge, As, Sb, and Te. The properties of the metalloids have characteristics in between that of the metals and the nonmetals. They are good conductors of heat and electricity, but they are not perfect good conductors or insulators.

Problem set Unit 1 (accounts for 7 %).

1. How many periods are there in the periodic table?
a) 7, b) 18, c) 8, d) 10.
2. How many valence electrons does argon have?
a) 18, b) 8, c) 3, d) 6
3. What is the group number of the halogens?
a) 7 b) 17, c) VIIA, d) 18
4. How many elements are in period 4?
a) 6 b) 8 c) 18 d) 4
5. The elements in the modern/present periodic table are arranged according to their,
a) atomic numbers,
b) atomic masses,
c) mass numbers,
d) oxidation states.
6. The majority of the elements in the periodic table are
a) metals, b) non-metals, c) metalloids, d) Noble gases.
7. Elements in period 3 are alike in that they all have the same number of
a) Protons,
b) neutrons,
c) electrons in the valence shell
d) occupied principal energy levels.

Answer Key to Problem set Unit 1

1. a, 2. b, 3. b, 4. c, 5. a, 6.a, 7.d.



Learning activity # 2

Title of Learning Activity : **ATOMIC STRUCTURES AND TRENDS IN THE s AND p BLOCK ELEMENTS**

At the end of this Unit, the learner should be able to

1. State the general trend and predict the atomic sizes of an elements an a group/period.
2. State the general trend and predict the melting/boiling points of elements in a group/period..
3. State the general trend and predict the ionization energies of elements in a group/period.
4. State the general trend and predict the electronegativities of elements in a group/period.
5. State the general trend and predict the electron affinities of elements in a group/period.
6. State the general trend and predict the nuclear binding energies of elements in a group/period.
7. State the general trend and predict the effective nuclear charge of elements in a group/period.

Summary of the learning activity

In Unit 1 of this module we constructed, placed and classified elements accordingly in the periodic table. This unit futher explores common trends in atomic sizes, ionization enegies, electron affinities, electronegativities, and the nuclear binding energies of the elements across the table. The learner will use excel program to graphically follow and appreciate these trends. At the end of each topic, relevant worked examples and excersises will follow to aid in the development of conceptual and quantitative problem solving skills.

List of Required Readings

1. Alan G. Sharpe; *Inorganic Chemistry, 3rd Edition*. Longman Singapore Publisher, (1992).
2. Catherine E. Housecroft and Alan G. Sharpe; *Inorganic Chemistry*. Prentice-Hall International, USA. (2000).
3. J. D. Lee, *Concise Inorganic Chemistry, 4th edition*. Chapman & Hall, New York. USA. (1993).
4. Thomas R. Gilbert, Rein V. Kirss, and Geoffrey Davies; *Chemistry, The science in context*. W.W. Norton and company NY, USA. (2004).
5. William L. Jolly, *Modern inorganic Chemistry 2nd Ed*. McGraw-Hill. Inc. New York, USA (1991).



List of relevant useful links

<http://www.webelements.com/webelements/elements/text/periodic-table/econ.html>

For the properties of the elements and it is a live table. Once the table figure is active, identify the property of interest on the left and just click on any element in the Table and the details of the property sought for come along. It will display the information wanted.

<http://www.webelements.com/webelements/elements/text/Cu/ionz.html>

This gives the particular properties of the element. In this case it is Cu. You may substitute for Cu in the above URL and insert Cl for instance and get the details for Chlorine.

www.chemicool.com

Provides an interactive Periodic Table that includes information about elements and their states, energies, characteristics, abundance, and costs.

education.jlab.org/itselemental/index.html

Allows user to search facts and information about elements by atomic number, chemical name, or chemical symbol.

www.webelements.com

Includes information on elements and their properties, discovery, history, and uses.

<http://en.wikipedia.org/wiki/Gold>

Gives the different methods of calculating the electronegativity values.

<http://www.chemguide.co.uk/atoms/bonding/electroneg.html>

This page explains what electronegativity is, and how and why it varies around the Periodic Table. It looks at the way that electronegativity differences affect bond type and explains what is meant by polar bonds and polar molecules.

http://en.wikipedia.org/wiki/Electron_affinity

Shows electron affinity of various elements in a table form.

<http://hyperphysics.phy-astr.gsu.edu/hbase/chemical/ionize.html>, and http://en.wikipedia.org/wiki/Ionization_energy

For interactive tables on ionization trends and data.

http://en.wikipedia.org/wiki/Binding_energy

Details on nuclear binding energy

http://en.wikipedia.org/wiki/Effective_nuclear_charge

For detailed effective nuclear charge



List of relevant MULTIMEDIA resources

- Computer with internet connecting facility to access relevant links and free source resources.
- Multi-media resources such as CD players, VCD etc.
- CD-ROM for this module for compulsory reading and demonstrations.

Learning activities

TRENDS IN s- AND p-BLOCK ELEMENTS

The main value of the periodic table is the ability to present the predictable properties of an element based on its location on the table. It should be noted that the properties vary differently when moving vertically along the columns of the table, than when moving horizontally along the rows.

Consider the properties shown in **table 2.1** below

Element	Atomic No.	Group	Period	Melting points (°C)	Van de waals/atomic radius (pm)	Pauli electro-negativity (X ^p)	1 st ionization energy (KJmol ⁻¹)
Li	2	1	2	180	157	1.0	520.2
Na	11		3	98	191	0.9	495.8
K	19		4	63	235	0.8	418.9
Rb	37		5	39	250	0.8	403.0
Cs	55		6	28	272	0.8	375.7
Be	4	2	2	1283	112	1.6	899
Mg	12		3	650	160	1.3	737.7
Ca	20		4	842	197	1.0	589.8
Sr	38		5	777	215	0.9	549.5
Ba	56		6	727	224	0.9	502.1
B	5	13	2	2300	-	2.0	800
Al	13		3	660	143	1.6	577.5
Ga	31		4	30	153	1.8	578.8
In	49		5	156	167	1.8	558.2
Tl	81		6	304	171	1.6	589.4
O	8	16	2	- 218	140	3.4	1314
S	16		3	115	185	2.6	999.6



Se	34		4	221	200	2.6	941.6
Te	52		5	450	220	2.1	869.3
F	9	17	2	- 250	135	4.0	1681
Cl	17		3	- 101	180	3.2	1251.2
Br	35		4	- 7.2	195	3.0	1139.9
I	53		5	113	215	2.7	1008.4
He	2	18	1	- 275	99		2372.3
Ne	10		2	- 249	160		2080.7
Ar	18		3	- 189	191		1520
Kr	36		4	- 160	197		1350.8

Atomic sizes across and down the periodic table.

The distance between the nucleus and the outermost electrons of an atom is the atomic radius that in covalently bonded atom is the covalent radius. The distance between the nucleus and the outermost electrons of an ion is the ionic radius. Note; Ionic radius is the radius that an ion has in an ionic crystal, where the ions are packed together to a point where their outermost electronic orbitals are in contact with each other.

Note: Distances on the atomic scale have traditionally been expressed in Ångstrom units ($1\text{Å} = 10^{-8}\text{cm}$), but nowadays the picometer is preferred; $1\text{ pm} = 10^{-12}\text{ m} = 10^{-10}\text{ cm} = 100\text{ Å}$. The radii of atoms and ions are typically in the range 70-400 pm.

Exercise

- Use the values in Table 2.1 to plot an excel graph for the atomic number (horizontal/X-axis) versus the Van der waals/atomic radius (pm) (vertical/Y-axis) for any three (3) periods shown in the table.
- Use the values in Table 2.1 to plot an excel graph for the atomic number (horizontal/X-axis) versus the Van der waals/atomic radius (pm) (vertical/Y-axis) for the groups shown in the table.



Solution

The two plots in the exercise 2.1 above should show that the Elements in the same period show trends in atomic radius, from top to bottom in a group, the atomic radii of the elements increase. Since there are more filled energy levels, electrons are found farther from the nucleus. Moving left to right across a period, atomic radius usually decreases. This occurs because each successive element has an added proton and electron which causes the electron to be drawn closer to the nucleus.

Melting and boiling points across and down the periodic table

The melting point of an element or compound means the temperatures at which the solid form of the element or compound is at equilibrium with the liquid form. While the boiling point of an element or compound means the temperature at which the liquid form of an element or compound is at equilibrium with the gaseous form. We usually presume the air pressure to be 1 atmosphere.

Exercise

- c) Use the values in Table 2.1 to plot an excel graph for the atomic number (horizontal/X-axis) versus the melting points (vertical/Y-axis) for any three (3) periods shown in the table.
- d) Use the values in Table 2.1 to plot an excel graph for the atomic number (horizontal/X-axis) versus the melting points (vertical/Y-axis) for the groups shown in the table.

For example consider Period 3. The trends in melting points and boiling points going across Period 3 are not straightforward, and need more detailed consideration than the trends in Group 2: We note that, the melting points generally increase going from sodium to silicon, then decrease going to argon (with a “bump” at sulphur). Boiling points generally increase going from sodium to aluminium, then decrease to argon (again with a “bump” at sulphur).

Explanation

Sodium, magnesium and aluminium are all **metals**. They have **metallic bonding**, in which positive metal ions are attracted to delocalised electrons. Going from sodium to aluminium: a) the charge on the metal ions increases from +1 to +3 (with magnesium at +2), b) the number of delocalised electrons increases, c) so the strength of the metallic bonding increases, and d) the melting points and boiling points increase.

Silicon is a **metalloid** (an element with some of the properties of metals and some of the properties of non-metals). Silicon has a very high melting point and boiling point because: a) all the silicon atoms are held together by strong covalent bonds, and b) which need a very large amount of energy to be broken.



Phosphorus, sulphur, chlorine and argon, these are all **non-metals**, and they exist as small, separate molecules. Phosphorus, sulphur and chlorine exist as simple molecules, with strong covalent bonds between their atoms. Argon exists as separate atoms (it is monatomic). Their melting and boiling points are very low because: a) when these four substances melt or boil, it is the van der Waals' forces between the molecules which are broken, b) are very weak bonds, and c) so little energy is needed to overcome them.

Sulphur has a higher melting point and boiling point than the other three because sulphur does exist as S_8 molecules:

Electronegativity (X^p) across and down the periodic table

Electronegativity, symbol χ , first proposed by Linus Pauling in 1932, is a chemical property that describes the power of an atom (or, more rarely, a functional group) to attract electrons towards itself. Electronegativity, as it is usually calculated, is not strictly an atomic property, but rather a property of an atom *in a molecule*. Electronegativity of an element varies with its chemical environment, though it is considered to be a transferable property, meaning that similar values will be valid in a variety of situations.

Electronegativity cannot be directly measured and must be calculated from other atomic or molecular properties. Several methods of calculation have been proposed and, although there may be small differences in the numerical values of the electronegativity, all methods show the same periodic trends between elements.

Exercise

- Use the values in Table 2.1 to plot an excel graph for the atomic number (horizontal/X-axis) versus corresponding Pauli electronegativity (X^p) (vertical/Y-axis) for any three (3) periods shown in the table.
- Use the values in Table 2.1 to plot an excel graph for the atomic number (horizontal/X-axis) versus the corresponding Pauli electronegativity (X^p) (vertical/Y-axis) for the groups shown in the table.

Solution

For all the periods across the table, electronegativity increases from left to right. However, it decreases down any given group (See also Figure 2.1). You have to ignore inert gases group. It doesn't have an electronegativity, because it is not good at forming bonds.



Methods of calculation

Pauling electronegativity

Pauling first proposed the concept of electronegativity in 1932 as an explanation of the fact that the covalent bond between two different atoms (A–B) is stronger than would be expected by taking the average of the strengths of the A–A and B–B bonds. Consider a bond between two atoms, A and B. Each atom may be forming other bonds as well. If the atoms are equally electronegative, both have the same tendency to attract the bonding pair of electrons, and so it will be found *on average* half way between the two atoms. To get a bond like this, A and B would usually have to be the same atom. You will find this sort of bond in, for example, H₂ or Cl₂ molecules.

According to valence bond theory, of which Pauling was a notable proponent, the «additional stabilization» of the heteronuclear bond is due to the contribution of ionic canonical forms to the bonding.

The difference in electronegativity between atoms A and B is given by:

$$\chi_A - \chi_B = (\text{eV})^{-1/2} \sqrt{E_d(\text{AB}) - [E_d(\text{AA}) + E_d(\text{BB})]/2}$$

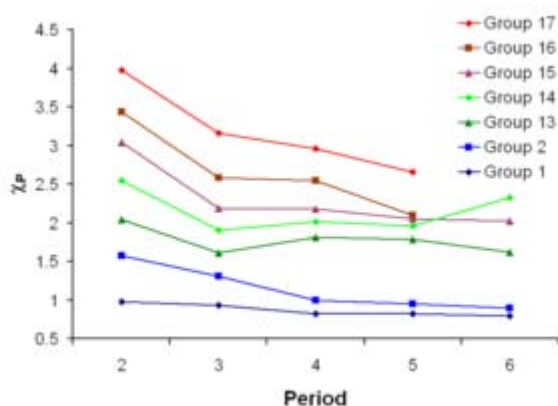
where the dissociation energies, E_d , of the A–B, A–A and B–B bonds are expressed in electronvolts, the factor $(\text{eV})^{-1/2}$ being included to ensure a dimensionless result. Hence, the difference in Pauling electronegativity between hydrogen and bromine is 0.73 (dissociation energies: H–Br, 3.79 eV; H–H, 4.52 eV; Br–Br 2.00 eV).

As only differences in electronegativity are defined, it is necessary to choose an arbitrary reference point in order to construct a scale. Hydrogen was chosen as the reference, as it forms covalent bonds with a large variety of elements: its electronegativity is fixed at 2.20 on a relative scale running from 0.7 to 4.0. It is also necessary to decide which of the two elements is the more electronegative (equivalent to choosing one of the two possible signs for the square root). This is done by «chemical intuition»: in the above example, hydrogen bromide dissolves in water to form H⁺ and Br[−] ions, so it may be assumed that bromine is more electronegative than hydrogen.

To calculate Pauling electronegativity for an element, it is necessary to have data on the dissociation energies of at least two types of covalent bond formed by that element. Allred updated Pauling's original values in 1961 to take account of the greater availability of thermodynamic data, and it is these «revised Pauling» values of the electronegativity which are most usually used.



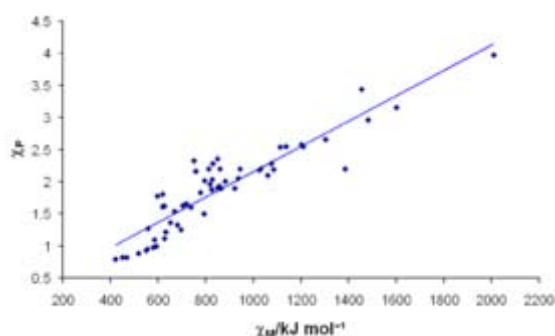
Figure 2.1. Periodic trends of electronegativity using the Pauling scale



The variation of Pauling electronegativity (y -axis) as one descends the main groups of the Periodic table from the second period to the sixth period.

Mulliken electronegativity

Mulliken proposed that the arithmetic mean of the first ionization energy and the electron affinity should be a measure of the tendency of an atom to attract electrons. As this definition is not dependent on an arbitrary relative scale, it has also been termed **absolute electronegativity**, with the units of kilojoules per mole or electronvolts.



The correlation between Mulliken electronegativities (x -axis, in kJ/mol) and Pauling electronegativities (y -axis).



However, it is more usual to make use of a linear transformation to transform these absolute values into values which resemble the more familiar Pauling values. For ionization energies and electron affinities in electronvolts,

$$\chi = 0.187(E_i + E_{ea}) + 0.17$$

and for energies in kilojoules per mole,

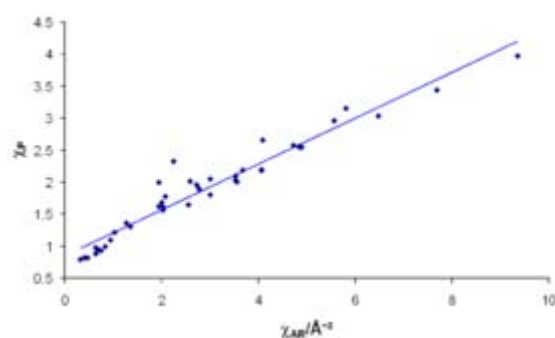
$$\chi = (1.97 \times 10^{-3})(E_i + E_{ea}) + 0.19$$

The Mulliken electronegativity can only be calculated for an element for which the electron affinity is known, fifty-seven elements as of 2006.

Allred-Rochow electronegativity

Allred and Rochow considered that electronegativity should be related to the charge experienced by an electron on the «surface» of an atom: the higher the charge per unit area of atomic surface, the greater the tendency of that atom to attract electrons. The effective nuclear charge, Z^* experienced by valence electrons can be estimated using Slater's rules, while the surface area of an atom in a molecule can be taken to be proportional to the square of the covalent radius, r_{cov} . When r_{cov} is expressed in ångströms,

$$\chi = 0.359 \frac{Z^*}{r_{cov}^2} + 0.744$$



The correlation between Allred-Rochow electronegativities (x -axis, in \AA^{-2}) and Pauling electronegativities (y -axis).

**The other methods are;**

1. Sanderson electronegativity: Sanderson has also noted the relationship between electronegativity and atomic size, and has proposed a method of calculation based on the reciprocal of the atomic volume. With a knowledge of bond lengths, Sanderson electronegativities allow the estimation of bond energies in a wide range of compounds. Also Sanderson electronegativities are used for different investigations in organic chemistry.
2. Allen electronegativity: Perhaps the simplest definition of electronegativity is that of Allen, who has proposed that it is related to the average energy of the valence electrons in a free atom.

$$\chi = \frac{n_s \varepsilon_s + n_p \varepsilon_p}{n_s + n_p}$$

where $\varepsilon_{s,p}$ are the one-electron energies of s- and p-electrons in the free atom and $n_{s,p}$ are the number of s- and p-electrons in the valence shell. It is usual to apply a scaling factor, 1.75×10^{-3} for energies expressed in kilojoules per mole or 0.169 for energies measured in electronvolts, to give values which are numerically similar to Pauling electronegativities.

In general, electronegativity increases on passing from left to right along a period, and decreases on descending a group. Hence, fluorine is undoubtedly the most electronegative of the elements while caesium is the least electronegative, at least of those elements for which substantial data are available.

There are some exceptions to this general rule. Gallium and germanium have higher electronegativities than aluminium and silicon respectively because of the d-block contraction. Elements of the fourth period immediately after the first row of the transition metals have unusually small atomic radii because the 3d-electrons are not effective at shielding the increased nuclear charge, and smaller atomic size correlates with higher electronegativity.



Correlation of electronegativity with oxidation number

In inorganic chemistry it is common to consider a single value of the electronegativity to be valid for most «normal» situations. While this approach has the advantage of simplicity, it is clear that the electronegativity of an element is *not* an invariable atomic property and, in particular, increases with the oxidation state of the element.

Consider:

1. What happens if two atoms of equal electronegativity bond together?

Consider a bond between two atoms, A and B. Each atom may be forming other bonds as well as the one shown - but these are irrelevant to the argument. If the atoms are equally electronegative, both have the same tendency to attract the bonding pair of electrons, and so it will be found *on average* half way between the two atoms.

Reminder: To get a bond like this, A and B would usually have to be the same atom. You will find this sort of bond in, for example, H₂ or Cl₂ molecules.

This sort of bond could be thought of as being a «pure» covalent bond - where the electrons are shared evenly between the two atoms. The molecule is said to be non-polar.

A polar bond is a covalent bond in which there is a separation of charge between one end and the other - in other words in which one end is slightly positive and the other slightly negative. Examples include most covalent bonds. The hydrogen-chlorine bond in HCl or the hydrogen-oxygen bonds in water are typical. In general, the greater the difference in electronegativity between two atoms, the more polar the bond that will be formed between them, with the atom having the higher electronegativity being at the negative end of the dipole

2. What happens if B is slightly more electronegative than A?

B will attract the electron pair rather more than A does.

That means that the B end of the bond has more than its fair share of electron density and so becomes slightly negative. At the same time, the A end (rather short of electrons) becomes slightly positive.

3. What happens if B is a lot more electronegative than A?

In this case, the electron pair is dragged right over to B's end of the bond. To all intents and purposes, A has lost control of its electron, and B has complete control over both electrons. Ions have been formed.

The implication of all this is that there is no clear-cut division between covalent and ionic bonds. In a pure covalent bond, the electrons are held on average exactly half way between the atoms. In a polar bond, the electrons have been dragged slightly towards one end.

How far does this dragging have to go before the bond counts as ionic? There is no real answer to that. You normally think of sodium chloride as being a typically



ionic solid, but even here the sodium hasn't *completely* lost control of its electron. Because of the properties of sodium chloride, however, we tend to count it as if it were purely ionic.

Lithium iodide, on the other hand, would be described as being «ionic with some covalent character». In this case, the pair of electrons hasn't moved entirely over to the iodine end of the bond. Lithium iodide, for example, dissolves in organic solvents like ethanol - not something which ionic substances normally do.

In Summary

- No electronegativity difference between two atoms leads to a pure non-polar covalent bond.
- A small electronegativity difference leads to a polar covalent bond.
- A large electronegativity difference leads to an ionic bond.
- $\Delta X \leq 1.2$ covalent bond
- $\Delta X \approx 1.5$ moderately ionic
- $\Delta X \geq 2.0$ ionic

Quiz. Is CHCl_3 polar or non-polar?.

Answer: CHCl_3 is polar. The hydrogen at the top of the molecule is less electronegative than carbon and so is slightly positive. This means that the molecule now has a slightly positive «top» and a slightly negative «bottom», and so is overall a polar molecule.

Explaining the patterns in electronegativity

The attraction that a bonding pair of electrons feels for a particular nucleus depends on: the number of protons in the nucleus, the distance from the nucleus, and the amount of screening by inner electrons.

Why does electronegativity increase across a period?

Consider sodium at the beginning of period 3 and chlorine at the end (ignoring the noble gas, argon). Think of sodium chloride as if it were covalently bonded.

Both sodium and chlorine have their bonding electrons in the 3-level. The electron pair is screened from both nuclei by the 1s, 2s and 2p electrons, but the chlorine nucleus has 6 more protons in it. It is no wonder the electron pair gets dragged so far towards the chlorine that ions are formed.

Electronegativity increases across a period because the number of charges on the nucleus increases. That attracts the bonding pair of electrons more strongly.

Why does electronegativity fall as you go down a group?

Think of hydrogen fluoride and hydrogen chloride.

The bonding pair is shielded from the fluorine's nucleus only by the $1s^2$ electrons. In the chlorine case it is shielded by all the $1s^2 2s^2 2p^6$ electrons. In each case there



is a net pull from the centre of the fluorine or chlorine of +7. But fluorine has the bonding pair in the 2-level rather than the 3-level as it is in chlorine. If it is closer to the nucleus, the attraction is greater.

As you go down a group, electronegativity decreases because the bonding pair of electrons is increasingly distant from the attraction of the nucleus.

NOTE; Positive ions can have the effect of polarising (electrically distorting) nearby negative ions. The polarising ability depends on the charge density in the positive ion.

Polarising ability increases as the positive ion gets smaller and the number of charges gets larger.

As a negative ion gets bigger, it becomes easier to polarise. For example, in an iodide ion, I^- , the outer electrons are in the 5-level - relatively distant from the nucleus. A positive ion would be more effective in attracting a pair of electrons from an iodide ion than the corresponding electrons in, say, a fluoride ion where they are much closer to the nucleus.

Aluminium iodide is covalent because the electron pair is easily dragged away from the iodide ion. On the other hand, aluminium fluoride is ionic because the aluminium ion can't polarise the small fluoride ion sufficiently to form a covalent bond.

Ionization Energy/POTENTIAL

The ionization energy potential or E_1 of an atom or molecule is the energy required to remove one mole of electrons from one mole of isolated gaseous atoms or ions. More generally, the n^{th} ionization energy is the energy required to strip of an n^{th} mole of electrons after the first $(n+1)$ mole of electrons have already been removed. It is considered as a measure of the «reluctance» of an atom or ion to surrender an electron, or the «strength» by which the electron is bound; the greater the ionization energy, the more difficult it is to remove an electron.

Usually, any subsequent ionization energy involves removing an electron from an orbital closer to the nucleus. Electrons in the closer orbital experience greater forces of electrostatic attraction, and thus, require more energy.

Generally, any subsequent ionization energy involves removing an electron from an orbital closer to the nucleus in an already more positively charged ion. Electrons in the closer orbital experience greater forces of electrostatic attraction, and thus, require more energy to be removed.

Exercise

Use the values in Table 2.1 to plot an excel graph for the atomic number (horizontal/X-axis) versus corresponding 1st ionization energies (vertical/Y-axis) for any three (3) periods shown in the table. Repeat the same for 1st ionization energies for the groups shown in the table.



Solution

It is a minimum for the alkali metals which have a single electron outside a closed shell. It generally increases across a row (period) with the periodic maximum for the noble gases which have closed shells. For example, sodium requires only 496 kJ/mol or 5.14 eV/atom to ionize it while neon, the noble gas immediately preceding it in the periodic table, requires 2081 kJ/mol or 21.56 eV/atom. The ionization energy is one of the primary energy considerations used in quantifying chemical bonds.

The ionization potential is an indicator of the reactivity of an element. Elements with a low ionization energy tend to be reducing agents and to form salts.

Table 2.4. Successive ionization energies of *in kJ/mol* ($96.485 \text{ kJ/mol} = 1 \text{ eV}$) some selected elements.

Element	First	Second	Third	Fourth	Fifth	Sixth	Seventh
Na	496	4,560					
Mg	738	1,450	7,730				
Al	577	1,816	2,881	11,600			
Si	786	1,577	3,228	4,354	16,100		
P	1,060	1,890	2,905	4,950	6,270	21,200	
S	999.6	2,260	3,375	4,565	6,950	8,490	27,107
Cl	1,256	2,295	3,850	5,160	6,560	9,360	11,000
Ar	1,520	2,665	3,945	5,770	7,230	8,780	12,000

Application

In order to determine how many electrons are in the outermost shell of an element, one can use the ionization energy. If, for example, it required 1,500 kJ/mol to remove one mole of electrons and required 6,000 kJ/mol to remove another mole of electrons and then 5,000 kJ/mol, etc. this means that the element had one electron in its outermost shell. This means that the element is a metal and in order for this element to achieve a stable complete outer shell, it needs to lose one electron. Thus, the first electron is easy to remove and consequently the ionization energy is low. Notice, however, that once the stable complete outer shell has been formed, it becomes *much* more difficult to remove the next electron. If that electron can be removed the consequent one can be removed a bit more easily.



Electrostatic explanation

Atomic ionization energy can be predicted by an analysis using electrostatic potential and the Bohr model of the atom, as follows.

Consider an electron of charge $-e$, and an ion with charge $+ne$, where n is the number of electrons missing from the ion. According to the Bohr model, if the electron were to approach and bind with the atom, it would come to rest at a certain radius a . The electrostatic potential V at distance a from the ionic nucleus, referenced to a point infinitely far away, is:

$$V = \frac{1}{4\pi\epsilon_0} \frac{ne}{a}$$

Since the electron is negatively charged, it is drawn to this positive potential. (The value of this potential is called the *ionization potential*). The energy required for it to «climb out» and leave the atom is:

$$E = eV = \frac{1}{4\pi\epsilon_0} \frac{ne^2}{a}$$

This analysis is incomplete, as it leaves the distance a as an unknown variable. It can be made more rigorous by assigning to each electron of every chemical element a characteristic distance, chosen so that this relation agrees with experimental data.

It is possible to expand this model considerably by taking a semi-classical approach, in which momentum is quantized. This approach works very well for the hydrogen atom, which only has one electron. The magnitude of the angular momentum for a circular orbit is:

$$L = |\mathbf{r} \times \mathbf{p}| = r m v = n \hbar$$

The total energy of the atom is the sum of the kinetic and potential energies, that is:

$$E = T + U = \frac{p^2}{2m_e} - \frac{ke^2}{r} = \frac{m_e v^2}{2} - \frac{ke^2}{r}$$

Velocity can be eliminated from the kinetic energy term by setting the Coulomb attraction equal to the centripetal force, giving:

$$T = \frac{ke^2}{2r}$$



Now the energy can be found in terms of k , e , and r . Using the new value for the kinetic energy in the total energy equation above, it is found that:

$$E = -\frac{ke^2}{2r}$$

Solving the angular momentum for v and substituting this into the expression for kinetic energy, we have:

$$\frac{n^2\hbar^2}{rm_e} = ke^2$$

This establishes the dependence of the radius on n . That is:

$$r(n) = \frac{n^2\hbar^2}{km_e e^2}$$

At its smallest value, n is equal to 1 and r is the Bohr radius a_0 . Now, the equation for the energy can be established in terms of the Bohr radius. Doing so gives the result:

$$E = -\frac{1}{n^2} \frac{ke^2}{2a_0} = -\frac{13.6eV}{n^2}$$

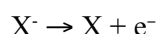
This can be expanded to larger nuclei by incorporating the atomic number into the equation.

$$E = -\frac{Z^2}{n^2} \frac{ke^2}{2a_0} = -\frac{13.6Z^2}{n^2} eV$$



Electron Affinity (EA).

The **electron affinity**, (EA), of an atom or molecule is the energy required to detach an electron from a singly charged negative ion, i.e., the energy change for the process



An equivalent definition is the energy released ($E_{\text{initial}} - E_{\text{final}}$) when an electron is attached to a neutral atom or molecule. It should be noted that the sign convention for E_{ca} is the opposite to most thermodynamic quantities: a positive electron affinity indicates that energy is *released* on going from atom to anion.

All elements whose EA have been measured using modern methods have a positive electron affinity, but older texts mistakenly report that some elements such as alkaline earth metals have negative E_{ca} , meaning they would repel electrons. The electron affinity of the noble gases have not been conclusively measured, so they may or may not have slightly negative EAs. Atoms whose anions are relatively more stable than neutral atoms have a greater E_{ca} . Chlorine most strongly attracts extra electrons; mercury most weakly attracts an extra electron. E_{ca} of noble gases are close to 0.

Although E_{ca} vary in a chaotic manner across the table, some patterns emerge. Generally, nonmetals have more positive E_{ca} than metals.

E_{ca} is not limited to the elements but also applies to molecules. For instance the electron affinity for benzene is negative, as is that of naphthalene.

N.B. Don't confuse Electron affinity with work function. The **work function** is the minimum energy (usually measured in electron volts) needed to remove an electron from a solid to a point immediately outside the solid surface (or energy needed to move an electron from the Fermi energy level into vacuum). Here «immediately» means that the final electron position is far from the surface on the atomic scale but still close to the solid on the macroscopic scale. The work function is an important property of metals. The magnitude of the work function is usually about a half of the ionization energy of a free atom of the metal.

Nuclear binding energies

Nuclear binding energy is derived from the strong nuclear force and is the energy required to disassemble a nucleus into free unbound neutrons and protons, strictly so that the relative distances of the particles from each other are infinite (essentially far enough so that the strong nuclear force can no longer cause the particles to interact). At the atomic level, the binding energy of the atom is derived from electromagnetic interaction and is the energy required to disassemble an atom into free electrons and a nucleus.



The binding energy for a single nucleus is given by

$$\Delta E = \Delta m \cdot c^2 = (m_s - m_b) \cdot c^2 = (Z \cdot m_p + N \cdot m_n - m_b) \cdot c^2$$

where

c is the speed of light;

m_s is the mass of the separated nucleons;

m_b is the mass of the bound nucleus;

Z is the atomic number of the bound nucleus;

m_p is the mass of one proton;

N is the number of neutrons;

m_n is the mass of one neutron.

In atoms with high or very low ratios of protons to neutrons, the binding energy becomes negative, resulting in an unstable nucleus.

Specific example: a deuteron

A deuteron is the nucleus of a deuterium atom, and consists of one proton and one neutron. The experimentally-measured masses of the constituents as free particles are

$$m_{\text{proton}} = 1.007825 \text{ u};$$

$$m_{\text{neutron}} = 1.008665 \text{ u};$$

$$m_{\text{proton}} + m_{\text{neutron}} = 1.007825 + 1.008665 = 2.01649 \text{ u}.$$

The mass of the deuteron (also an experimentally measured quantity) is

$$\text{Atomic mass } {}^2\text{H} = 2.014102 \text{ u}.$$

The mass difference = $2.01649 - 2.014102 \text{ u} = 0.002388 \text{ u}$. Since the conversion between rest mass and energy is 931.494 MeV/u , a deuteron's binding energy is calculated to be

$$0.002388 \text{ u} \times 931.494 \text{ MeV/u} = 2.224 \text{ MeV}.$$

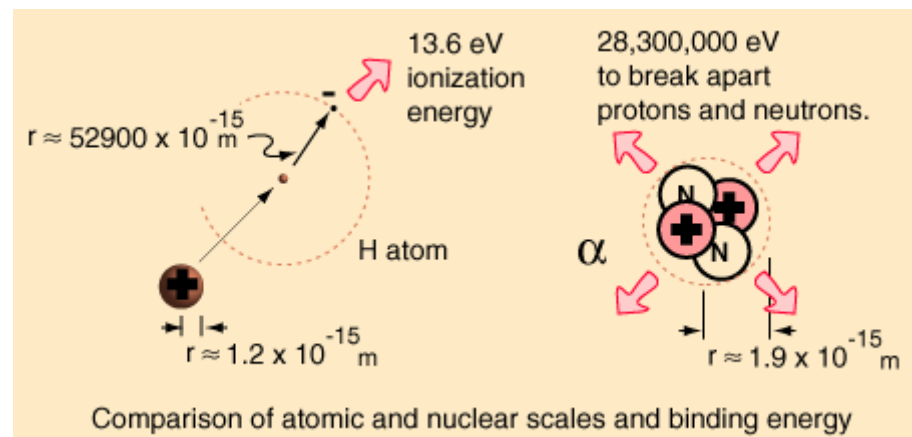
Thus, expressed in another way, the binding energy is $[0.002388/2.01649] \times 100\% = \text{about } 0.1184\%$ of the total energy corresponding to the mass. This corresponds to $1.07 \times 10^{14} \text{ J/kg} = 107 \text{ TJ/kg}$.



For the alpha particle $\Delta m = 0.0304$ u which gives a binding energy of 28.3 MeV.

	protons	2×1.00728 u		Alpha particle
	neutrons	2×1.00866 u		
	Mass of parts	<u>4.03188</u> u	Mass of alpha	4.00153 u
$1 \text{ u} = 1.66054 \times 10^{-27} \text{ kg} = 931.494 \text{ MeV}/c^2$				

The enormity of the nuclear binding energy can perhaps be better appreciated by comparing it to the binding energy of an electron in an atom. The comparison of the alpha particle binding energy with the binding energy of the electron in a hydrogen atom is shown below. The nuclear binding energies are on the order of a million times greater than the electron binding energies of atoms.



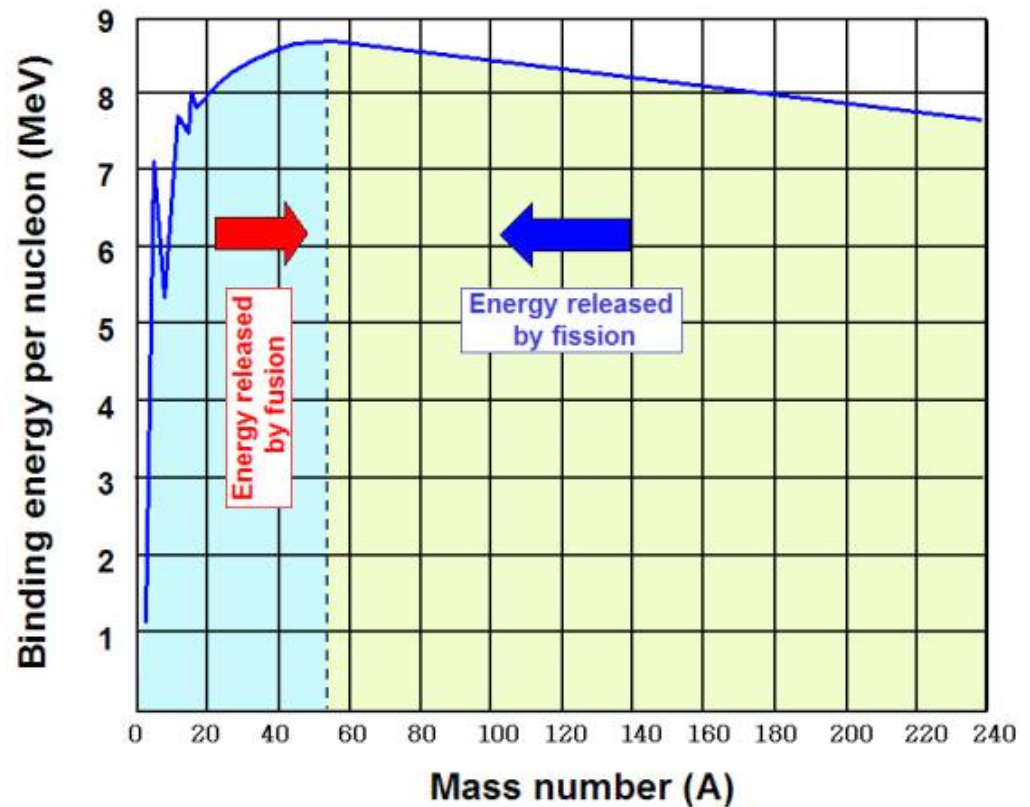


Figure 2.6. Nuclear binding energy curve

In the periodic table, elements, exhibit a general increase in binding energy per nucleon as the atomic mass increases. This increase is generated by increasing forces per nucleon in the nucleus, as each additional nucleon is attracted by all of the other nucleons, and thus more tightly bound to the whole.

The region of increasing binding energy is followed by a region of relative stability (saturation) in the sequence from magnesium through xenon. In this region, the nucleus has become large enough that nuclear forces no longer completely extend efficiently across its width. Attractive nuclear forces in this region, as atomic mass increases, are nearly balanced by repellent electromagnetic forces between protons, as atomic number increases.

Finally, in elements heavier than xenon, there is a decrease in binding energy per nucleon as atomic number increases. In this region of nuclear size, electromagnetic repulsive forces are beginning to gain against the strong nuclear force.

At the peak of binding energy, nickel-62 is the most tightly-bound nucleus, followed by iron-58 and iron-56 (This is the basic reason why iron and nickel are very common metals in planetary cores, since they are produced profusely as end products in supernovae).



The existence of a *maximum* in binding energy in medium-sized nuclei is a consequence of the trade-off in the effects of two opposing forces which have different range characteristics. The attractive nuclear force (strong nuclear force), which binds protons and neutrons equally to each other, has a limited range due to a rapid exponential decrease in this force with distance. However, the repelling electromagnetic force, which acts between protons to force nuclei apart, falls off with distance much more slowly (as the inverse square of distance). For nuclei larger than about four nucleons in diameter, the additional repelling force of additional protons more than offsets any binding energy which results between further added nucleons as a result of additional strong force interactions; such nuclei become less and less tightly bound as their size increases, though most of them are still stable. Finally, nuclei containing more than 209 nucleons (larger than about 6 nucleons in diameter) are all too large to be stable, and are subject to spontaneous decay to smaller nuclei.

Effective Nuclear charge

Definition: The effective nuclear charge, also known as the kernel charge, is the net positive charge experienced by an electron in a multi-electron atom. The term «effective» is used because the shielding effect of negative electrons prevents higher orbital electrons from experiencing the full nuclear charge.

In an atom with one electron, that electron experiences the full charge of the positive nucleus. In this case, the electron experiences the charge as it is from the nucleus without any interference. However, in an atom with many electrons the outer electrons are simultaneously attracted to the positive nucleus and repelled by the negatively charged electrons. The effective nuclear charge on such an electron is given by the following equation:

$$Z_{eff} = Z - S$$

Where ; Z is the number of protons in the nucleus and S is the average number of electrons between the nucleus and the electron in question,

Note: Z_{eff} is also often known as « Z^* ».

A simple way to calculate the effective nuclear charge is to take the total protons minus all electrons excluding the valence electrons.

Consider a fluorine atom with a +9 nuclear charge and electronic configuration of $1s^2 2s^2 2p^5$. The two electrons ($1s^2$) in the first energy level as they look at the nucleus feel a +9 charge because that is the charge on the nucleus. But the electrons that are in the valence energy level ($2s^2 2p^5$) would be shielded from the nucleus by the 2 shielding electrons. The +9 nuclear charge is shielded by 2 electrons to give an effective nuclear charge of +7 that is felt by the valence electrons. If you get out beyond the valence electrons, then the effective charge is 0 simply because the +9 charge of the nucleus is surrounded by 9 electrons.



Trends

Down the Periodic Table (Top to Bottom), effective Nuclear Charge for the outermost valence electron remains constant moving down a column or family.

Across the Periodic Table (Left to Right)

Effective Nuclear Charge increases, going across the Periodic Table. Elements have more protons and the same number of core electrons.

The increase in the electrostatic force provided by the protons is greater than the increase in repulsion due to shielding and distance. Therefore, across the Periodic Table, Effective Nuclear Charge increases.

These trends can be used to explain other trends between elements. Properties such as atomic radius, Melting and Boiling Points, electronegativity and ionization energy can be discussed using the Effective Nuclear Charge theory.

Generally, as Z_{eff} increases:

- 1) Atomic Radii decrease (Greater pull on electrons means that they are held on «tighter») but only as you move from the left to the right of the periodic table; atomic radii increase from top to bottom because of added shells (despite the decreasing Z_{eff})
- 2) Melting and Boiling Points increase (Higher intermolecular forces as a result of high Z_{eff} means that more energy is needed to break intermolecular forces)
- 3) Electronegativity increases (Directly related to Z_{eff})
- 4) Ionization energy increases (If Z_{eff} is the net positive force pulling force experienced by an electron, and the ionization energy is the energy required to remove the outer most electron, the more Z_{eff} the more energy required to remove it).

It is important to bear in mind that when speaking about the location of electrons with respect to the nucleus, we are talking about a probability (see Heisenberg's uncertainty principle). The actual effective nuclear charge is therefore in fluctuation due to the location of an electron relative to the nucleus, and also due to electron-electron repulsions. Any theorized effective nuclear charge is a kind of mean between larger and smaller positive charges experienced during the path of an electron.

Practice

Find out what the effective nuclear charge would be for neon and sodium for the electrons in each energy level.

- a) What charge is felt by the electrons in the first level of a neon atom?
- b) What charge is felt by the electrons in the second level (valence level) of a neon atom?
- c) What charge is felt by the electrons in the first level of a sodium atom?



- d) What charge is felt by the electrons in the second level of a sodium atom?
e) What charge is felt by the electrons in the third (valence) level of a sodium atom?

Answers; a) 10, b) 8, c) 9, d) 1.

Problem set Unit 2 (accounts for 15 %).

- Determine the number of valence electrons for the elements in group II:
a) 1, b) 2, c) 5 d) 8
- What group or family of elements (IUPAC criteria) does silicon belong to?
a) 3, b) 14, c) IVA, d) 4.
- What is the name of the Group IIA elements?
a) alkali metals b) alkaline-earth metals c) halogens, d) Noble metals
- Which of the following is an inert gas?
a) hydrogen b) krypton c) chlorine, d) Nitrogen
- In the Periodic Trends, which element has the smallest atomic radius?
a) Mg, b) Ca, c) Sr, d) Be
- What group of elements on the periodic table has the greatest atomic size?
a) The halogens b) The noble gases c) The alkali metals d) p-block elements
- Which of the following sets are isoelectronic with each other? (T).
a) Rb^+ , Sr^{+2} , Xe b) Al^{+3} , Na^+ , Ne c) Br, Na^+ , Se^{-2}
- Arrange the atoms and ions in order of increasing size, Rb^+ , Na^+ , and K^+ .
a) $\text{Na}^+ < \text{K}^+ < \text{Rb}^+$ b) $\text{Rb}^+ < \text{K}^+ < \text{Na}^+$ c) $\text{Na}^+ < \text{Rb}^+ < \text{K}^+$



9. Which is the electronic configuration of a ${}_{1}\text{H}^3$ atom in the ground state?
a) $1s^1$, b) $1s^2$, c) $1s^22s^1$, d) $1s^22s^2$.
10. In which sub shell will the electrons have the highest energy?
a) 3p, b) 2p, c) 3s, d) 4s.
11. Which atom will form the most polar bond with hydrogen?
a) F, b) Cl, c) Br, d) I.
12. Which electron configuration represents the atom in the period 2 with the largest covalent atomic radius
a) $1s^22s^1$, b) $1s^12s^2$, c) $1s^22s^22p^1$, d) $1s^22s^22p^2$
13. The elements of period 3 that has the highest ionization energy is;
a) a noble gas, b) an alkali metal, c) halogen, d) an alkali earth metal.
14. Magnesium has a smaller atomic radius than Sodium because the magnesium atom has more,
a) Valency electrons, b) energy levels, c) Protons, d) neutrons.
15. As one proceeds from the left to the right across period 2, ionization energy generally, a) decreases and atomic sizes also decrease, b) decrease and atomic size increase, c) increase and atomic size decrease, d) increase and atomic size increase.

Answer key

1. b, 2. b, 3. b, 4. b, 5. d, 6. c, 7. b, 8. a, 9. a, 10. d, 11. a, 12. d, 13. a, 14. c, 15. c.



Learning activity # 3

Title of Learning Activity : **GENERAL PROPERTIES OF s AND p BLOCK ELEMENTS**

At the end of this Unit, the learner should be able to;

1. Predict the valency and oxidation states of the s-, and p-block elements
2. State the name(s) of simple s-, and p-block elements
3. Predict the occurrence of inert pair effect
4. State and predict the general behaviour of metals, non-metals, and the metalloids
5. Predict and explain the trend in the chemical and physical properties of s-block elements
6. Predict and explain the trend in the chemical and physical properties of p-block elements
7. Explain, and appreciate the anomalous behaviour of the first elements of a group
8. State and give example of diagonal relationships in the periodic table

Summary of the learning activity

In the first two Units of this module, we placed and monitored some trends in properties of the s-, and p-block elements of the periodic table. In this Unit and in furtherance of the above discussed trends, we monitor the physical and chemical properties of the s-, and p-block elements. Calculation of the common oxidation state, describing inert pair effects, and explaining the anomalous behaviour some elements will follow. Diagonal relationship is one subject that most students find fascinating, it will be considered before tackling relevant worked examples and exercises at the end of the Unit.

List of Required Readings

1. John C. Kotz and Paul Trichel, Jr. *Chemistry & Chemical reactivity, 3rd edition*. Saunders college publishing, NewYork, USA. (1996).
2. Alan G. Sharpe; *Inorganic Chemistry, 3rd Edition*. Longman Singapore Publisher (1992).
3. Catherine E. Housecroft and Alan G. Sharpe; *Inorganic Chemistry*. Prentice-Hall International, USA. (2000).
4. J. D. Lee, *Concise Inorganic Chemistry, 4th edition*. Chapman & Hall, New York. USA. (1993).



5. Thomas R. Gilbert, Rein V. Kirss, and Geoffrey Davies; *Chemistry, The science in context*. W.W. Norton and company NY, USA.(2004).
6. William L. Jolly, *Modern inorganic Chemistry 2nd Ed*. McGraw-Hill. Inc. New York, USA (1991).

List of relevant useful links

<http://chemistry.about.com/od/elementgroups/a/metals.htm>

Outlines what properties distinguishes metals from the other elements.

<http://www.docbrown.info/page01/ElCpdMix/ElCmdMix3.htm>

Valency of elements and compounds formation.

<http://chemistry.about.com/od/elementgroups/a/alkalimetals.htm>

Gives properties for alkali metals

<http://www.infoplease.com/ce6/sci/A0859586.htm>

Properties of metals

<http://chemistry.about.com/od/elementgroups/a/alkalineearths.htm>

About Alkali earth metals

http://nobel.scas.bcit.ca/chem0010/unit4/4.3.2_property_nonmetals.htm

Properties of non-metals

<http://en.wikipedia.org/wiki/Chalcogen>

For chalcogens and their details.

http://www.chemsoc.org/viselements/pages/data/intro_groupvii_data.html

For halogens and their properties.

<http://chemistry.about.com/od/elementgroups/a/noble gases.htm>

For noble gases

<http://www.google.de/search?client=firefox-a&rls=org.mozilla%3Aen-US%3Aofficial&channel=s&hl=de&q=chemical+properties+of+group+13+elements&meta=&btnG=Google-Suche> (pdf version)

http://www.chemsoc.org/viselements/pages/data/intro_groupiii_data.

Data on general properties of elements of group 13.



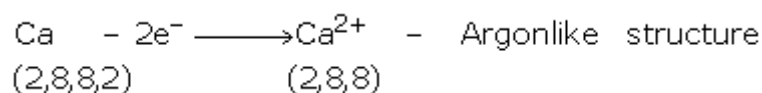
List of relevant MULTIMEDIA resources

- Computer with internet connecting facility to access relevant links and free source resources.
- Multi-media resources such as CD players, VCD etc.
- CD-ROM for this module for compulsory reading and demonstrations.

Learning activities

Valancy and formulae of compounds

Group IA, IIA and IIIA consists of metals of normal elements. They have 1, 2, or 3 valence electrons, which they can easily donate to acquire the configuration of nearest noble gas. They become positively charged i.e., cations.



Covalent compounds are formed by the sharing of electrons by atoms of non-metallic elements for example between groups IVA and VIIA. The elements of these groups react with one another by sharing electron pairs. Thus, the valency of an element in a covalent molecule is equal to the number of electron pairs shared by an atom of the element. Normally metals donate electrons from their valence shell so as to form positively charged ions such that the charge on the ion is equal to its electropositive valency.

Formula of Ionic Compounds

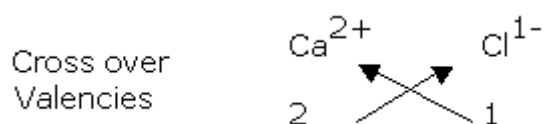
A chemical compound is always electrically neutral. The positive and negative valencies of the ions/radicals in a compound are equal and balanced.



Steps

- 1) Write the symbol of the ions of the compound.
- 2) Cation (positive ion) is written on the left hand side and the anion (negative ion) is written on the right hand side.
- 3) Put the valency number of the radical or ion below the symbol of the element.
Alternatively cross the positive and negative charges on ions to give subscripts.

Example: Calcium Chloride $\text{Ca}^{2+} \text{Cl}^{-}$ Valencies 2 1

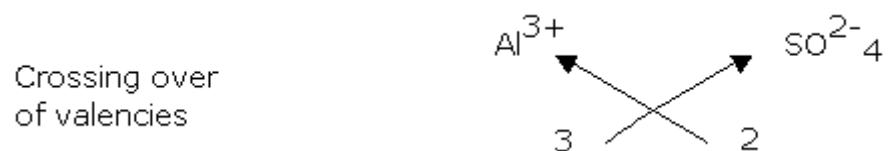


Formula CaCl_2 .

Ensure that the charges are balanced. $1(\text{Ca}^{2+}) = +2$; $2(\text{Cl}^{-}) = -2$

- 4) Radicals need to be enclosed within brackets if their number exceeds one.

Example: Aluminium Sulphate $\text{Al}^{3+} \text{SO}_4^{2-}$ Valency 3 2



$2(\text{Al}^{3+}) = +6$; $3(\text{SO}_4^{2-}) = -6$ $\text{Al}_2(\text{SO}_4)_3$.

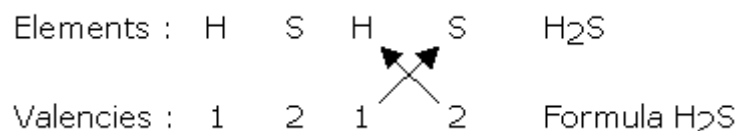


For Molecular compounds.

Formula of Molecular Compounds

Steps

- 1) Write the symbols of the elements which form the compound.
- 2) Below the symbol of each element write down its valency.
- 3) Cross over the valencies of the atoms.



Note: covalent compounds are formed by the combination between two different non-metals. While writing the formula the less electronegative nonmetal element is written on the left hand side, whereas, the more electronegative non-metal is written on the right hand side. Example: HCl

Naming of Simple Compounds

Compounds formed by the combination of two different elements are called binary compounds. Example: CO, CCl₄, H₂S, NaCl, NH₃, HBr. Binary compounds may be ionic or molecular.

Rules for naming a chemical compound.

Rule 1 The name of the element that occurs first in the compound is written first without any change in its spelling. In naming MgO, the first part of the formula is magnesium. If the metal has a variable valency the Roman numeral is included. For example: PbO would be named as Lead (II) oxide. The name of the second element in the formula is written last and is modified to end in *ide*. For example: MgO is named Magnesium oxide.

Rule 2 When two nonmetals combine covalently there is a strong chance that more than one binary compound will be formed. Example: when nitrogen and oxygen combine depending on the experimental conditions it can form any of the following compounds N₂O, NO, N₂O₃, NO₂, N₂O₄, N₂O₅. In such cases the proportion of the various elements are given by the Greek prefix like mono, di, tri, tetra, penta etc. N₂O₅ - Di-nitrogen pentoxide NO₂ - Nitrogen dioxide

Rule 3 Compounds containing three elements (ternary compounds), one of which is oxygen are named with suffix ate at the end, provided there is only one such compound. If there are two compounds, the one with more oxygen is named with the suffix ate



and the one with less oxygen is named with -ite ending. Examples: i) a) NaNO_3 - Sodium nitrate b) NaNO_2 - Sodium nitrite ii) a) CaSO_4 - Calcium sulphate b) CaSO_3 - Calcium sulphite iii) a) KClO_3 - Potassium chlorate b) KClO_2 - Potassium chlorite iv) a) $\text{Ca}_3(\text{PO}_4)_2$ - Calcium phosphate b) $\text{Ca}_3(\text{PO}_3)_2$ - Calcium Phosphite

Rule 4 If in a compound, oxygen is less than the oxygen present in a compound ending with ite then, it is given the prefix hypo- in the beginning and if oxygen present in a compound ending with ate is more then it is given the prefix per- in the beginning. Example: KClO - potassium hypochlorite KClO_2 - potassium chlorite KClO_4 - potassium per chlorate KClO_3 - potassium chlorate

Charateristics of metals and non-metals

Metals

Metals fall into groups in the periodic table determined by similar arrangements of their orbital electrons and a consequent similarity in chemical properties.

Summary of Common Properties of metals are:

1. shiny metallic luster
2. conduct electricity and heat
3. malleable so they can be hammered into different shapes
4. ductile, so they can be drawn into thin wire
5. high in density. Of all the metals, lead is the highest in density
6. metals have high melting points and most are solids at room temperature
7. mercury is the only metal that is a liquid at room temperature
8. cesium and gallium have low melting points; (melting points are less than 30°C)
9. metals do not combine chemically with each other, but react with nonmetals to form compounds
10. of all the metals, Group 1 (i.e. alkali metals) are the most reactive
11. metals near the center of the periodic table are less reactive
12. less reactive metals are more likely to be found in nature as uncombined elements (i.e. copper, silver, gold)
13. reactive metals such as the alkali metals are never found as uncombined elements (i.e. lithium, sodium)

Many of the properties of metals, including large atomic radius, low ionization energy, and low electronegativity, are due to the fact that the electrons in the valence shell of a metal atoms can be removed easily. One characteristic of metals is their ability to be deformed without breaking. Malleability is the ability of a metal to be hammered into shapes. Ductility is the ability of a metal to be drawn into wire. Because the valence electrons can move freely, metals are good heat conductors and electrical conductors.



Chemical Properties

Chemically, the metals differ from the nonmetals in that they form positive ions and basic oxides and hydroxides. Upon exposure to moist air, a great many undergo corrosion, i.e., enter into a chemical reaction; e.g., iron rusts when exposed to moist air, the oxygen of the atmosphere uniting with the metal to form the oxide of the metal. Aluminum and zinc do not appear to be affected, but in fact a thin coating of the oxide is formed almost at once, stopping further action and appearing unnoticeable because of its close resemblance to the metal. Tin, lead, and copper react slowly under ordinary conditions. Silver is affected by compounds such as sulfur dioxide and becomes tarnished when exposed to air containing them. The metals are combined with nonmetals in their salts, as in carbides, carbonates, chlorides, nitrates, phosphates, silicates, sulfides, and sulfates.

The Electromotive Series

(Kindly look up the series in any inorganic text and study it)

On the basis of their ability to be oxidized, i.e., lose electrons, metals can be arranged in a list called the electromotive series, or replacement series. Metals toward the beginning of the series, like cesium and lithium, are more readily oxidized than those toward the end, like silver and gold. In general, a metal will replace any other metal, or hydrogen, in a compound that it precedes in the series, and under ordinary circumstances it will be replaced by any metal, or hydrogen, that it follows.

Non-Metals

The Properties of the Nonmetals

1. nonmetals have low densities.
2. nonmetals are not ductile or malleable and brittle.
3. nonmetals are not lustrous and are dull.
4. nonmetals are poor conductors of heat and electricity except carbon in the form of graphite.
5. nonmetals can be found in nature in the uncombined elemental form
6. fluorine is the most reactive of all the nonmetals.
7. nonmetals can combine with other nonmetals to form compounds such as: carbon dioxide (CO_2), carbon monoxide (CO), methane (CH_4), and ammonia (NH_3).

Metalloids / semimetals

The Properties of the Semimetals

1. semimetals have both metallic and nonmetallic properties
2. semimetals are semiconductors, in other words, they are not as conducting as metals and are not insulators like the nonmetals (i.e. silicon and gallium arsenide (GaAs) are used widely in the semiconductor industry)



Block elements

Alkali metals

The alkali metals are lithium, sodium, potassium, rubidium, cesium, and francium (Li, Na, K, Rb, Cs, and Fr respectively). Hydrogen is often considered to be an alkali metal (especially at low temperatures or high pressures), although some scientists think of it as a nonmetal. The alkali metals are the elements located in Group IA (first column) of the periodic table.

In many of its properties, lithium, exhibits the same characteristics as do the more common alkali metals sodium and potassium; thus, the metal is highly reactive and forms strong hydroxide solutions. In many respects lithium also shows similarities to the elements of the alkaline-earth group, especially magnesium, which has similar atomic and ionic radii.

Summary of Common Properties

1. Lower densities than other metals,
2. One loosely bound valence electrons
3. Largest atomic radii in their periods,
4. Low ionization energies,
5. Low electronegativities,
6. Highly reactive

Properties

The alkali metals exhibit many of the physical properties common to metals, although their densities are lower than those of other metals. Alkali metals have one electron in their outer shell, which is loosely bound. This gives them the largest atomic radii of the elements in their respective periods. Their low ionization energies result in their metallic properties and high reactivities. It is easy to remove valence electrons of an alkali metal to form a univalent cation. Alkali metals have low electronegativities. They react readily with nonmetals, particularly halogens.

Physical Properties

1. They are grey and soft, and can be easily cut with a knife to expose a shiny surface which dulls on oxidation.
2. They have low densities - Li, Na and K are less dense than water. They have low standard enthalpies of melting and vaporization. They show relatively weak metallic bonding as only one electron is available from each atom.
3. Good conductors of heat and electricity
4. They have low standard enthalpies of melting and vaporization. They show relatively weak metallic bonding as only one electron is available from each atom.
5. Alkali metals colour flames. When the element is placed in a flame the heat provides sufficient energy to promote the outermost electron to a higher energy



level. On returning to ground level, energy is emitted and this energy has a wavelength in the visible region

Li red Na yellow K lilac Rb red Cs blue

- The ionic radii of the alkali metals are all much smaller than the corresponding atomic radii. This is because the atom contains one electron in an s level relatively far from the nucleus in a new quantum shell, and when it is removed to form the ion the remaining electrons are in levels closer to the nucleus. In addition, the increased effective nuclear charge attracts the electrons towards the nucleus and decreases the size of the ion.

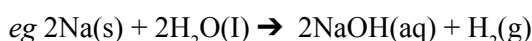
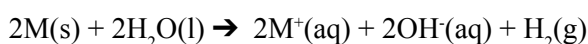
Chemical Properties

- Alkali metals react with water, oxygen, and other chemicals.
- Alkali metals are never found as free elements in nature (i.e. - sodium metal does not exist as Na metal in nature, sodium exists in compounds such as NaCl, Na₂CO₃, etc...). The reactivity increases on descending the Group from lithium to cesium. There is a closer similarity between the elements of this Group than in any other Group of the periodic table.
- Alkali metal compounds are water soluble and are present in seawater and salt deposits.
- All alkali metals have one valence electron which form ions with a positive (1+) charge

Storage: since alkali metals react quickly with oxygen, they are stored under mineral oil or kerosene

Chemical Properties

The alkali metals are strong reducing agents. The standard electrode potentials all lie between -2.7V and -3.0V, indicating a strong tendency to form cations in solution. They can reduce oxygen, chlorine, ammonia and hydrogen. The reaction with oxygen tarnishes the metals in air, so they are stored under oil. They cannot be stored under water because they react with it to produce hydrogen and alkali hydroxides:



This reaction illustrates the increasing reactivity on descending the Group. Li reacts steadily with water, with effervescence; sodium reacts more violently and can burn with an orange flame; K ignites on contact with water and burns with a lilac flame; Cs sinks in water, and the rapid generation of hydrogen gas under water produces a shock wave that can shatter a glass container.



Na dissolves in liquid ammonia to give a deep blue solution of sodium cations and solvated electrons. This solution is used as a reducing agent. At higher concentrations the colour of the solution changes to bronze and it conducts electricity like a metal.

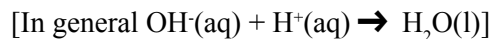
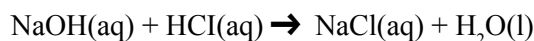
The chemistry of Li shows some anomalies, as the cation Li^+ is so small it polarises anions and so introduces a covalent character to its compounds. Li has a diagonal relationship with magnesium.

Oxides

The alkali metals form ionic solid oxides of composition M_2O when burnt in air. However, Na also forms the peroxide Na_2O_2 as the main product, and K forms the superoxide KO_2 , also as the main product.

Hydroxides

Alkali metal hydroxides are white ionic crystalline solids of formula MOH , and are soluble in water. They are all deliquescent except LiOH . The aqueous solutions are all strongly alkaline (hence the name of this Group) and therefore dangerous to handle. They neutralise acids to form salts, *eg*:



Halides

Alkali metal halides are white ionic crystalline solids. They are all soluble in water except LiF , which has a very high lattice enthalpy arising from the strong electrostatic interaction of the small Li^+ and F^- ions.

Oxidation States and Ionisation Energies

Alkali metals have oxidation states of 0 and +1. All the common compounds are based on the M^+ ion. This is because the first ionisation energy of these elements is low, and the second ionisation energy much higher. The outermost electron is well shielded from the attraction of the nucleus by filled inner electron levels and so is relatively easy to remove. The next electron is much more difficult to remove as it is part of a full level and is also closer to the nucleus.

The first ionisation energy decreases down the Group because the outermost electron is progressively further from the nucleus and so is easier to remove.



	1st	2nd	3rd
Li	513.3	7298.0	11814.8
Na	495.8	4562.4	6912.0
K	418.8	3051.4	4411.0
Rb	403.0	2632.0	3900.0
Cs	375.7	2420.0	3400.0

Solubility And Hydration

All the simple salts of the group 1 elements dissolve in water, producing ions, and consequently the solutions conduct electricity. Since the Li^+ is small, it is expected that the solutions of lithium salts would conduct electricity better than the rest. However, the trend is the opposite; $\text{Cs}^+ > \text{Rb}^+ > \text{K}^+ > \text{Na}^+ > \text{Li}^+$. The reason is that the ions are hydrated in solution. Since Li^+ is very small, it has higher attraction to the water molecules and hence more water molecules adhere to its ions making it highly hydrated. This makes the effective ionic radius large and hence it moves very slowly. The contrast happens for the Cs^+ .

For a salt to dissolve, i.e NaCl, it must break into its constituents:



If a salt is soluble, then its lattice energy is smaller than the energy of hydration. The solubility of most group 1 salts in water decreases on descending the group. This is because, the lattice energy only changes very slightly but the free energy of hydration changes drastically. For example, the differences in the lattice energy between NaCl and KCl is 67 KJ/Mol, yet the difference in free Energy of hydration is 76 KJ/Mol. Thus KCl is less soluble than NaCl.

N.B. The fluorides and carbonates have a trend that increases as one descends the group.

Difference between Lithium and the other group 1 elements

Lithium chemistry is closer to group 2 elements (more so Magnesium) than it is to the rest of group 1. In comparison, Li has higher melting point, it is harder, reacts least readily with oxygen, its hydroxide is less basic and many of its salts are less stable, and are more heavily hydrated than the rest of group 1 elements.

In general, you will realise that first elements (Li, Be, B, C, N, O, and F) in almost all groups differ from the rest. This is partly because the first element is always much



smaller and is more likely to form covalent compounds (Fajan's rule).

The similarity between the Li of group 1 and Mg of group 2 is called ***diagonal relationship***.

Such a relationship also exists between other other pairs of elements; Be and Al, and B and Si. This relationship a rises partly because of the similarities if ionic sizes between the pairs and majorly is due to similarities in their electronegativities.

Alkali Earth metals

The members of alkaline earth metals group are: Be, Mg, Ca, Sr, and Ba. The alkaline earth family is the second most reactive group. They are called alkaline earth metals because they form "alkaline" solutions (hydroxides) when they react with water. This term "alkaline" means that the solution has a pH greater than seven and is basic. Thus the alkaline earth metals form very basic solutions and are excellent reducing agents.

The alkaline earths have two electrons in the outer shell. They have smaller atomic radii than the alkali metals. The two valence electrons are not tightly bound to the nucleus, so the alkaline earths readily lose the electrons to form divalent cations. Alkaline earths have low electron affinities and low electronegativities. As with the alkali metals, the properties depend on the ease with which electrons are lost.

Summary of Common Properties

Physical Properties:

- low densities, but higher than densities of the comparable alkali metal.
- Generally, alkaline earth metals are silvery white metals with high melting and boiling points.
- Are often stronger than most metals.
- All are ductile, lustrous, and malleable metals.

The metals of Group 2 are harder and denser than sodium and potassium, and have higher melting points. These properties are due largely to the presence of two valence electrons on each atom, which leads to stronger metallic bonding than occurs in Group 1.

Three of these elements give characteristic colours when heated in a flame:

Mg brilliant white Ca brick-red Sr crimson Ba apple green

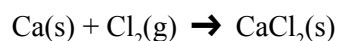
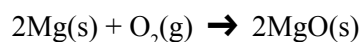


Atomic and ionic radii increase smoothly down the Group. The ionic radii are all much smaller than the corresponding atomic radii. This is because the atom contains two electrons in an s level relatively far from the nucleus, and it is these electrons which are removed to form the ion. Remaining electrons are thus in levels closer to the nucleus, and in addition the increased effective nuclear charge attracts the electrons towards the nucleus and decreases the size of the ion.

Chemical Properties

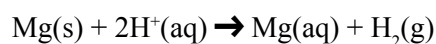
The chemical properties of Group 2 elements are dominated by the strong reducing power of the metals. The elements become increasingly electropositive on descending the Group.

Once started, the reactions with oxygen and chlorine are vigorous:



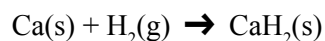
All the metals except beryllium form oxides in air at room temperature which dulls the surface of the metal. Barium is so reactive it is stored under oil.

All the metals except beryllium reduce water and dilute acids to hydrogen:

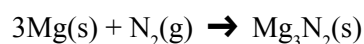


Magnesium reacts only slowly with water unless the water is boiling, but calcium reacts rapidly even at room temperature, and forms a cloudy white suspension of sparingly soluble calcium hydroxide.

Calcium, strontium and barium can reduce hydrogen gas when heated, forming the hydride:

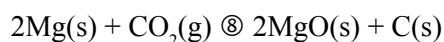


The hot metals are also sufficiently strong reducing agents to reduce nitrogen gas and form nitrides:





Magnesium can reduce, and burn in, carbon dioxide:



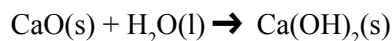
This means that magnesium fires cannot be extinguished using carbon dioxide fire extinguishers

Oxides

The oxides of alkaline earth metals have the general formula MO and are basic. They are normally prepared by heating the hydroxide or carbonate to release carbon dioxide gas. They have high lattice enthalpies and melting points. Peroxides, MO₂, are known for all these elements except beryllium, as the Be²⁺ cation is too small to accommodate the peroxide anion.

Hydroxides

Calcium, strontium and barium oxides react with water to form hydroxides:



Calcium hydroxide is known as slaked lime. It is sparingly soluble in water and the resulting mildly alkaline solution is known as lime water which is used to test for the acidic gas carbon dioxide.

Halides

The Group 2 halides are normally found in the hydrated form. They are all ionic except beryllium chloride. Anhydrous calcium chloride has such a strong affinity for water it is used as a drying agent.

Oxidation States and Ionisation Energies

In all their compounds these metals have an oxidation number of +2 and, with few exceptions, their compounds are ionic. The reason for this can be seen by examination of the electron configuration, which always has two electrons in an outer quantum level. These electrons are relatively easy to remove, but removing the third electron is much more difficult, as it is close to the nucleus and in a filled quantum shell. This results in the formation of M²⁺. The ionisation energies reflect this electron arrangement. The first two ionisation energies are relatively low, and the third very much higher.



Hydration energy

The hydration energy of the group II ions are four to five times greater than those of group I. This is largely due to their smaller sizes and increased charge. The hydration energy does decrease down the group as the size of the ions increase. In the case of Be, a further factor is the very strong complex $[\text{Be}(\text{H}_2\text{O})_4]^{2+}$ that is formed. Thus the crystalline compounds of group II contain more water of crystallization than the crystalline compounds of group I. Thus NaCl and KCl are anhydrous but $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$, $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$, and $\text{BeCl}_2 \cdot 2\text{H}_2\text{O}$ have waters of crystallization. Note that the numbers of the waters decrease as the ion becomes larger.

Solubility and Lattice energies

The solubility of salts decrease with increasing atomic weight, though the usual trend is reversed in this group with regards to fluorides and hydroxides. Generally the lattice energy decrease as the size of the ion increases. The hydration energy also decreases as the ion becomes larger. A decrease in lattice energy favours increased dissolution, but a decrease in hydration energy favours decreased dissolution. With most compounds, hydration energy changes more rapidly than lattice energy hence the compounds become less soluble as the metal gets larger. However, for fluorides and hydroxides, their lattice energies change more rapidly than hydration energy, and so their solubility increases on descending the group. This explains why Ca and Mg ions are the ones causing water hardness.



Block elements

Group 13

The Group 13 consists of five elements: boron (B), aluminum (Al), gallium (Ga), indium (In), and thallium (Tl) (Figure 1.1). These elements are characterized by having three electrons in their outer energy levels, but reflect a wide range in the occurrence and behavior. Among them, only B is metalloid and plays a significant role in plants. Aluminum, being one of the basic constituents of the lithosphere, reveals amphoteric properties.

The elements of Group 13 are:

Name	symbol	electron configuration
boron	B	[He]2s ² 2p ¹
aluminium	Al	[Ne]3s ² 3p ¹
gallium	Ga	[Ar]3d ¹⁰ 4s ² 4p ¹
indium	In	[Kr]4d ¹⁰ 5s ² 5p ¹
thallium	Tl	[Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ¹

Physical Properties

1. Boron is a non-metallic grey powder.
2. All the other members of Group 13 are soft, silvery metals.
3. Thallium develops a bluish tinge on oxidation.
4. The densities of all the Group 13 elements are higher than those of Group 2 elements.
5. The melting points of all the elements are high, but the melting point of boron is much higher than that of beryllium in Group 2, whereas the melting point of aluminium is similar to that of magnesium in Group 2.
6. Their ionic radii are much smaller than the atomic radii. This is because the atom contains three electrons in valence shell and when ionized the remaining electrons are in levels closer to the nucleus. In addition, the increased effective nuclear charge attracts the electrons towards the nucleus and decreases the size of the ion.



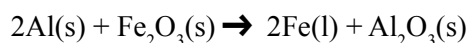
Chemical Properties

General Reactivity trend

The general trend down Group 13 is from non-metallic to metallic character. Boron is a non-metal with a covalent network structure. The other elements are considerably larger than boron and consequently are more ionic and metallic in character. Aluminium is at borderline between ionic and covalent character in its compounds. The remaining Group 13 members are generally considered to be metals, although some compounds exhibit covalent characteristics.

The chemical properties of Group 13 elements reflect the increasingly metallic character of descending members of the Group. Here only boron and aluminium will be considered.

Boron is chemically unreactive except at high temperatures. Aluminium is a highly reactive metal which is readily oxidised in air. This oxide coating is resistant to acids but is moderately soluble in alkalis. Aluminium can therefore reduce strong alkalis, a product being the tetrahydroxoaluminate ion, $\text{Al}(\text{OH})_4^-$. Aluminium also reacts violently with iron (III) oxide to produce iron in the Thermit process:



Oxides

Boron oxide, B_2O_3 , is an acidic oxide and an insoluble white solid with a very high boiling point (over 2000K) because of its extended covalently-bonded network structure. Aluminium oxide, Al_2O_3 , is amphoteric.

Halides

The most important halide of boron is boron trifluoride, which is a gas. Aluminium chloride, AlCl_3 , is a volatile solid which sublimes at 458K. The vapour formed on sublimation consists of an equilibrium mixture of monomers (AlCl_3) and dimers (Al_2Cl_6). It is used to prepare the powerful and versatile reducing agent lithium tetrahydridoaluminate, LiAlH_4 .

Both boron chloride and aluminium chloride act as Lewis acids to a wide range of electron-pair donors, and this has led to their widespread use as catalysts. Aluminium chloride is used in the important Friedel-Crafts reaction.

Hydrides

Boron forms an extensive and interesting series of hydrides, the boranes. The simplest of these is not BH_3 as expected, but its dimer B_2H_6 .

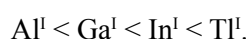


Oxidation States and Ionisation Energies

Boron and aluminium occur only with oxidation number +3 in their compounds, and with a few exceptions their compounds are best described as ionic. The electron configuration shows three electrons outside a noble gas configuration, two in an s shell and one in a p shell. The outermost p electron is easy to remove as it is furthest from the nucleus and well shielded from the effective nuclear charge. The next two s electrons are also relatively easy to remove. Removal of any further electrons disturbs a filled quantum shell so is difficult. This is reflected in the ionization energies. The first three ionization energies are low, and the fourth very much higher.

Inert pair effect

The term **inert pair effect** is often used in relation to the increasing stability of oxidation states that are 2 less than the group valency for the heavier elements of groups 13, 14, 15 and 16. As an example in group 13 the +1 oxidation state of Tl is the most stable and Tl^{III} compounds comparatively rare. The stability increases in the following sequence:



The situation in groups 14, 15 and 16 is that the stability trend is similar going down the group, but for the heaviest members, e.g. lead, bismuth and polonium both oxidation states are known.

The lower oxidation state in each of the elements in question has 2 valence electrons in s orbitals. On the face of it a simple explanation could be that the valence electrons in an s orbital are more tightly bound and of higher energy than electrons in p orbitals and therefore less likely to be involved in bonding. Unfortunately this explanation does not stand up.



Group 14

The members of this group are C, Si, Ge, Sn, and Pb.

Physical properties

1. Going across a period, ionic radii decrease, however reaching group 14 (IVA), anions decrease in size as you head towards the noble gases.
2. The atomic radii increases as we descend the group. The difference in size between the Si and Ge is less than expected because Ge has a full 3d shell which shields the nuclear charge rather ineffectively.
3. They show gradation downward. Carbon is non-metallic to Lead (Pb) whose oxides are amphoteric but the element is mainly metallic.
4. 'Diagonal line' in the p-block that separates the metals from non-metals passes through Si and Ge. Indicating that Si is more non-metallic whereas Ge is mainly metallic in behaviour. Thus both are considered metalloids.
5. The first four ionization energies suggest that the elements form +4 oxidation state
6. They exhibit +4 oxidation state though +2 becomes stable as one descends the group. In this case Pb resembles its metallic neighbours.

Chemical Properties

These elements are relatively unreactive, but the reactivity increases down the group. The M^{2+} oxidation state becomes more stable with lower members. This unreactivity is partly contributed by the surface coating of oxides or due to high overpotential of reducing the H^+ to H_2 on the elements surface.

Reactions with water

C, Si, and Ge are unaffected by water. Sn reacts with steam to give SnO_2 and H_2 . Pb is unaffected by water probably because of the oxide coating.

Reactions with acids

C, Si, and Ge are unaffected by dilute acids. Sn dissolves in dilute HNO_3 , forming $Sn(NO_3)_2$. Pb dissolves slowly in dilute HCl, forming the sparingly soluble $PbCl_2$, though very readily in HNO_3 , forming $Pb(NO_3)_2$ and oxides of nitrogen. Pb does not dissolve in dilute H_2SO_4 because of the surface coating of $PbSO_4$ is formed.

Diamond (C) is unaffected by concentrated acids, but graphite reacts with both HNO_3 and HF/HNO_3 . Si, too, is oxidized by HF/HNO_3 . Ge dissolves slowly in hot



concentrated H_2SO_4 and in HNO_3 . Sn dissolves in several concentrated acids. Pb does not dissolve in concentrated HCl because during the reaction, a surface coating of PbCl_2 is formed.

Inert pair effect

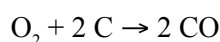
Here too, the inert pair effect shows itself increasingly with the lower members of the group. There is a decrease in the stability of the (+4) oxidation state and an increase in the (+2) state on descending the group. Ge(+2) is a strong reducing agent whereas Ge(+4) is stable. Pb(+2) is ionic, stable and more common than Pb(+4), which is oxidizing. The lower valencies are more ionic because the radius of M^{2+} is greater than that of M^{4+} and according to Fajans' rule, the smaller the ion the greater is the tendency for covalency.

Oxides and Oxoacids of Carbon and silicon

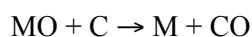
Unlike the later members of group 14, carbon forms stable, volatile monomeric oxides; CO (in limited oxygen supply) and CO_2 . Silicon is also stable as SiO_2 .

(Exercise: In discussions with any module student list all the properties of CO and CO_2 that you may remember)

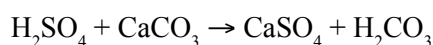
Carbon monoxide, with the chemical formula CO, is a colorless, odorless, and tasteless gas. It is the product of the partial combustion of carbon-containing compounds. Carbon monoxide has significant fuel value, burning in air with a characteristic blue flame, producing carbon dioxide. Despite its serious toxicity, CO plays a highly useful role in modern technology, being a precursor to a myriad of products.



CO also is a byproduct of the reduction of metal oxide ores with carbon, shown in a simplified form as follows:

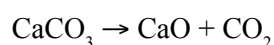


A large variety of chemical reactions yield carbon dioxide, such as the reaction between most acids and most metal carbonates. For example, the reaction between sulfuric acid and calcium carbonate (limestone or chalk) is depicted below:

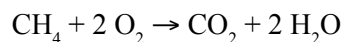




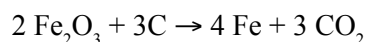
The H_2CO_3 then decomposes to water and CO_2 . Foaming or bubbling, or both accompany such reactions. In industry such reactions are widespread because they can be used to neutralize waste acid streams. The production of quicklime (CaO) a chemical that has widespread use, from limestone by heating at about 850°C also produces CO_2 :



The combustion of all carbon containing fuels, such as methane (natural gas), petroleum distillates (gasoline, diesel, kerosene, propane), but also of coal and wood, will yield carbon dioxide and, in most cases, water. As an example the chemical reaction between methane and oxygen is given below.



Iron is reduced from its oxides with coke in a blast furnace, producing pig iron and carbon dioxide:

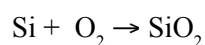


Silicon Dioxide

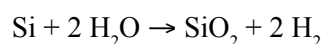
The chemical compound silicon dioxide, also known as silica or silox (from the Latin «silex»), is the oxide of silicon, chemical formula SiO_2 and has been known for its hardness since the 9th century. Silica is most commonly found in nature as sand or quartz. It is a principal component of most types of glass and substances such as concrete.

Silicon dioxide is formed when silicon is exposed to oxygen (or air). A very thin layer (approximately 1 nm or 10 Å) of so-called 'native oxide' is formed on the surface when silicon is exposed to air under ambient conditions. Higher temperatures and alternate environments are used to grow well-controlled layers of silicon dioxide on silicon.

Thermal oxidation of silicon is easily achieved by heating the substrate to temperatures typically in the range of 900-1200 degrees C. The atmosphere in the furnace where oxidation takes place can either contain pure oxygen or water vapor. Both of these molecules diffuse easily through the growing SiO_2 layer at these high temperatures. Oxygen arriving at the silicon surface can then combine with silicon to form silicon dioxide. The chemical reactions that take place are either



or so-called «dry oxidation» or





for «wet oxidation». Due to the stoichiometric relationships in these reactions and the difference between the densities of Si and SiO_2 , about 46% of the silicon surface is «consumed» during oxidation. That is, for every 1 μm of SiO_2 grown, about 0.46 μm of silicon is consumed (see Figure 1).

Consider Lead (Pb)

Lead is a dense, relatively soft, malleable metal with low tensile strength. It is a poor conductor of electricity and heat. Lead has a face-centered cubic crystalline structure. It is below tin in Group 14 of the periodic table. Although lead has a lustrous silver-blue appearance when freshly cut, it darkens upon exposure to moist air because of the rapid formation of an oxide film; the film protects the metal from further oxidation or corrosion. All lead compounds are poisonous (see lead poisoning). Lead resists reaction with cold concentrated sulfuric acid but reacts slowly with hydrochloric acid and readily with nitric acid.

The element has four naturally occurring stable isotopes, three of which result from the decay of naturally occurring radioactive elements (thorium and uranium). Since this decay takes place at a constant rate, it is possible to predict either the maximum age of a lead-containing rock or its composition at some earlier date, as long as the rock has not been chemically altered. There are 25 known radioactive isotopes of lead, some of which occur naturally in small amounts.

Group 15

The elements of this group all have five electrons in their outer most shell. They there exhibit a maximum oxidation state of five towards oxygen by using all the five outer electrons in forming bonds. The tendency for the pair of s electrons to remain inert (inert pair effect) increases with the increasing atomic weight. Thus only the p electrons are used in bonding and trivalency results. Valency of three and five are shown with halogens and sulphur.

The group shows the usual trends, that metallic character increases on descending the group. Thus N and P are non-metals, As and Sb (antimony) are metalloids, and Bi is metallic. Thus since metallic oxides are basic but those of non-metals are acidic, the oxides of N and P are acidic, whereas those of As and Sb are amphoteric and that of Bi is basic.



General properties of Nitrogen and Phosphorus

Properties of Nitrogen

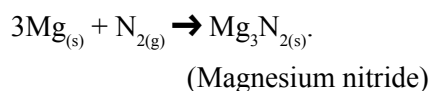
Physical

As the first element of the group, it differs from the rest. Nitrogen is a colorless gas, odorless, and a tasteless gas. The vapor density of nitrogen is 14 (Vapor density of air = 14.4). Hence it is slightly lighter than air. Nitrogen is slightly soluble in water. 1 liter of water dissolves about 22 ml of nitrogen at 0°C. When cooled and compressed nitrogen liquefies to a colorless liquid, which boils at -196°C. When further cooled, it freezes to a white snow like solid, which melts at -210°C.

Chemical

Nitrogen being a neutral gas that does not affect the color of indicators, It is neither combustible not a supporter of combustion. Nitrogen does not easily combine with other elements under ordinary conditions. A molecule of nitrogen is diatomic. These two atoms have combined by mutually sharing three pairs of electrons. The triple bond formed between the atoms is so strong, that a large amount of energy is needed to break this bond. Hence special conditions such as temperature, pressure, catalyst, promoters, high voltage etc. may be needed to make nitrogen combine with other elements.

- With hydrogen; Nitrogen combines with hydrogen in presence of electric sparks to form ammonia. Usually ammonia is obtained by treating a 3:1 mixture (by volume) of hydrogen and nitrogen at about 450°C, and a pressure of 200 to 1000 atm, in the presence of finely divided iron as catalyst and molybdenum as promoter. Promoter is a substance, added along with a catalyst to increase the efficiency of the catalyst. Promoters are also called activators.
- With oxygen; Nitrogen combines with oxygen only in presence of an electric arc at a temperature of 3000 - 5000°C to form nitric oxide. Also when lightning takes place in the sky, electric sparks of high voltage are formed that make nitrogen and oxygen of the atmosphere reactive enough to form nitric oxide.
- With metals; Red hot or burning magnesium, calcium and aluminium combine with nitrogen to form their respective nitrides (N⁻³).



A binary compound of an element with nitrogen is called a «nitride». When these nitrides react with hot water, they yield their respective hydroxides and ammonia.



With calcium carbide; When calcium carbide is heated in presence of nitrogen at 800°C to 1000°C it gives calcium cyanamide an important fertilizer. The commercial name of calcium cyanamide is nitrolim.

Properties of Phosphorus

Physical

Phosphorus is solid at room temperature. White phosphorus is soft, waxy and reactive.

Chemical

It reacts with moist air and gives out light (chemiluminescence). It ignites spontaneously in air at approximately 35°C, and is stored under water to prevent this. It is highly toxic too.

If white phosphorus is heated to approximately 250°C or at a lower temperature in the presence of sunlight, then red phosphorus is formed. This is a polymeric solid, which is much less reactive than white phosphorus. It is stable in air and does not ignite unless it is heated to over 400°C. Thus it does not need to be stored under water, but is stored in organic (paraffin) solvents.

Heating the white phosphorus under high pressure results in a highly polymeric form of black phosphorus. This is thermodynamically the most stable allotrope of phosphorus.

N.B. White phosphorus is much more toxic than red phosphorus. Necrosis of the jaw, also called phossy-jaw, is caused by chronic exposure to white phosphorus.

As is stable in dry air, but tarnishes in moist air giving first a bronze then black tarnish.

Sb is less reactive, and is stable towards water and to air at room temperature. On heating in air it forms Sb_4O_6 or Sb_4O_8 or Sb_4O_{10} .

Bi forms Bi_2O_3 on heating.



Group 16 (Chalcogens)

The **chalcogens** (with the «ch» pronounced like a k) is sometimes known as the **oxygen family**. It consists of oxygen (**O**), sulfur (**S**), selenium (**Se**), tellurium (**Te**), the radioactive polonium (**Po**), and the synthetic ununhexium (**Uuh**). The compounds of the heavier chalcogens (particularly the sulfides, selenides, and tellurides) are collectively known as **chalcogenides**. Unless grouped with a heavier chalcogen, oxides are not considered chalcogenides. The name is generally considered to mean «ore former» from the Greek *chalcos* «ore» and *-gen* «formation».

Oxygen and sulfur are nonmetals, and polonium, selenium and tellurium are metalloids semiconductors (i.e., their electrical properties are between those of a metal and an insulator). Nevertheless, tellurium, as well as selenium, are often referred to as metals when in elemental form. The formal oxidation number of the chalcogen is generally -2 in a chalcogenide but other values, such as -1 in pyrite, can be attained.

The elements have electronic structure of s^2p^4 . Thus they may attain the noble gas structure by either gaining 2 electrons, forming M^{2-} or by sharing the 2 electrons, thus forming covalent bonds. The electronegativity of Oxygen is second only to that of F. Meaning that the electronegativity difference between most metals and the oxygen is very large making most metallic oxides ionic in character. The highest formal oxidation number +6 is found in sulfates, selenates and tellurates, such as in sodium selenate (Na_2SeO_4).

Note: Modern chemical understanding based on quantum theory somewhat outdates the use of formal oxidation numbers in favour of a many-electron wavefunction approach allowing detailed computer simulation, though the concept, while flawed, is still useful in thought experiments.

Oxygen

Oxygen shows several differences with the rest of the group. These are associated with its smaller size, higher electronegativity, and lack of suitable d-orbital for bonding compared to the rest.

At room temperature and pressure, oxygen is a colorless gas, but it is pale blue when condensed to the liquid at -183°C . Diatomic oxygen (O_2) is paramagnetic because it has two unpaired electrons.

Ozone (O_3), is a second, less stable allotrope of oxygen. It is a blue, diamagnetic gas with an odor so strong that it can be detected even in very minute concentrations (< 0.05 ppm).

Sulfur

In 1817 a Swedish chemist named Jöns Jacob Berzelius noted a red substance resulting from sulfide ores from mines of Fahlun (Falun), Sweden. When this red material was investigated in the following year, it proved to be an element and was named after the Moon or the Moon goddess Selene. Compounds that contain two elements, one of which is sulfur, are called sulfides. For example, a beautiful gold-colored mineral



is called pyrite, or «fool's gold,» because it looks so much like real gold. Pyrite is iron sulfide (FeS₂).

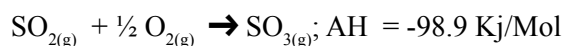
Its physical and chemical properties are very distinctive. It often occurs as a brilliant yellow powder. When it burns, it produces a clear blue flame and a very strong odor.

Sulfur has the most number of allotropes of all the elements. The most common and stable is the yellow, orthorhombic form which consists of S₈ molecules with the sulfur atoms arranged in a crown-shaped ring. Less stable allotropes have rings of 6 to 20 sulfur atoms.

Chemistry

Hydrogen sulfide, H₂S, has a structure resembling that of water. Unlike water, however, H₂S is a gas under normal conditions. This is because only very weak hydrogen bonding occurs between the molecules compared with the bonding in water. Hydrogen sulfide is a deadly poison, comparable to Hydrogen cyanide, but fortunately, H₂S has a terrible smell (of rotten egg) and can be detected in very minute concentrations.

Sulfur dioxide (SO₂) and trioxide (SO₃) are the most important oxides of the element. SO₂ is a colorless toxic gas with a choking odor. It readily dissolves in water and is easily converted to SO₃.



SO₃ is extremely reactive and is very difficult to handle, it is always deliberately converted to sulfuric acid by reactions with water.

Selenium

Selenium is a metalloid. Selenium and tellurium are often associated with each other. They tend to occur together in the Earth and have somewhat similar properties.



Group 17 (Halogens)

As a group, halogens (F, Cl, Br, I, and At) exhibit highly variable physical properties. Halogens range from solid (I_2) to liquid (Br_2) to gaseous (F_2 and Cl_2) at room temperature.

Name	symbol	electron configuration
fluorine	F	$[He]2s^2 2p^5$
chlorine	Cl	$[Ne]3s^2 3p^5$
bromine	Br	$[Ar]3d^{10} 4s^2 4p^5$
iodine	I	$[Kr]4d^{10} 5s^2 5p^5$
astatine	At	$[Xe]4f^{14} 5d^{10} 6s^2 6p^5$

The chemical properties are more uniform. The halogens have very high electronegativities. Fluorine has the highest electronegativity of all elements. The halogens are particularly reactive with the alkali metals and alkaline earths, forming stable ionic crystals.

Summary of Common Properties

1. Very high electronegativities
2. Seven valence electrons (one short of a stable octet)
3. Highly reactive, especially with alkali metals and alkaline earths

Physical Properties

1. fluorine, a poisonous pale yellow gas
2. chlorine, a poisonous greenish gas, has an irritating odor
3. bromine, a toxic and caustic red-brown volatile liquid at room temperature
4. iodine, a shiny dark grey crystalline solid at room temperature but which easily sublimates to form a violet vapour on heating

N.B. All the isotopes of astatine are radioactive, and so this element will not be considered any further here.



Chemical Properties

1. All halogens have seven valence electrons which form ions with a negative (-1) charge
2. Fluorine is the most reactive of all the nonmetals
3. Iodine sublimates when heated

In general, the Halogens are a very similar set of non-metals. They all exist as diatomic molecules, X_2 , and oxidise metals to form halides. The halogen oxides are acidic, and the hydrides are covalent. Fluorine is the most electronegative element in the periodic table. Generally, electronegativity and oxidising ability decrease on descending the Group. The result of this decreasing electronegativity is increased covalent character in the compounds, so that AlF_3 is ionic whereas $AlCl_3$ is covalent.

N.B. Fluorine shows some anomalies because of the small size of its atom and ion. This allows several F atoms to pack around a different central atom, as in AlF_6^{3-} compared with $AlCl_4^-$. The F-F bond is also unexpectedly weak because the small size of the F atom brings the lone pairs closer together than in other halogens, and repulsion weakens the bond.

The most characteristic chemical feature of the halogens is their ability to oxidise. Fluorine has the strongest oxidising ability, so other elements which combine with fluorine have their highest possible oxidation number. Fluorine is such a strong oxidising agent that it must be prepared by electrolysis. Chlorine is the next strongest oxidising agent, but it can be prepared by chemical oxidation. Most elements react directly with chlorine, bromine and iodine, with decreasing reactivity going down the Group, but often the reaction must be activated by heat or UV light. Chlorine, bromine and iodine disproportionate in the presence of water and alkalis.

Oxides and Oxoacids

There are no fluorine oxides as F is more electronegative than oxygen (O). Chlorine, bromine and iodine each form several oxides which are thermally unstable, such as chlorine dioxide ClO_2 . The only fluorine oxoacid, HOF, is unstable at room temperature, but there are many oxoacids of the other halogens. The best known salts of these are; hypochlorite, chlorate(I) ClO^- , chlorite, chlorate(III) ClO_2^- , hypochlorate, chlorate(V) ClO_3^- , perchlorate, chlorate(VII) ClO_4^- . These are all powerful oxidising agents.

Halides

The halogens can combine with each other to form interhalogens and polyhalide ions. Polyhalide ions have the general formula $[Y-X-Y]^-$. It is not possible for F to represent X in a polyhalide ion as it cannot expand its octet.



Hydrides

Hydrogen halides have the general formula HX. HF is a colourless liquid which boils at 19.5°C , and all the other hydrogen halides are colourless gases. HF is a liquid due to the extensive hydrogen bonding which occurs between molecules. All the hydrogen halides dissolve easily to give acidic solutions, the most widely used being hydrochloric acid, HCl. All except HF are typical acids; they liberate carbon dioxide from carbonates and form salts with basic oxides. HF is a weak acid because the H-F bond is very strong, and because hydrogen-bonding occurs between F^- and HF in solution.

Organic Compounds

The halogens form organic compounds which are best known for their industrial and environmental impact, such as PVC, DDT and TCP.

Oxidation States and Electron Affinities

Fluorine in all its compounds has an assigned oxidation number of -1, as it is the most electronegative of all the elements. The other halogens show a wide range of oxidation numbers, and the redox chemistry of these halogens is important. The oxidation numbers most commonly shown are odd; there are few compounds with even oxidation numbers and they are often thermally unstable. Chlorine is the third most electronegative element after F and O. The halide ions are readily formed by accepting one electron, as this completes an octet of valence electrons. The electron affinity decreases on descending the Group



NOBLE/INERT GASES (Group 0)

The noble gases (He, Ne, Ar, Kr, Xe, and Rn) also known as the inert gases, are located in Group VIII of the periodic table.

Properties

The noble gases are relatively non reactive. This is because they have a complete valence shell. They have little tendency to gain or lose electrons. The noble gases have high ionization energies and negligible electronegativities. The noble gases have low boiling points and are all gases at room temperature.

Summary of Common Properties

Physical Properties:

1. **Helium** has a boiling point of 4.2K under 1 atm pressure, the lowest boiling point of any substance.
2. When an electric current is passed through a tube which contains **Neon** gas at a low pressure gas, it has a red-orange glow.
3. **Argon** is colourless, odourless and tasteless. It makes up 0.934% of air.
4. **Krypton** is a dense, colorless, odorless atmospheric gas.
5. **Xenon** is a heavy, unreactive stable gas. When Xenon atoms are sufficiently energized, they produce a brilliant white light.

Chemical properties:

1. noble gases exist as single, monatomic gases atoms
2. all noble gases have eight valence electrons, a completely filled outer energy level of electrons ('filled' outer shell means two electrons in the outer shell for He and eight electrons in the outer shell for the others). This stable electron arrangement accounts for the nonreactive nature of the noble gases.

Problem set Unit 3 (accounts for 20 %).

1. $1s^2 2s^2 2p^6 3s^2 3p^4$, is the electronic configuration that belongs to which period and group of the periodic table of elements?
A) 3, 4, b) 4, 16, c) 3, 16, d) 4, 3.
2. Which element in the period 3 has both metallic and nonmetallic properties
a) Na, b) Al, c) Si, d) Ar.
3. Which element is more likely to form a compound with Krypton,
a) F, b) Cl, c) Xe, I, Br.



4. Which group contains an element that is a liquid at room temperature,
a) 1, b) 7, c) 8, d) 18.
5. In which group do elements usually form oxides with the formula M_2O_3 ?
a) 1, b) 3, c) 13, d) 2.
6. Which group contains two metalloids
a) 2, b) 14, c) 16, d) 18
7. Which element in group 15 is most metallic?
a) N, b) P, c) Ti, d) Bi
8. What is the total number of elements in period 2 that are all gases at STP,
a) 1, b) 2, c) 4, d) 8
9. Which represents the electronic configuration of a metalloid in ground state
a) 2-3, b) 2-5, c) 2-8-5, d) 2-8-6
10. Which element in group 14 will more likely to form Chlorides of the formula MCl_2 .
a) C, b) Si, c) Sn, d) Pb
11. To what oxidation state would we assign the elements X and Y respectively in the table below?
- | | | | | | | |
|----------|-------|-------|-------|-------|--------|--------|
| <u>X</u> | 786 | 1,577 | 3,228 | 4,354 | 16,100 | |
| <u>Y</u> | 1,060 | 1,890 | 2,905 | 4,950 | 6,270 | 21,200 |
- a) 2 & 3, b) 4 & 5, c) 5 & 6, d) 1 & 1.
12. By Paulings electronegativity method, what is the likely bond type for molecule AB, where bond energies for A-B = 3.8, A-A = 4.5 and B-B = 2.7 eV?.
a) moderately ionic, b) pure covalent, c) ionic, d) covalent.
13. Inert pair effect is likely to be experienced in,
a) group 1, b) upper group 2, c) lower group 14, d) lower group 17.



14. Calculate the effective nuclear charge for a sulphur atom,
a) 16, b) 8, c) 6, d) 10.
15. What is the electronic configuration for bromine atom in ground state?,
a) $[\text{Kr}]4s^24p^5$, b) $[\text{Ar}]4s^24p^7$, c) $[\text{Ar}]4s^24p^5$, d) $[\text{Ar}]3s^23p^5$.
16. A polar bond is more likely to be formed between,
a) Na and Cl, b) Mg and O, c) Na and F. d) H and O,
17. Electronegativity falls down a group because,
a) the atoms become bigger
b) more protons are added
c) the atomic radius becomes less defined
d) more electrons are added.
18. Aluminium flouride is ionic but aluminium iodide is covalent this is more likely because,
a) Flourine is more electronegative than iodine,
b) iodine is bigger hence easily polarized by aluminium,
c) iodine is more of a metal than flourine,
d) aluminium is so big that it polarizes flourine.
19. Most naturally abundant elements have,
a) high electronegativity,
b) higher melting points,
c) higher nuclear charge,
d) higher nuclear binding energy.
20. As the atomic mass number increases, the nuclear binding energy,
a) increase,
b) increases then decreases,
c) remains constant,
d) increases then remains constant.

Answer key

1.c 2. c 3. a 4. d 5. c 6. b, 7. d, 8. c, 9. a, 10. d, 11.b, 12. a, 13. c, 14. c, 15. c, 16. d, 17. a, 18. b, 19. d, 20. b



Learning activity # 4

Title of Learning Activity : **OCCURRENCE, ABUNDANCE AND EXTRACTION**

At the end of this Unit, the learner should be able to

1. Appreciate the natural abundances of the elements.
2. State, predict, and explain the methods and ways of extraction of s-block elements
3. State, predict, and explain the general methods and of extractions of the metallic p-block elements according the electrochemical series.
4. State, predict, and explain the general methods of extraction of non-metallic p-block elements.
5. State, predict and explain the general methods of extraction and purification of the p-block gaseous elements.
6. Carry out a research project based on the extraction of some s- and p-block elements and write a report.

Summary of the learning activity

So far we have considered the positions and trends in properties for the s-, and p-block elements of the periodic table. The current unit branches to occurrences, abundance and extractions methods compounds of the s-, and p-block elements. The unit will consider the elements as a group in the periodic table with a few representative members of the group considered. A general extraction trend, using the position of the element in the electrochemical series, will be appreciated. The extraction of the gaseous elements especially in the atmosphere will also be discussed.

List of Required Readings

1. Alan G. Sharpe; *Inorganic Chemistry, 3rd Edition*. Longman Singapore Publisher (1992).
2. William L. Jolly, *Modern inorganic Chemistry 2nd Ed*. McGraw-Hill. Inc. New York, USA (1991).
3. Catherine E. Housecroft and Alan G. Sharpe; *Inorganic Chemistry*. Prentice-Hall International, USA. (2000).
4. F. Albert Cotton and Geoffrey Wilkinson, F. R. S. *Advanced Inorganic Chemistry, A Comprehensive text, 6th edition*. Wiley Eastern Limited, New Delhi. India. (1984).
5. J. D. Lee, *Concise Inorganic Chemistry, 4th edition*. Chapman & Hall, New York. USA. (1993).



6. Thomas R. Gilbert, Rein V. Kirss, and Geoffrey Davies; *Chemistry, The science in context*. W.W. Norton and company NY, USA.(2004).
7. Carole H. McQuarrie, Donald A. McQuarrie and Peter A. Rock. *General Chemistry, 3rd edition*. W. H. Freeman and Company, New York, USA, (1990).

List of relevant useful links

- <http://www.infoplease.com/ce6/sci/A0859587.html> Useful for the introduction to metal extractions
- <http://www.citycollegiate.com/sblock1.htm>
- <http://www.wissensdrang.com/auf1na1.htm>
- www.citycollegiate.com . For free computer courses and notes
- http://en.wikipedia.org/wiki/Frasch_process For sulphur process
- <http://www.chemsoc.org/viselements/pages/fluorine.html> details on fluorine

List of relevant MULTIMEDIA resources

- Computer with internet connecting facility to access relevant links and free source resources.
- Multi-media resources such as CD players, VCD etc.
- CD-ROM for this module for compulsory reading and demonstrations.

Learning activities

Occurrence, Abundance and Extraction.

A solid element or compound which occurs naturally in the Earth's crust is called a mineral. A mineral which contains a high enough percentage of a metal for economic extraction is called a metal ore. The most common metal ores are oxides and sulphides. Sulphides are the oldest ores, formed in the Earth's history when there was a lot of sulphur from volcanic activity. Oxides formed later when photosynthesis in plants released large amounts of oxygen into the atmosphere. Metal ore deposits are a finite resource (there are only a certain amount of them) and non-renewable. Many metals are obtained today from recycling (melting and refining) scrap metals. About half of the aluminium, copper, lead, steel and tin which are used in the UK today come from recycled scrap metal.

A metal above carbon in the reactivity series may be extracted from its ore by electrolysis. A metal below carbon in the reactivity series (zinc to silver) may be extracted from its ore by heating with carbon. The metal is displaced from its non-metal anion by the more reactive carbon. Carbon is used because it is readily available and cheap (coke or charcoal are both carbon). The metal in the ore is said to be reduced by reaction with carbon. Hydrogen may be used to reduce metals which are lower than itself in the reactivity series, but since it is more expensive than carbon it is



only used on a large scale for the extraction of tungsten, to avoid the formation of tungsten carbide.

Below is a general summary of the extraction methods and where they are applicable.

Method	Type of metal
Mining the pure metal	Noble metals such as Gold, Silver, and platinum
Roasting the sulphide, and reduction of the oxides	Mostly metals of the p-block elements
Reduction of the oxides	Less reactive especially in the p-block elements
Electrolysis of the molten solids	Reactive elements in Groups 1, 2, & 3, eg Na, Mg, and Al.

Alkali Metals

Despite being close family members, the elements of group 1 do not occur together largely because of their different ionic sizes. These elements are too reactive to be found free in nature. Sodium occurs mainly as NaCl (salt) in sea-water and dried-up sea beds. Potassium is more widely distributed in minerals such as sylvite, KCl, but is also extracted from sea-water. The alkali metals are so reactive they cannot be displaced by another element, so are isolated by electrolysis of their molten salts

Lithium is the thirty-fifth most abundant element in the earth crust (approximately 18 ppm). Sodium and potassium are in numbers seven (approximately 22 700 ppm) and eight (approximately 18 400 ppm) respectively. Rb and Cs are mainly obtained as by-products from Lithium processing. Francium is radioactive and does not occur appreciably in nature.

Extraction

The compounds of group 1 are very stable to heat and very high in electronegative series and they react with water, so both thermal decomposition and displacement by another element in solution are impractical. The metals are also very strong chemical reducing agents, meaning that they can not be replaced by reducing their oxides. Also, electrolysis of aqueous solutions in order to obtain the element may only be successful if mercury cathode is used.

Thus the metals are all isolated by electrolysis of their fused salts, mostly halides, and often with impurities added to lower the melting point.



The Electrolytic Production of Lithium

Electrolyte LiCl–KCl eutectic feed LiCl temperature 400–460°C. Anode carbon cathode Mild Steel

Overall Cell Reaction $\text{LiCl (l)} \rightarrow \text{Li (l)} + 1/2 \text{Cl}_2 \text{ (g)}$ for which $E_{427^\circ\text{C}} = 3.6 \text{ V}$

Anode $\text{Cl}^- \rightarrow 1/2 \text{Cl}_2 \text{ (g)} + \text{e}^-$

Cathode $\text{Li}^+ + \text{e}^- \rightarrow \text{Li (l)}$

Current Density 2 A/cm²

Energy Consumption 35 kWh/kg

Extraction of sodium

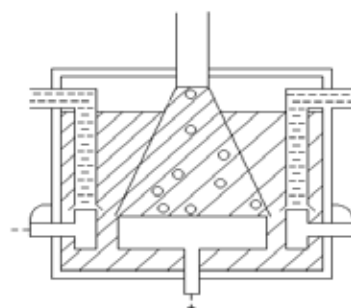
On industrial scale sodium metal is extracted by “Down’s Process”. Sodium is extracted by the electrolytic reduction of purified, molten rock salt. Down’s cell consists of a rectangular container of steel. Inside of the tank is lined with firebricks. Anode is a graphite rod which projects centrally up through the base of the cell. Cathode is a ring of iron, which surrounds the anode. The anode and cathode are separated from each other by a cylindrical steel gauze diaphragm so that Na and Cl₂ are kept apart. A bell like hood is submerged over the anode.

The main draw back is that the Melting point of NaCl is 801°C. At this temperature, molten NaCl and Na form a metallic fog in the container, which is impossible to separate.

In order to overcome this difficulty instead of only NaCl, a mixture of NaCl (42%) and CaCl₂ (58%) is electrolyzed in down’s cell. The melting point of this mixture is 600°C. At 600°C no metallic fog is formed.

When an electric current is passed through the molten mixture of NaCl and CaCl₂, NaCl decomposes into Na⁺ and Cl⁻ ions. Na⁺ ions migrate towards cathode while Cl⁻ ions towards the anode. The molten sodium collects in the cathode compartment where it rises to the top and is tapped off by a pipe. Chlorine is collected at the anode.

Label this diagram of a Downs electrolytic cell with: Carbon anode; Chlorine gas; Insulator; Molten electrolyte; Molten sodium; and, Steel cathode.





The chemistry.

Fused NaCl contains sodium and chloride ions.



Electrochemical changes

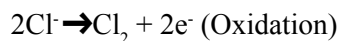
At cathode

Na^+ -ions migrate to cathode where they are reduced to Na.

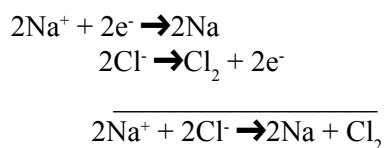


At anode

Cl^- -ions migrate to anode and oxidised to form chlorine gas.



Overall Reaction



After manufacture, sodium is stored under oil because it is rapidly oxidized by atmospheric oxygen. Furthermore, sodium reacts explosively with many other non-metallic elements and compounds.

A similar cell can be used for the extraction of Potassium from KCl, though the cell must be operated at higher temperatures. Modern method uses cheaper Sodium vapour to reduce the K at 850oC in large floatation tanks.





Question Explain why during electrolysis calcium is also obtained but it does not mix with Sodium?.

Answer: During electrolysis calcium is also obtained at cathode but sodium and calcium are separated from each other due to difference in density. Density of Na is 0.67gm/cc and the density of Ca is much higher than that of Na i.e. 2.54gm/cc. That's why they do not mix with each other.

Alkali earth metals

Occurrence and Extraction

These elements are all found in the Earth's crust, but not in the elemental form as they are so reactive. Instead, they are widely distributed in rock structures. The main minerals in which magnesium is found are carnellite, magnesite and dolomite. Calcium is found in chalk, limestone, gypsum and anhydrite. Magnesium is the eighth most abundant element in the Earth's crust, and calcium is the fifth.

Industrial Information

Magnesium is the only Group 2 element used on a large scale. It is used in flares, tracer bullets and incendiary bombs as it burns with a brilliant white light. It is also alloyed with aluminium to produce a low-density, strong material used in aircraft. Magnesium oxide has such a high melting point it is used to line furnaces.

Two methods of producing magnesium commercially are used. The principal method is the **Dow Sea water process**; electrolysis of fused magnesium chloride, which is used in the extraction of magnesium from seawater (the principal source) and from dolomite. In recovery from seawater, the magnesium is precipitated as magnesium hydroxide by treatment with lime (calcium oxide) obtained from oyster shells. The hydroxide is collected and treated with hydrochloric acid to form the chloride. The chloride is fused and electrolyzed, forming magnesium metal and chlorine gas. The molten metal is cast into ingots for further processing; the chlorine gas is made into hydrochloric acid and is reused to form magnesium chloride. About 1 lb of magnesium is recovered from each 100 gal of seawater; the oceans are a virtually inexhaustible source of this metal (0.13% Mg). A second method of magnesium production, called the ferrosilicon process, involves the reduction of magnesium oxide (prepared by calcining dolomite) with an iron-silicon alloy.

All the factors that were considered in choosing the extraction method for Group 1 metals are also relevant here. All the metals can be obtained by electrolysis of their fused chloride, with sodium chloride added to lower the melting point.

(Write the chemical equation for the reaction described above).



Group 13

Occurrence and Extraction

These elements are not found free in nature, but all are present in various minerals and ores. The most important aluminium-containing minerals are bauxite and cryolite.

Aluminium is the most widely used element in this Group. It is obtained by the electrolysis of aluminium oxide, which is purified from bauxite. The melting point of the aluminium oxide is too high for electrolysis of the melt, so instead it is dissolved in molten cryolite.

Industrial Information

Boron.

Discovered : by J.L. Gay-Lussac and L.J. Thenard in Paris, France, and Sir Humphry Davy in London, UK in 1808. The name is derived from the Arabic 'buraq', borax, its principal ore. Pure boron is a little-used dark powder, but boron compounds are important in many industries, such as glass and detergent manufacture and agriculture. Pyrex glass is tough and heat resistant because of the boric acid used to make it. Boron is an essential mineral for plants but not animals - in fact it can be toxic in excess. We take in about 2 milligrams each day from our food (about 60g in a lifetime).

Boron of low quality is obtained by reducing the B_2O_3 with Mg and Na at high temperatures. Pure Boron is not easy to obtain as it has very high temperatures of 2180°C.

Aluminium

Aluminium is obtained from its Bauxite ore; which may be $AlO.OH.(Al_2O_3.H_2O)$ or $Al(OH)_3.(Al_2O_3.3H_2O)$. The ore is first purified by removing the waste (usually Iron and silicon compounds) by dissolving the ore in NaOH. The amphoteric Al dissolves but not the impurities. Then the $Al(OH)_3$ is precipitated from the alkaline aluminate either by bubbling through the solution CO_2 gas or by seeding the solution with Al_2O_3 . The $Al(OH)_3$ is then calcined (heated strongly) to produce pure Al_2O_3 . Al_2O_3 is then melted with cryolite $Na_3[AlF_6]$ and electrolyzed in graphite lined steel tank, which also serves as the cathode. The anodes are all made of graphite. The cell runs continuously, and at intervals molten aluminium (m.p 600°C) is tapped from the bottom of the cell and more bauxite is added.



Quiz 1. What are the uses of the cryolite?

Solution 1 it improves the electrical conductivity of the cell as Al_2O_3 is a poor conductor and 2) it serves as added impurity and lowers the melting point of the mixture to about 950°C .

Quiz 2. Find the products formed at the anode terminal of the cell.

Solution: O_2 , CO_2 , and F_2 .

Quiz 3. what is the main draw back of this process?.

Solution 3. It is energy intensive and hence is only economical where cheap hydroelectricity power is available.

Group 14

The group 14 elements are well known except for germanium, carbon is the seventh, and silicon is the second most abundant element by weight in the earth crust. Ge, too, occurs as traces in mineral ores. Though the abundance of Tin and Lead are comparatively low, they occur in concentrated deposits and hence easy to extract.

Carbon occurs in deposits as Coal, Crude oil, and carbonate salts. It similarly occurs as graphite, diamond, soot, CO_2 , and even as CO gas.

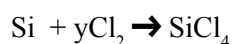
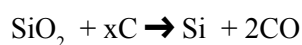
Coke is produced in large amounts by high temperature carbonization of coal. Here the coal is heated in large ovens in the absence of air. The Coke is extensively used in the metallurgical extraction of iron and many other metals.

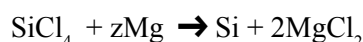
Quiz. How would diamond and graphite be extracted?.

Solution. For diamond and graphite, they are mined then just purified/cleaned as they occur 'native'-uncombined in most cases.

Silicon

The elemental Si is obtained by reducing the SiO_2 with excess high purity coke. Though high purity Si is obtained by converting the Si to SiCl_4 , purifying this by distillation, and then reducing the chloride with Mg or Zn.





N.B. Kindly balance the equations above by getting the numerical values (coefficients) x, y, and z. ($x = y = z = 2$).

Semiconductor quality Si can be made by sodium reduction of $\text{Na}_2[\text{SiF}_6]$.

Under exposure to oxygen, a silicon surface oxidizes to form silicon dioxide (SiO_2). Native silicon dioxide is a high-quality electrical insulator and can be used as a barrier material during impurity implants or diffusion, for electrical isolation of semiconductor devices, as a component in transistors, or as an interlayer dielectric in multilevel metallization structures such as multichip modules. The ability to form a native oxide was one of the primary processing considerations which led to silicon becoming the dominant semiconductor material used in integrated circuits today.

N.B. Kindly, get the extraction process for Lead (from the ore galena) and Tin (Cassiterite ore).

Group 15

Nitrogen comprises 78% of the earth's atmosphere, though not very abundant in the earth crust. This could be partly because nitrates are very soluble in water hence not 'stable' in the crust, though a few are found in desert deposits (NaNO_3) like in Chile. Phosphorus is the eleventh most abundant element in the earth crust. It also constitutes 60% of our human beings bones and teeth. Mostly it is found as fluoroapatite [$\text{Ca}_3(\text{PO}_4)_2 \cdot \text{CaF}_2$] and hydroxyapatite [$\text{Ca}_3(\text{PO}_4)_2 \cdot \text{Ca}(\text{OH})_2$]. The elements of As, Sb and Bi are not very abundant. Their most important source is the sulphides occurring as traces in mineral ores. Care should be taken as both As and Sb are poisonous.

Nitrogen

Condensing air to liquid state, and then fractionally distilling the liquid air commercially obtain the nitrogen gas. Since, nitrogen has a lower boiling point than O_2 , it distills off first.

Quiz 1. What are the major components of liquid air?.

Answer: N_2 , O_2 , Ar, H_2O , CO_2 , Ne, H_2 , He, etc.

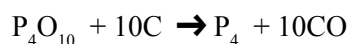
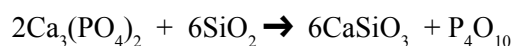
Quiz. 2. What are the boiling points for liquid N_2 , O_2 , CO_2 , and H_2O ?



Answer: N₂ (-196 °C) O₂ (-183 °C), H₂O (0 °C), CO₂ (-78 °C sublimes).

Phosphorus

Found most often in phosphate rock. Primary mining areas are Russia, USA, Morocco, Tunisia, Tongo and Nauru. Commercially obtained by reducing the calcium phosphate with C in an electric furnace at 1400 – 1500oC. Sand (silica SiO₂) is added to remove the calcium as liquiod slag (calcium silicate) and to drive off the phosphorus as P₄O₁₀. Carbon reduces the P₄O₁₀ to P. At higher temperatures, gaseous P distills off, mainly as P₄ but with some P₂. This is then condensed to white phosphorus by passing it through water.





Group 16

Oxygen is the most abundant of all the elements. In free form, it forms 21 % by composition but 23 % by weight of the earth's atmosphere. However, oxygen makes up 47 % of the earth crust where its major component is the silicate mineral, oxides, sulphates and carbonates. It makes 89 % by weight of the world ocean waters.

Sulfur is the sixteenth most abundant element and it constitutes 0.03% by weight of the earth crust. It occurs mainly as sulphide and sulphate (gypsum) ores. The remainder elements of this group (Se, Te, and Po) occur in trace amounts.

Oxygen

Oxygen is commercial produced at industrial scale by frictional distillation of liquid air (see separation of Nitrogen from liquefied air). The gas produced by this method contains traces of N_2 , Ar and other gases. It is sometimes prepared at laboratory scale by heating the $KClO_3$ mixed with MnO_2 as a catalyst. Heating $NaClO_3$ at $150^\circ C$ with the same MnO_2 catalyst otherwise produces small amounts for emergency breathing in aircrafts.

Sulphur

Major sources of S are; 1) recovery from natural gas and petroleum (50%), 2) mined by Frasch Process (19%), 3) from pyrites (17%), 4) recovery from smelter gases (12%), mined as sulphur ore (4%).

N.B. It is morally and environmentally essential that all traces of S are removed from the petroleum product for the, H_2S gas, has objectionable smell and is environmentally undesirable.

The underground deposits of S are mined by **Frasch process**, which produces sulphur of very high purity (99%).

Here sulfur is obtained from underground deposits by holes, which are drilled down through the overlying rock into the sulfur deposits. A special series of concentric pipes are then inserted into the drill hole. The outer pipes will contain superheated steam (usually about $160^\circ C$), which is pumped down into the deposit. Since the melting point of sulfur is so low ($115.21^\circ C$, just a little over the boiling point of water), it readily liquefies. As the sulfur becomes molten, it is removed by pumping air down the central pipe. When the molten sulfur reaches the surface, it is pumped onto wooden blocks where the sulfur again solidifies.

Exercise. Get the diagram of the Frasch process, draw and label its parts.



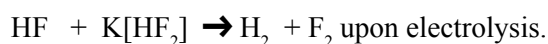
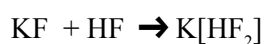
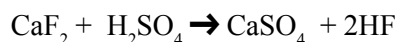
Group 17

The halogens are very reactive and therefore do not occur in free state. However, all except Astatine (radioactive and hence short lifetime) are found in combined form in the earth's crust. Fluorine is the thirteenth most abundant element by weight in the earth crust and chlorine is the twentieth. Bromine and iodine are comparatively rare. The main source of Fluorine is Fluorspar (CaF_2) or fluorite and fluoroapatite. The name fluorspar is because the mineral fluoresces, that is it emits light when heated. Interestingly, seawater contains negligible amounts of Fluorine; this is because the sea contains large quantities of Ca^{2+} contributing to the insoluble CaF_2 . The most abundant compound of Chlorine is NaCl , and it is used to produce virtually all the Cl_2 and HCl .

Bromine also occurs in seawater. Though Iodine is present in negligible amounts, though is normally concentrated by the seaweeds.

Fluorine

Fluorine is very reactive, and this causes great difficulty in the preparation and handling of the element. It is obtained by treating CaF_2 with concentrated H_2SO_4 to give an aqueous mixture of HF . This is then distilled, yielding the anhydrous HF liquid. Then a cooled solution of KHF_2 in anhydrous HF is electrolyzed, giving F_2 and H_2 .



But there are several difficulties in obtaining the fluorine;

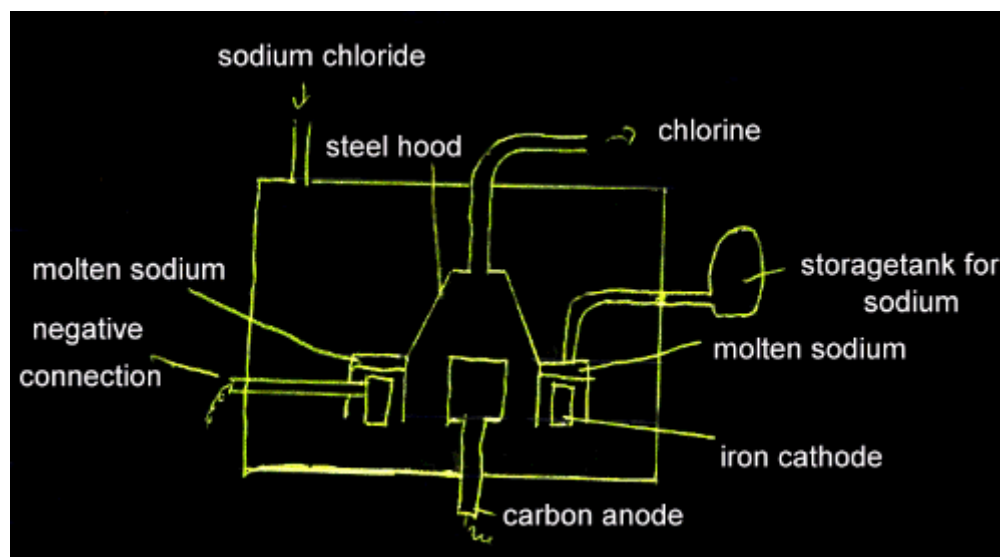
HF is toxic and corrosive both to glass and tissues.

Fluorine is very reactive; it attacks glasses and catches fire with traces of gases and crystalline silicon.

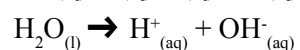
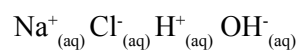
Chlorine

The name is derived from the Greek 'chloros', meaning pale green. A yellowy-green dense gas with a choking smell. It is very poisonous and was used as a weapon during the First World War. The gas is made on a large scale from electrolysis of either aqueous or anhydrous NaCl salt (sodium chloride) during the production of NaOH and Na metal respectively.

The diagram below shows for aqueous Sodium Chloride.



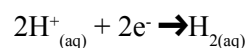
Particles Present in the electrolyte



The Hydrogen ions and Hydroxide ions are present because at any given moment, 1 in 10 million water molecules will split up into ions. Although the proportion of Hydrogen ions is low, the number of Hydrogen and Hydroxide ions is significant.

At the cathode (-)

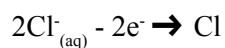
Both Na^+ and H^+ are attracted, but Sodium is above Hydrogen in the reactivity series, and so it is easier to give electrons to H^+ than Na^+ . Thus, Na^+ are not discharged.



Since Hydrogen ions are constantly being removed from the water, the OH^- concentration around the cathode is constantly rising, i.e. the solution is becoming more alkaline, the pH paper turns blue purple. Sodium ions are also present in the solution around the cathode and so the important strong alkali Sodium Hydroxide can be extracted from the solution there.



At the anode (+)



Chloride ions are attracted and discharged.

N.B. Chlorine reacts with Sodium Hydroxide and so there is a need to keep apart the Sodium Hydroxide produced at the anode. One of the ways that this is achieved in industries is that the electrolyte is kept higher than the anode side so that diffusion of Hydroxide ions through a partition between the electrodes (Diaphragm) is minimised. The cell used in industries is called a «Diaphragm Cell.»

Otherwise for the rest of the remaining elements, bromine is obtained by processing brine from wells and from seawater, iodine is obtained from brine wells and sea plants such as kelp. But astatine is radioactive. Less than 30 grams exist in the Earth's crust. For greater quantities, this element needs to be synthesized in a laboratory.



Group 18

Under normal conditions, the elements of group 18 have little if no tendency to gain or lose electrons. Thus they have little tendency to form bonds, and so they mostly exist as single atoms.

Exercise. Get as many compounds of Xenon as you can.

Solution: Xe[PtF₆], XeF, XeF₄, XeO₃, XeO₂F₂, etc.

The gases He, Ne, Ar, Kr, and Xe all occur in the atmosphere.

Ar is abundant (0.93% in atmosphere) and can be recovered by fractional distillation of air. See details on how N₂ and O₂ were obtained. It originates in the air mostly from electron capture (β^+ decay) of potassium.

For the other gases, helium is extracted from natural gas wells, argon are present in about 1% of the Earth's atmosphere, other noble gases are present in trace amounts, and radon is a radioactive gas produced from the radioactive decay of the element radium.

Problem set Unit 4 (accounts for 8 %).

1. In order to obtain metals from their sulfide ores, the ores are usually converted into, a) sulfates, b) oxides, c) sulfites, d) carbonates
2. Which of the following occurs during the electrolysis of fused CaCl₂?, a) calcium ions are reduces, b) calcium atoms are reduced, c) Chloride ions are reduced, d) Chloride atoms are reduced.
3. Which element is most likely to be found uncombined in nature?, a) Na, b) O, c) N, d) Ar.
4. Aluminium is majorly obtained form which ore, a) bauxite, b) oxide c) appatite, d) Alumina
5. The most abundant element in the earth crust is more likly to be, a) aluminium, b) oxygen, c) water, d) silica
6. Nitrogen is more abundant in the atmosphere than in the earth crust mainly because, a) it is a light gas, b) it is inert, c) its salts are very soluble in water, d) more stable in the presence of oxygen in the atmosphere.
7. Which of the following is not true about phosphorus?, a) It is solid at room temperature, b) white phosphorus is more toxic than the rest, c) needs to be stored under water, d) black phosphorus is the most thermodynamically stable.
8. Water (H₂O) and hydrogen sulphide (H₂S), share members of chalcogen family (O and S). Whereas water is a liquid at room temperature, H₂O is a gas because?,



a) water is denser than H_2S , b) H_2S is smelly so it has to be a gas, c) S as opposed to O has a larger atomic radius, d) hydrogen bondings are stronger in H_2O than in H_2S .

Answer Key for Unit 4.

1) b, 2) a, 3) d, 4) a, 5) b, 6) c, 7) c, 8) d.



Learning activity # 5

Title of Learning Activity : **ECONOMIC USES AND APPLICATIONS OF s-, AND p-BLOCK ELEMENTS.**

At the end of this Unit, the learner should be able to;

1. State, appreciate and predict economic uses and applications of some s-block elements.
2. State, appreciate and predict economic uses and applications of some p-block elements.
3. Predict some economic uses of a given hypothetical s- and p-block elements.

Summary of the learning activity

This is the last unit of this module, so far we have appreciated most of the chemistry of the s and p block elements. However, the ultimate aim of any technology or knowledge is its application. In this unit we turn to discuss the various uses of the s and p elements. In furtherance of the above knowledge, we will discuss the possible uses of some hypothetical elements and their compounds. All the above will be crowned with some relevant examples and exercises.

List of Required readings

1. Philip Mathews, *Advanced Chemistry 1, physical and industrial*. Cambridge University Press, (1992).
2. Alan G. Sharpe; *Inorganic Chemistry, 3rd Edition*. Longman Singapore Publisher (1992).
3. Catherine E. Housecroft and Alan G. Sharpe; *Inorganic Chemistry*. Prentice-Hall International, USA. (2000).
4. F. Albert Cotton and Geoffrey Wilkinson, F. R. S. *Advanced Inorganic Chemistry, A Comprehensive text, 6th edition*. Wiley Eastern Limited, New Delhi. India. (1984).
5. J. D. Lee, *Concise Inorganic Chemistry, 4th edition*. Chapman & Hall, New York. USA. (1993).
6. Thomas R. Gilbert, Rein V. Kirss, and Geoffrey Davies; *Chemistry, The science in context*. W.W. Norton and company NY, USA.(2004).
7. Carole H. McQuarrie, Donald A. McQuarrie and Peter A. Rock. *General Chemistry, 3rd edition*. W. H. Freeman and Company, New York, USA, (1990).
8. Otto S. Wolfbeis, *Fiber Optics Chemical sensors and Biosensors, Vol. 1*. CRC Press, London, (1991).



List of relevant useful links

1. <http://www.webelements.com/webelements/scholar/elements/carbon/uses.html> Uses of Carbon
2. <http://www.corrosionsource.com/handbook/periodic/14.htm> Uses of silicon
3. <http://www.webelements.com/webelements/scholar/elements/nitrogen/uses.html> Nitrogen
4. <http://chandrasekaran.tripod.com/uses.html> Phosphorus
5. <http://www.infoplease.com/ce6/sci/A0860185.html> Oxygen
6. <http://library.thinkquest.org/C004050F/use.htm> Uses of chlorine

List of relevant MULTIMEDIA resources

- Computer with internet connecting facility to access relevant links and free source resources.
- Multi-media resources such as CD players, VCD etc.
- CD-ROM for this module for compulsory reading and demonstrations.

Learning activities

USES OF THE s-, AND p-BLOCK ELEMENTS

Basically, the use or application of any elements is determined by the chemistry and the physical properties of the particular element. In most cases, we will discuss only one or two group members as representatives of that group.

Alkali metals

This is the Group of Li, Na, K, Rb, and Cs and they are all metals

Lithium: Used for making alloys especially with aluminium to make aircraft parts, which are light and strong. Due to its size and electropositivity (opposite of electronegativity), lithium is used in both primary and secondary lithium batteries. Otherwise its compound, Li_2CO_3 , is used to toughen glass.

Sodium: Globally, about 60% of Na is used to form alloys with lead, which is used to produce PbEt_4 and PbMe_2 used as anti-knock additive in Petrol. Liquid sodium is used as nuclear power plant coolant. Caustic soda, NaOH, and soda ash, Na_2CO_3 are the most important alkali used in industries. Both find applications in paper making, alumina, soap, and rayon. NaOCl is used as bleach and disinfectants.

Quiz. Why would it be possible to use the liquid Na, as a new clear power plant coolant?. (Hint used the operational temperature of a plant at 600°C).

Solution: Sodium's use as a coolant in nuclear reactors depends on three of the metal's properties. First, it is a liquid at temperatures between 98°C and 883°C, so the metal flows easily around the hot reactor core (which operates at a temperature



of about 660°C). Second, it is a good thermal conductor (because of its free moving delocalized electrons), so excess heat generated in the core is conducted away efficiently. And third, it has a relatively high thermal capacity, so large amounts of heat energy are absorbed for any given temperature rise.

Potassium

Potassium is one essential element. Almost 95% of the world potassium is used as fertilizer. KOH is used in soap and detergent making, KNO_3 finds use in explosives, while KBr is used in photography.

Alkali earth metals

In this group are Be, Mg, Ca, Sr, Ba, and Ra. Magnesium is the only Group 2 element used on a large scale. It is used in flares, tracer bullets and incendiary bombs as it burns with a brilliant white light. It is also alloyed with aluminium to produce a low-density, strong material used in aircraft. Magnesium oxide has such a high melting point it is used to line furnaces. In addition, magnesium metal is used in flash powder and photographic light bulbs. It is worthy to note that hard water contains calcium and magnesium ions. Chemically it is used as the famous Grignard reagent; $\text{C}_2\text{H}_5\text{MgBr}$.

Calcium ores are found as calcium carbonate, CaCO_3 in chalk, limestone, marble, and calcite whereas beryllium and its compounds are toxic, however, barium compounds are used as white paint pigments

Otherwise, Be has very low neutron capture cross sectional area and it finds use in nuclear power industry and as windows for X-ray tubes.

Quiz. What is milk of magnesia and what are its applications?.

Group 13

This is the group of B, Al, Ga, In, Tl. These elements are not found free in nature, but are all present in various minerals and ores.

Boron has limited uses, but is used in flares to provide a highly visible green colour. It is an essential component of nuclear reactor rods. Boron filaments are now used extensively in the aerospace industry as a lightweight yet strong material. Boric acid is used as a mild antiseptic, and borax as a water softener in washing powders. Borosilicate glass contains boron compounds. Pyrex glass is tough and heat resistant because of the boric acid used to make it. Boron is an essential mineral for plants but not animals - in fact it can be toxic in excess. We take in about 2 milligrams each day from our food (about 60g in a lifetime).

Aluminium is the most widely used element in this Group and most industrially important. It is light, non-toxic, has a high thermal conductivity, can be easily worked and does not corrode due to its oxide coating, which is very effective although only 10nm thick. It has several domestic uses such as cooking utensils, aluminium



foil and bottle tops, and is widely used in the building industry where a strong, light, easily-constructed material is required. These properties also make it invaluable in the building of aeroplanes and spacecraft. Above all it is used as an anti-acid for indigestion. $\text{Al}_2(\text{SO}_4)_3$ is used in water treatments.

N.B. Al may not be as harmless as previously thought, it is toxic to people with kidney failure, and those with Alzheimer's (which causes senility) disease have deposits of Al in the brain.

Ga, also finds important use in semiconductor and transistors industry and to make memory chips for computers.

Group 14

Group 14 comprises C, Si, Ge, Sn, and Pb. All of the elements are somewhat useful and have wide applications. However, this discussion will be restricted to Carbon and Silicon.

Carbon (C), has many industrial uses for its three forms; amorphous, diamond, and graphite. The amorphous form (coal) is used in power plants, and is used in the iron industry to make steel. Coal is used as fuel. The diamond form is used as jewelry, and is used as part of cutting tools. The graphite form is a great conductor, a lubricant and is used for pencils and sports equipment. A new form of carbon, buckminsterfullerene (C_{60}) has been made. It forms with treatments of graphite by lasers and is now commercially available in small quantities. Carbon dioxide is one that human's exhale after breathing oxygen. Don't forget that Carbon dioxide is used to produce carbonated soft drinks and soda water. C-14 isotope used in archaeological dating (telling how old something is by determining the amount of Carbon-14 present in the item being tested as compared to a standard value for a similar object which is new).

Otherwise, because of the many combinations, carbon is capable of forming chains and the compounds are important in many facets of the chemical industry.

Quiz. What property of Carbon makes it be used as a reducing agent in extraction of metals?.

Solution: In electrochemical series, carbon is a reducing agent to these metals.

Quiz. Why is graphite used as a conductor electricity yet diamond is not ?.

Solution: In the structure of graphite, there is a loose electron for every carbon atom that conducts electricity. Diamond uses all its four electrons for bonding.

N.B. By the end of 2007, the global car industry is expected to decide on the next-generation refrigerant in car air conditioning. CO_2 is one discussed option

**Silicon:**

Silicon is one of man's most useful elements. In the form of sand and clay it is used to make concrete and brick; it is a useful refractory material for high-temperature work, and in the form of silicates it is used in making enamels, pottery, etc. Silica, as sand, is a principal ingredient of glass, one of the most inexpensive of materials with excellent mechanical, optical, thermal, and electrical properties.

Hyperpure silicon can be doped with boron, gallium, phosphorus, or arsenic to produce silicon for use in transistors, solar cells, rectifiers, and other solid-state devices which are used extensively in the electronics and space-age industries. Hydrogenated amorphous silicon has shown promise in producing economical cells for converting solar energy into electricity.

Silicon is important to plant and animal life. Diatoms in both fresh and salt water extract Silica from the water to build their cell walls. Silica is present in the ashes of plants and in the human skeleton. Silicon is an important ingredient in steel; silicon carbide is one of the most important abrasives and has been used in lasers to produce coherent light of 4560 Å.

In addition, silicon finds use in medicine, silicone implants (Breast implant).

Quiz. What property of silicon would be exploited in making medicinal implants?

Solution: Silicon network is generally stable and therefore does not react with human tissues.

Exercise: Take a walk around your neighborhood and note down all the applications of Tin and Lead.

Group 15

Group 15 comprises N, P, As, Sb (Antimony), Bi. This is one group where the metallic property emerges as one moves down the group. The discussions here will be restricted to Nitrogen and Phosphorus.

Nitrogen.

Nitrogen has many uses, most of the air we breathe is nitrogen. Everything needs nitrogen, plants breathe mostly nitrogen with partly carbon dioxide so they would be gone without nitrogen. One of these uses is making light bulbs. Light bulbs are almost always filled with Nitrogen. Used in making of ammonia, NH_3 , production (Haber process). This is the greatest use of nitrogen. Ammonia is used for fertilizer production and to produce nitric acid (Ostwald process). Other uses are:

1. Used by the electronics industry, which uses the gas as a blanketing medium during production of such components as transistors, diodes, etc.
2. Used in annealing stainless steel and other steel mill products
3. Used as a refrigerant both for the immersion freezing of food products and for transportation of foods



4. Liquid nitrogen is used by the oil industry to build up pressure in wells to force crude oil upward
5. Used as an inert atmosphere in tanks of explosive liquid storage tanks, both in ground-based tanks and in ships
6. Another use is explosives. Nitrogen can be used to make four different explosives. You probably know Trinitrotoluene better by the initials TNT. Yet another use for Nitrogen is that it is used to make Laughing Gas.

Phosphorus

In living beings, phosphorus is found in bones, teeth, nucleic acids (DNA and RNA), in the energy carriers (such as ATP), lipids, proteins and enzymes. This shows the importance of phosphorus to the health of body in general and bones and teeth in particular.

Additional uses are:

1. Used in the manufacture of safety matches, pyrotechnics, incendiary shells, smoke bombs, tracer bullets, etc. Since phosphorus catches fire readily, it is used in all matches; earlier white phosphorus was used and in today's "safety matches" red phosphorus is used. The red phosphorus is safe because it is less volatile and less poisonous (0.1 g of white phosphorus can kill a person!) and eliminates the accidental fire as it catches fire much less readily.
2. Fire stoppers; One interesting irony is that phosphorus that is used to make fires (in its elemental form) is used to stop or reduce the hazards of fire (in the form of chemicals derived from it). Phosphorus compounds are among the best flame-retardants for many flammable materials; many are used as plasticizers and flame-retardants as two-in-one.
3. Fertilizers
4. Water treatments; Many soluble phosphates are used to remove unwanted metal salts from the water (for many uses); to convert hard water to soft water by removing the metal ions as insoluble metal phosphates. For much better performance some organophosphates are used, even though they are much costlier than the normal inorganic phosphates
5. Phosphates are used in the production of special glasses, such as those used for sodium lamps.
6. Bone-ash, calcium phosphate, is used to produce fine chinaware and to produce monocalcium phosphate used in baking powder.
7. Important in the production of steels, phosphor bronze, and many other products.
8. Na_3PO_4 is important as a cleaning agent, as a water softener, and for preventing boiler scale and corrosion of pipes and boiler tubes
9. Pesticides.
10. Daily use; Phosphorus compounds can be found in tooth pastes (cavity protection or tartar control), in shampoos (Head&Shoulder) and many more commonly used items



Group 16

This group is made of (oxygen, sulfur, selenium, and tellurium). We will concentrate on the uses of Oxygen, and sulfur. On your own, kindly get to know the uses of the other remaining three elements.

Oxygen.

Oxygen was formerly the official standard for the atomic weights of elements. The chemists used natural oxygen, a mixture of three isotopes, to which the value of 16.

Quiz. Which element then replaced Oxygen in this standardization?.

Answer; In 1961 carbon-12 replaced oxygen as the standard.

The most important use of oxygen is for breathing. Animals require oxygen in their air (or water) to live. Another use for oxygen is for oxidation reactions -- the most obvious example is burning. Fires require oxygen to burn. Some fire extinguishers work by preventing oxygen from reaching the burning materials.

Other notable uses are;

1. Oxygen is utilized in medicine in the treatment of respiratory diseases and is mixed with other gases for respiration in submarines, high-flying aircraft, and spacecraft.
2. Oxy-acetylene welding
3. Frequently used to aid respiration of patients in hospitals
4. Used in making methanol and ethene oxide
5. Rocket fuel oxidant
6. Steel manufacture
7. Ozone (O_3) in the atmosphere is protection against the sun's ultraviolet rays

Sulphur

Most of the sulphur produced is used to produce sulphuric acid, H_2SO_4 , the most important manufactured chemical in the world. Sulphuric acid has many uses, including for the synthesis of fertilisers and polyamides. It is used in batteries («battery acid»).

Additionally, Sulphur finds use in;

1. In the manufacture of sulphur dioxide and sodium sulphite which are used for bleaching straw and wood fibers, and for removing lignin from wood pulp for the paper industry,
2. In the manufacture of organic chemicals, which contain sulphur.
3. Sulphur is a component of black gunpowder (a mixture of potassium nitrate, KNO_3 , carbon, and sulphur). It is used in the vulcanisation of natural rubber, as a fungicide, and as a fumigant.



4. Sulphur compounds are used in the bleaching of dried fruits and for paper products.

N.B. Selenium and tellurium are often associated with each other. They tend to occur together in the Earth and have somewhat similar properties. They have many uses in common. In recent years, some important new uses have been found for selenium. It is now used in the manufacture of plain paper photocopiers.

Group 17

Comprises F, Cl, Br, I, and At. Chemically, the halogens are too reactive to occur free in nature. The halogens are probably the most important Group of the Periodic Table used in industry. However this discussion is mainly be restricted to F and Cl.

Flourine

Named and recognized by Andre Marie Ampere of France in 1811 and isolated by Henri Moissan of France in 1886. Despite blowing a lot of things up, Flourine is still useful. Fluorine is an essential ingredient in Teflon and Freon (a refrigerant and propellant). Freon gases are the ingredients in aerosol containers, which may be damaging the ozone layer in the atmosphere by changing ozone (O_3) to oxygen (O_2). Soluble fluorides are very lethal. Small quantities taken over a period of time produce fluorosis: hypercalcification of bones, permanent stiffness of the spinal column, and stiff and painful joints. In lesser amounts, fluorine in water can cause a mottling or irregular browning of people's teeth. But at one part per million in drinking water, fluorine causes a substantial reduction in children's cavities.

In addition, Fluorine and its compounds are used in producing isotopically fractionated uranium (from UF_6), commercial fluorochemicals, including many well known high-temperature plastics, and hydrofluoric acid is extensively used for etching the glass of light bulbs, etc.

Chlorine

Chlorine is widely used in making many everyday products. It is used for producing safe drinking water the world over. Even the smallest water supplies are now usually chlorinated.

It is also extensively used in the production of paper products, dyestuffs, textiles, petroleum products, medicines, antiseptics, insecticides (you remember DDT?), food, solvents, paints, plastics, and many other consumer products.

Most of the chlorine produced is used in the manufacture of chlorinated compounds for sanitation, pulp bleaching, disinfectants, and textile processing. Further use is in the manufacture of chlorates, chloroform, carbon tetrachloride, and in the extraction of bromine.

Organic chemistry demands much from chlorine, both as an oxidizing agent and in substitution, since it often brings many desired properties in an organic compound when substituted for hydrogen, as in one form of synthetic rubber.



As for the remaining members of the group, bromine is used in photographic chemicals, dyes, fire retardants and in pharmaceutical products. Iodine compounds are used in producing iodized salt, photographic chemicals and in medicines and is required in human body in order to make the hormone thyroxin

Group 18

All the gases He, Ne, Ar, Kr, and Xe are relatively uncreative and occur in the atmosphere. Here we discuss, He, Ar, and Xe.

Helium;

1. Helium is used to fill balloons, blimps, pressurize liquid fuel in rockets.
2. Helium and argon are used in arc welding and metallurgical processes to protect materials from reacting with nitrogen and oxygen in the air

Argon

Argon finds application in:

1. Used in electric light bulbs and in fluorescent tubes at a pressure of about 3 mm, photo tubes, glow tubes, etc.,
2. Used as an inert gas shield for arc welding and cutting, blanket for the production of titanium and other reactive elements.
3. Protective atmosphere for growing silicon and germanium crystals, lasers, light bulbs electric light bulbs and fluorescent tubes, photo tubes, glow tubes,
4. welding, cutting, as a protective gas for other substances,
5. Argon fills incandescent light bulbs. This gas replaces the oxygen-rich air that corrodes the Tungsten filament and causes blackening of the bulb.
6. Argon provides the atmosphere for growing crystals to be used in the semiconductor industry, such as Silicon and Germanium.
7. Argon is used in dating rocks with a process called Potassium-Argon dating. By comparing the amount of K-40 to Ar-40, the age of rocks can be found. The larger the ratio of Argon to Potassium, the older the samples.
8. Argon is also used in radio tubes and Geiger counters (the probe that senses radiation is made of Argon), as well as an inert atmosphere for arc-welding metals like aluminum and stainless steel.

Xenon.

1. used in making electron tubes, stroboscopic lamps, bactericidal lamps, and lamps used to excite ruby lasers for generating coherent light.
2. used in the atomic energy field in bubble chambers, probes, and other applications where its high molecular weight is of value.
3. potentially useful as a gas for ion engines.
4. the perxenates are used in analytical chemistry as oxidizing agents



5. Xenon can also be used in flash cameras and “high-intensity” lamps for projecting movies.
6. It has most recently been used to replace the halogen in the headlights. Xenon was found to be more efficient and brighter.

PROJECT

Since most of the elements/minerals in our neighbourhood occurs as earth/soil deposits, in not more than 1000 words, write how {i.e. medicinal, cosmetics, leisure, ceremonies etc} your country/community uses the soils/mineral deposits. Kindly give a brief description of the physical appearances of the mineral or soil. It be presented as a joint write up for people from one community or country but submitted separately [This will account for 5% of your grade in this unit]

Problem set Unit 5 (accounts for 10 %).

1. which group is for elements that are best extracted by electrolysis, which one is not ?
a) Na, Mg, Al, b) Pb, Sn, C, c) Li, Mg, K, d) Ca, K, Cs
2. For Down’s method to be used for extraction of K
a) current must be increased, b) operational temperatures must be increased,
c) the walls must be made more inert, d) no modification is required.
3. In Downs’ method, CaCl_2 is added to so that
a) the boiling point of CaCl_2 is lowered, b) the boiling point of NaCl is lowered,
c) the fogging is eliminated in the chambers, d) less heat may be used.
4. Which substance is used as a moderator in a nuclear reactor?
a) Al, b) C {graphite}, c) helium, d) Plutonium.
5. Nitrogen containing minerals are not so much in the earth crust because
a) Nitrogen is a light gas, b) most of its salts are not stable,
c) most of its salts are soluble in water, d) It is only in the atmosphere.
6. 5 points from the project: introduction, chemistry involved, physical characterization of the minerals, relevance and conclusion (1 mark for each)

Answer Key for Unit 5

- 1) b, 2) b, 3) c, 4) b, 5) c. 6 (see above).



XI. Compiled List of all Key Concepts (Glossary)

ELECTRONIC CONFIGURATION. A summary of the distribution of electrons about a nucleus, specifying which orbitals these electrons are most likely to be found.

A MINERAL. is a natural, inorganic crystalline material, which is formed by geological processes and contains the substance of interest.

AN ORE. A mineral which contains a high enough percentage of a metal for economic extraction is called a metal ore.

VALENCE ELECTRONS. **electrons** in an outer shell of an atom that can participate in forming chemical bonds. They are easily shared or transferred during a chemical reaction.

ELECTRONEGATIVITY. is the power of an atom when in a molecule to attract electrons to itself. The Pauling scale is the most commonly used. Fluorine (the most electronegative element) is assigned a value of 4.0.

POLAR BONDS. A bond in which electrons are shared between elements having a difference in electronegativity of between 0.5 and ~1.5. The bond between hydrogen and oxygen in water is an example of a polar covalent bond.

EXTRACTION OF ELEMENTS. The process of obtaining the element(s) from a mixture or compound by chemical or physical or mechanical means

METALS. Simple definition is an element that loses electrons in a chemical reaction, An example of an alkali metal is sodium.

WATER OF HYDRATION. Water chemically combined with a substance in such a way that it can be removed, as by heating, without substantially changing the chemical composition of the substance.

INERT PAIR EFFECT. The inert pair effect explains why common ions of say Pb are Pb^{4+} and Pb^{2+} , and not just Pb^{4+} as we might expect from the octet rule. Valence electrons in an s orbital penetrate to the nucleus better than electrons in p orbitals, and as a result they're more tightly bound to the nucleus and less able to participate in bond formation. A pair of such electrons is called an «inert pair».

PHYSICAL PROPERTIES. Any characteristics of a material that can be observed without changing the identity of the material itself for example, density, electrical conductivity, melting points, etc.



CHEMICAL PROPERTIES. of matter describes its «potential» to undergo some chemical change or reaction by virtue of its composition. Chemical change results in one or more substances of entirely different composition from the original substances. For example hydrogen has the potential to ignite and explode given the right conditions. This is a chemical property.

CATHODE. is the terminal where current flows out. To avoid misconceptions, remember that the anode/cathode distinction is based on current, not voltage. Anode/cathode is not the same as positive/negative or vice versa. Illustrative example: for a battery being discharged, the positive terminal is the cathode, while for the same battery being recharged, the positive terminal is the anode

CATENATION. The ability of an element such as carbon (C—C), silicon (Si—Si), tin (Sn—Sn), etc. to form chains. This is because these bonds are very strong due to the use of some π bonding.

COVALENT BOND(S). An interatomic attraction resulting from the sharing of electrons between the atoms.

CONTACT PROCESS. A reaction sequence used for the commercial preparation of sulfuric acid.

DIAMAGNETISM. The physical property of being repelled by a magnetic field.

EXCITED STATE. The state of one atom in which at least one electron is not in the lowest possible energy level.



XII. Compiled List of Compulsory Readings

Unit 1. Properties Of Periodic Table Elements

Complete reference

PDF file: Name; Reading 1.

Abstract : The file introduces the learner to the basics of inorganic chemistry where electronic configuration and the orbital designations such as s, p, d, and f and their relative energies are the central theme. The rules that govern the electronic distribution are also introduced. Further it briefly outlines how to use Excel computer program to graphically monitor trends in both physical and chemical data values of the periodic table elements.

Rationale : Since the students come from different academic backgrounds, the site seeks to bridge that gap that may exist among the students by graphically presenting, from first principles, the electronic configuration of elements. The symbols of elements and their names are also presented so that the learner does not agonize on the identity of the elements. To fully appreciate the meaning of data associated with physical and chemical changes for the periodic table elements, Excel program is handy and graphical presentation of these changes is therefore paramount.

Unit 2. Atomic Structures and trends in S and P Block elements

Complete reference :

PDF file, Name: Reading 2.

Abstract : The file introduces how to measure the atomic and ionic sizes and the units of such measurements. The inherent differences between cationic and anionic sizes compared to their corresponding atomic sizes are also discussed. Also included are definition of the melting and boiling points.

Rationale: This site makes it possible to, in depth, appreciate the different methods of determining the atomic and ionic radii of elements. Usually, the minuteness of the atomic scale used in measurements of atomic and ionic radii is incomprehensible to beginners. This document vividly connects those atomic units to the common SI – units we are used to. Lastly, for the intension of showing, as in other fields of science, that chemistry retains common definition of some terms, this unit defines melting and boiling points of elements and briefly discusses how to determine them.



Unit 3. General Properties of S and P Block Elements

Complete reference:

PDF file, Name: Reading 3.

Abstract: This reading begins with the definitions and examples for Oxidation states of various elements in as a affected by the position of the elements in the periodic table. Highlighted next is the details of how to name the common elements and their chemical compounds. Here, both common and IUPAC naming criteria for the compounds are illustrated. At the end, the meaning and examples of both physical and chemical changes a substance can be subjected to are included.

Rationale: Since unit 3 deals with General Properties of S and P Block Elements, to understand the contents of the unit, the learner needs to have the knowledge of Oxidation states of any elements as a function of the elements position in the periodic table. Naturally, any learner would be interested in knowing how to name the chemical compounds they encounter in the course of this module. Such are the details included in the document. Lastly, defining with examples both physical and chemical changes would make Unit 3 more understandable to the learner.

UNIT 4 & 5 Occurrence, abundance, Extraction and Economic uses of s and p block elements.

PDF file, Name: Reading 4.

Abstract: The document introduces the meaning of the terms ore and minerals and other related terminologies. The next bit deals with the term extraction and all the considerations that are important for commercial exploitation of any mineral ore. Lastly, an exhaustive name list of major ores and minerals is attached.

Rationale: The section takes the learner on a tour of the terminologies associated with ores and mineralogy and extraction as a process. Since most of these terminologies are derived from fields other than chemistry, their mastery is usually a challenge to majority of young chemists. In addition, for good appreciation of details on extraction of minerals, the issues limiting such exploitation must be known to the learner. Included here too is a comprehensive name list for ores and minerals of common elements. The document is aimed at making mastery of both Units 4 and 5 easier to the learner.



XIII. Compiled List of Useful Links

Useful Link #1

Title : PERIODIC TABLE OF ELEMENTS

URL :http://en.wikipedia.org/wiki/Periodic_table

Screen capture :



Description: Wikipedia is a multilingual, web-based, free content encyclopedia project. Wikipedia's articles provide links to guide the user to related pages with additional information. Periodic as the most used root to understanding the chemistry inorganic substances is well presented here.

Rationale: The students will find this site useful, for one, it is frequently updates and besides is free. It contains relevant examples, questions and discussions that are strongly required for successful completion of this course.



Useful Link #2

Title : ORGANIZATION OF THE PERIODIC TABLE.

URL : <http://www.webelements.com/>

Screen capture :

Description: This site, which was developed for higher level chemistry students explains the periodicity of the elements in the periodic table. The site enables one to browse for different properties (both physical and chemical) of any one selected element. It underscores the Aufbau principles, shows where s, p, and d block s are in the periodic table and includes properties associated with elements of a given group.

Rationale: This site is core to the module. We understand things better when they are presented graphically and in colours, and that is what is in this site. The organization of the periodic table is well illustrated and vividly brought to life in relevant colours. In addition, the site dwells on the such properties of elements as electronegativity, atomic radii, etc, that comprise the major part of this module.



Useful Link #3

Title : PROPERTIES OF ELEMENTS

URL : <http://www.dayah.com/periodic/>

Screen capture :

Description: With an opportunity to decide on which language you would want to use, this site is for the properties of the elements on a live table. Once the table figure is active, identify the property of interest on the left and just click on any element in the Table and the details of the property sought for come along. It will display the information wanted. Very interactive and enjoyable.

Rationale: This site is a quite relevant for it may complement the notes in this module. In addition, it is also for leisure reading as there are more details and topics that due to the time duration and the calling for this module, could not be covers. For those interested in obtaining the most by selecting a language of competence, this is the place to visit.



Useful Link #4

Title : REGIONS OF PERIODIC TABLE & ELEMENTAL PROPERTIES.

URL : www.chemicool.com

Screen capture :

Legend

Li Solid	Cs Liquid	H Gas	Tf Synthetic
Alkali metals	Alkali earth metals	Transition metals	Rare earth metals
Other metals	Noble gases	Halogens	Other nonmetals

Periodic Table first discovered in 1869 by Dmitry I. Mendeleev is a way of presenting all the elements so as to show their similarities and differences. The elements are arranged in increasing order of **atomic number(Z)** as you go from left to right across the table. The horizontal rows are called **periods** and the vertical rows, **groups**.

A **noble gas** is found at the right hand side of each period. There is a progression from metals to non-metals across each period. Elements found in groups (e.g. alkali, halogens) have a similar electronic configuration. The number of electrons in outer shell is the same as the number of the group (e.g. lithium 2:1).

The block of elements between groups II and III are called **transition metals**. These are similar in many ways; they produce colored compounds, have variable valency and are often used as **catalysts**. Elements 50 to 71 are known as **lanthanide** or rare earth elements. These elements are found on earth in only very small amounts.

Elements 90 to 103 are known as the actinide elements. They include most of the well known elements which are found in **nuclear reactions**. The elements with larger atomic numbers than 92 do not occur naturally. They have all been produced artificially by bombarding other elements with particles.

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Description: Provides an interactive Periodic Table that includes information about elements and their states, energies, characteristics, abundance, and costs. It wraps up most of the properties discussed in Units 1, 2, 3, and 4 in a very simple and easy-to-understand graphic illustration.

Rationale: This site is a virtual summary of the module, the arrangement is such that each element is discussed here fully, from its place in the periodic table to its economic uses. It further classifies the periodic table into those regions of our discussions; regions such as s, p, d, actinides, etc. areas.

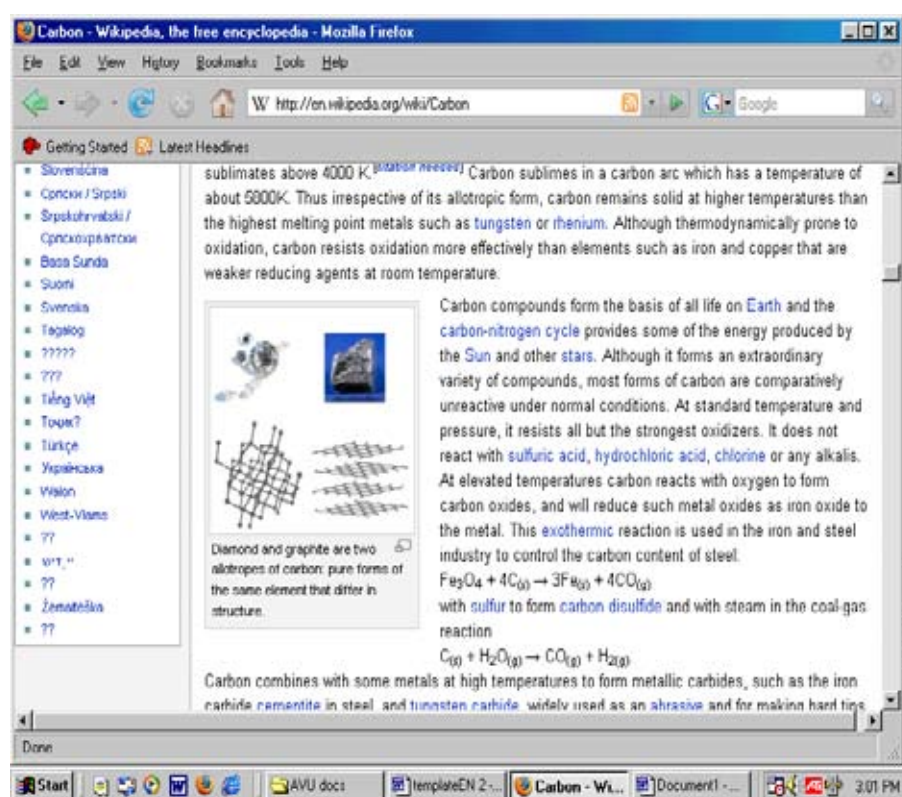


Useful Link #5

Title : PROPERTIES AND USES OF ELEMENTS

URL : <http://en.wikipedia.org/wiki/Carbon>

Screen capture :



Description: We get for any selected element (you have to substitute for the element in the URL) all the physical and chemical properties and the possible isotopes, and allotropes associated with it. It is a must-read site for a chemistry student.

Rationale: To get the full description of an element; its atomic mass, electronic configuration, isotopy, allotropy and even uses, this is the site to visit. Here I have it for carbon, you may seek for any element of your choice.

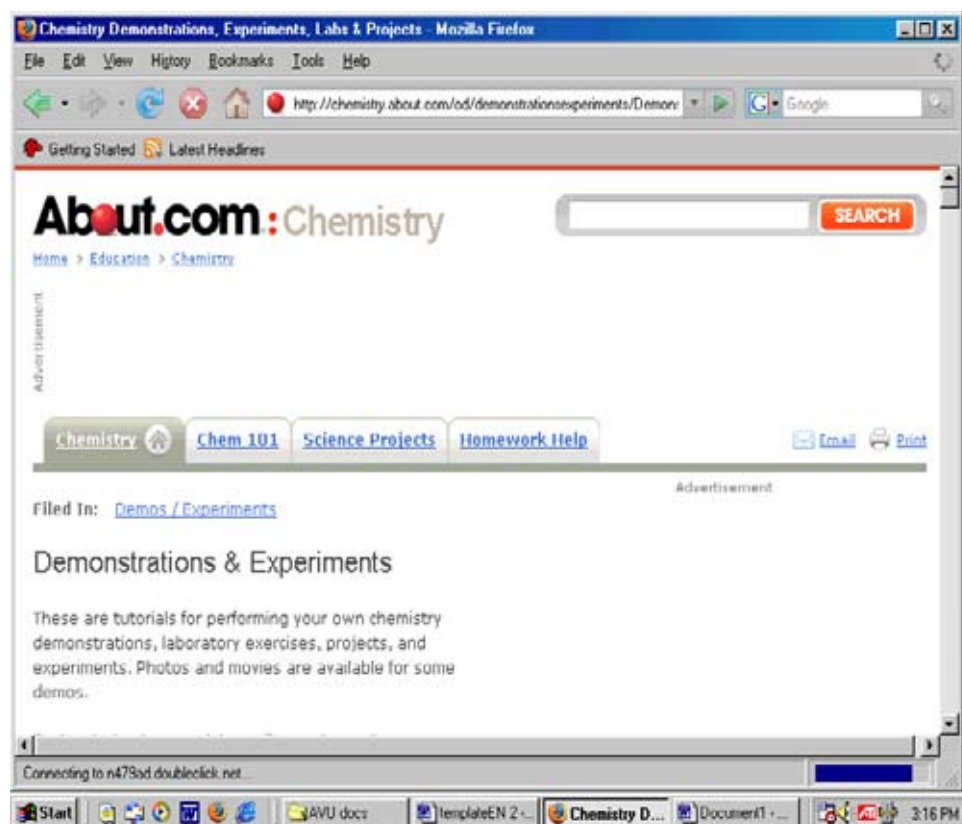


Useful Link #6

Title : CLASSIFICATION OF CHEMICAL ELEMENTS

URL : <http://chemistry.about.com/od/elementgroups/a/metals.htm>

Screen capture :



Description: Outlines what properties distinguishes metals from the other elements. The site similarly shows how to conduct simpler chemistry experiments at home and also in the labs.

Rationale: A part from the serious issues of differences in the characteristics of classes of elements, the site enables us to see the light and fun side of chemistry as investigative science. This site demystifies chemistry.



Useful Link #7

Title : THE PROPERTIES OF HALOGENS

URL : http://www.chemsoc.org/viselements/pages/data/intro_groupvii_data.html

Screen capture :

Visual Elements: Group 17 - The Halogens - Mozilla Firefox

http://www.chemsoc.org/viselements/pages/data/intro_groupvii_d

Group 17 - *The Halogens*

The elements of Group 17, the Halogens, are:

	symbol	electron configuration
fluorine	F	$[\text{He}]2s^2 2p^5$
chlorine	Cl	$[\text{Ne}]3s^2 3p^5$
bromine	Br	$[\text{Ar}]3d^{10} 4s^2 4p^5$
iodine	I	$[\text{Kr}]4d^{10} 5s^2 5p^5$
astatine	At	$[\text{Xe}]4f^{14} 5d^{10} 6s^2 6p^5$

All the isotopes of astatine are radioactive, and so this element will not be considered further here.

Appearance

Fluorine is a poisonous pale yellow gas, chlorine is a poisonous pale green gas, bromine is a toxic

Description: This is a complete site for the physical and chemical properties of halogens. It includes all that we ever, for the purposes of this module, about halogens and their chemistry.

Rationale: This site discusses the appearances, the general reactivity, occurrences and attractions and the compounds of halogens with other elements. It discusses the properties of their oxides, oxoacids, halides, organic compounds, electron affinity and many more properties of the halogens.

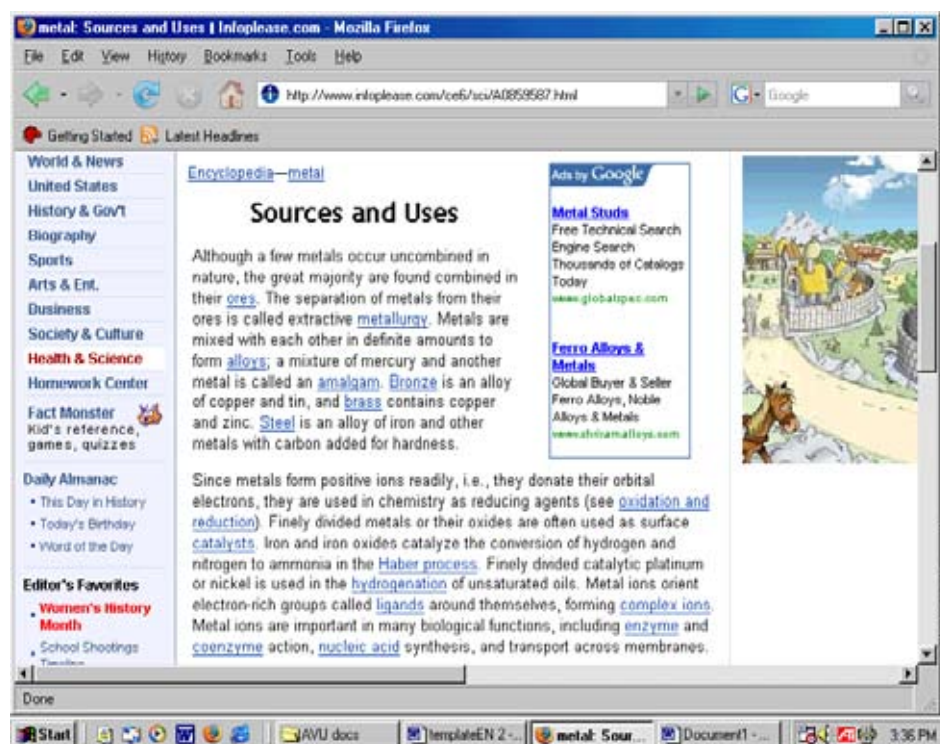


Useful Link #8

Title : EXCTRACTION OF ELEMENTS

URL : <http://www.infoplease.com/ce6/sci/A0859587.html>

Screen capture :



Description: The site is relevant and useful for the introduction to metal extractions. The details of some specific methods are also included.

Rationale: Largely, this site helps in better undertaking of the unit 5. The methodologies and the reasons for specific elements extraction details are all discussed herein. However, the learner is advised to explore all those pull-outs so as to get the full benefit of the site.

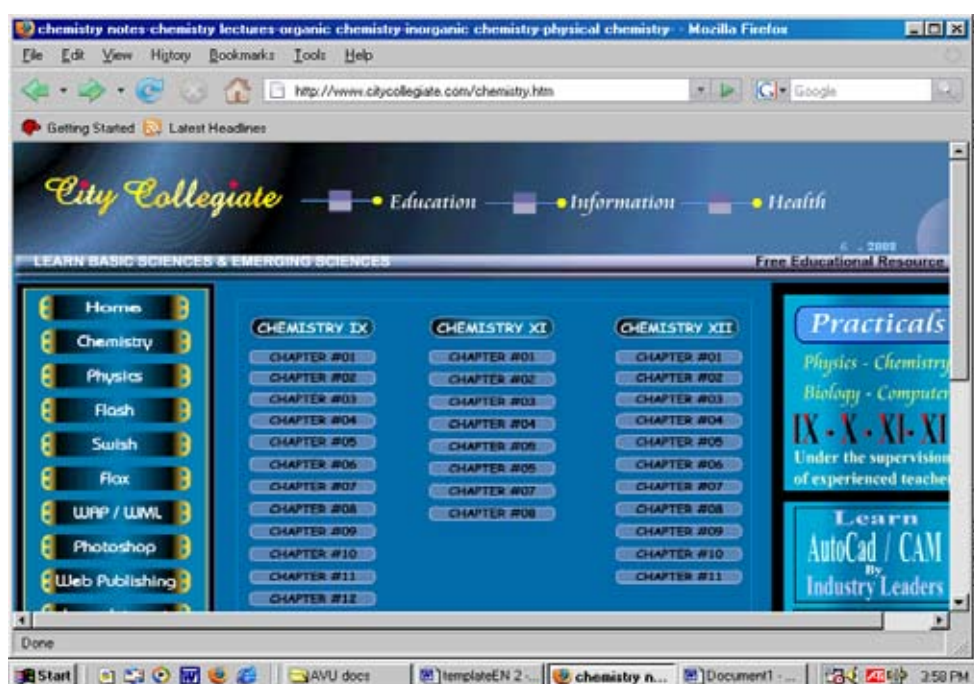


Useful Link #9

Title : INORGANIC CHEMISTRY COURSE NOTES

URL : <http://www.citycollegiate.com/chemistry.htm>

Screen capture :



Description: This site is for freshers chemistry. It is designated in relevant chapters. Take time to go through them. For example, chapters 3 to 10 would be free source notes and learning materials for this module.

Rationale: The site offers information in basic chemistry for beginners and also other subjects. This is a broad based site and would help the learner to navigate more than one topic at a time. Each chapter for different levels would give details commensurate to that particular level of chemistry.



Useful Link #10

Title : USES OF ELEMENTS

URL : <http://www.corrosionsource.com/handbook/periodic/14.htm>

Screen capture :

The screenshot shows a web browser window displaying the Corrosion Source website. The page title is "Chlorine". The browser address bar shows the URL: <http://www.corrosionsource.com/handbook/periodic/17.htm>. The website header includes the Corrosion Source logo, navigation links (Mission Statement, Site Map, Help Desk, Advertise), and a search bar. A banner for "Predict 4.0 All New Features & Lab Data" is visible. The breadcrumb trail reads: "You are here: Home > Handbook > Periodic Table of Elements".

Atomic Number:	17	Atomic Symbol:	Cl
Atomic Weight:	35.453	Electron Configuration:	2-8-7
Shells:	2,8,7	Filling Orbital:	3p ⁵
Melting Point:	97.8°C	Boiling Point:	882.9°C
Uses:	salt, water purification, bleaches, CFC's, etc.		

Description: The site makes it possible for one to browse on the uses and properties of any selected method from the presented periodic table of the elements.

Rationale: For a complete and updated uses of various elements of the periodic table, this site is necessary. The default element in the URL is sulphur, but take time to navigate for any element of choice. For example for Chlorine, in place of 14 in the URL, put 17.



XIV. Synthesis of the Module

Having gone through this course, you can appreciate that to better understand both the physical and chemical properties of elements and to infer their uses and applications, one needs the knowledge of electronic configurations and the periodic table of elements. We can now classify the elements as metals, non-metals and metalloids or simply as s and p block elements. At this point, as a learner, you are expected to be in a position to clearly explain why an element belongs to a particular family (group) of the periodic table, predict their relative melting points, atomic sizes, ionisation energies, and to appreciate the trends of their electronegativities, nuclear binding energies and effective nuclear charge among other properties. Above all, the concept of assigning a particular valency and oxidation numbers to an element is now apparent.

Unit 3 brings to light why different groups of elements behave the way they do and to understand their physical and chemical properties. These properties are linked to their positions in the periodic table. Furthermore, from Unit 4, you should be in a comfortable position to explain why an extraction process is preferred for a particular class of elements and also to appreciate their natural abundance and how they occur in nature.

Lastly, unit 5, captures the wish of all humanity and we should be in a position to help; how to better use these natural resources. This module makes it easier for us to appreciate, predict and understand the possible economic uses and applications of the various classes of the elements of the periodic table and their compounds.

As academic measure, and to help you develop the conceptual and quantitative problem solving skills, exercises and quizzes are interspersed within the text and a comprehensive one at the end.



XV. Summative Evaluation

The test below evaluates your overall understanding of the entire course and is therefore taken after the coverage of the entire coursework. Be with a periodic table and a calculator (Contributes 40 % of the total).

- How many elements are in period 6?
a) 12 b) 18 c) 32 d) 6
- Which of the following is a metalloid?
a) magnesium b) germanium c) phosphorus, d) Lead
- Which of the following is a non-metal?
a) magnesium b) osmium c) carbon d) Mercury
- Which property describes non-metals? a) conduct electricity very well, b) do not react at all, c) have negative oxidation states. d) are located on the right-hand-side of the periodic table
- Which element has the largest atomic radius?
a) Sn b) P c) Cs d) Xe
- What is an isoelectronic series?
a) series for atoms of same group,
b) similar electronic configuration,
c) isotopes of the same element
d) members of same period in periodic table.
- Which of the following sets are isoelectronic with each other?
a) K^+ , Ca^{+2} , Ar
b) Al, Mg^{2+} , Si
c) Ar, Ne, Kr
d) Li, Be, N, F
- Arrange the atoms and ions in order of increasing size, Li^+ , Be^{+2} , and Cl^- .
a) $Cl^- < Li^+ < Be^{+2}$
b) $Be^{+2} < Cl^- < Li^+$
c) $Be^{+2} < Li^+ < Cl^-$
d) $Cl^- < Be^{+2} < Li^+$



- 9 Arrange the following in increasing melting points, I, Cl, Br, F ?
a) $F < Cl < Br < I$, b) $I < Br < Cl < F$ c) $F < Br < Cl < I$
10. Arrange the following in increasing metallic characteristics, Na, Al, Rb, C ?
a) $Al < Cs < C < Na$
b) $Cs < Na < Al < C$
c) $Al < Na < Cs < C$
d) $C < Na < Cs < Al$
11. What is the formula for the product of a reaction between Germanium and excess oxygen.
a) Ge_4O_2 , b) GeO_2 , c) Ge_2O d) Ge_2O_4 .
12. What is the formula for the product of a reaction between Arsenic and Fluorine,
a) AsF_5 , b) As_5F , c) As_7F_5 , d) As_3F_7
13. What name is given to the amount of energy required to remove the outermost electron from an atom in a gaseous phase?. a) Ionization energy, b) electron affinity, c) electronegativity, d) Valence energy.
14. The transfer of electrons from sodium atoms to chloride atoms result in the formation of;
a) coordinate covalent bonds,
b) Polar covalent bonds,
c) non-polar bonds,
d) ionic bonds
15. What type of bond exists in a molecule of iodine?
a) ionic,
b) polar covalent,
c) non-polar covalent,
d) metallic bond
- 16 Do metal oxides behave as:
a) acids,
b) bases,
c) neutral,
d) amphoteric compounds when added to water?
17. Which compound contains both ionic and covalent bonds?.
a) HCl, b) NaCl, c) NH_4Cl , d) CCl_4
18. Which compound is a network solid?.
a) CH_4 , b) CO_2 , c) CaH_2 , d) SiO_2



19. A solid surface is soft, has a low melting point, and a poor conductor of electricity. The substance is most likely to be?
- an ionic solid,
 - a network solid,
 - a metallic solid,
 - a molecular solid.
20. Which statement best explains why CCl_4 is a non-polar molecule?.
- C and Cl are non-metals,
 - the C-Cl bonds are non-polar,
 - CCl_4 is a liquid at room STP,
 - The CCl_4 molecule is symmetrical
21. The formula for Lead (IV) oxide is
- PbO ,
 - Pb_2O ,
 - PbO_2 ,
 - Pb_4O_2
22. The correct formula for Chlorous acid is
- HClO ,
 - HClO_2 ,
 - HClO_3 ,
 - HClO_4 .
23. Using electronegativity as a guide, which formula is correctly written
- F_6S ,
 - Cl_2O ,
 - Br_4C ,
 - I_3P
24. The correct name for the compound NaClO_3 is sodium
- Chloride,
 - Chloric,
 - Chlorate,
 - Chlorite
25. Which electron configuration represents the atom in the period 2 with the lowest ionization energy
- $1s^22s^1$,
 - $1s^12s^2$,
 - $1s^22s^22p^1$,
 - $1s^22s^22p^2$
26. In general, the element with the lowest ionization energy would be classified as
- halogens,
 - noble gases,
 - metals,
 - non-metals
27. A neutral oxygen atom (O) differs from an ion of oxygen (O^{2-}) in that the atom has,
- more electrons,
 - fewer electrons,
 - more protons,
 - fewer protons
28. An element that has both a high ionization energy and high electronegativity is most likely to be
- metal,
 - nonmetal,
 - metalloid,
 - noble gas.
29. The atoms of the most active non-metals have,
- small atomic radius and high ionization energy,
 - small atomic radii and low ionization energy,
 - large atomic radii and low ionization energy,
 - large atomic radii and high ionization energy



30. As one proceeds from the left to right in period 3, of the periodic table, there is a decrease in
- ionization energy,
 - electronegativity,
 - metallic character,
 - valence electrons.
31. As one proceeds from right to left across period 3 of the periodic table, the number of electrons in the 2p subshell,
- decrease,
 - increase,
 - remain the same,
 - increase then decrease
32. Which of the group 17 elements has the greatest nuclear charge,
- F,
 - Cl,
 - Br,
 - I.
33. Which element will form a compound with the general formula M_2CO_3 , when it combines with carbonate ions,
- beryllium,
 - aluminum,
 - calcium,
 - lithium.
34. As the elements are considered from top to bottom of the periodic table in group 15, the ionization energy
- decreases,
 - increases,
 - remains the same,
 - increases then remains the same.
35. Considered in succession, the elements in period 2 of the table show decreased atomic radius with increasing atomic numbers. This may best be explained by the fact that the
- nuclear charge increases,
 - number of principal energy levels increase,
 - number of neutrons decreases,
 - electronegativity decrease.
36. During the electrolysis of fused NaCl, which reaction occurs at the positive electrode?,
- Chloride ions are oxidized,
 - Chloride ions are reduced,
 - Sodium ions are oxidized,
 - sodium ions are reduced.



37. Which group includes elements that are generally obtained from their compounds by electrolysis,
- a) Al, Mg, Na,
 - b) C, Ca, Na,
 - c) Cl, P, Na,
 - d) O, Al, Ar
38. The energy equivalent of the mass defect is called,
- a) an alpha particle,
 - b) binding energy,
 - c) half-life,
 - d) kinetic energy.
39. Which can be used as a coolant in a nuclear reactor,
- a) graphite,
 - b) sodium,
 - c) Carbon dioxide,
 - d) water
40. When neutrons and protons combine to form a stable nucleus, the actual mass of the nucleus compared to the total mass of its parts is always,
- a) less,
 - b) greater,
 - c) the same,
 - d) sometimes greater sometimes less.

Answer Key

- | | | | |
|--------|--------|--------|--------|
| 1. c, | 13. a, | 25. a, | 37. a, |
| 2. b, | 14. d, | 26. c, | 38. b, |
| 3. c, | 15. c, | 27. b, | 9. b, |
| 4. c, | 16. b, | 28. b, | 40. a. |
| 5. c, | 17. c, | 29. a, | |
| 6. b, | 18. d, | 30. c, | |
| 7. a, | 19. d, | 31. a, | |
| 8. c, | 20. d, | 32. d, | |
| 9. a, | 21. c, | 33. d, | |
| 10. b. | 22. b, | 34. a, | |
| 11. b, | 23. b, | 35. a, | |
| 12. a, | 24. c, | 36. a, | |



XVI. References

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XVII. Student Examination Records

Student Name	Continuous assessments and quizzes					Final Exams	Cumulative total Marks
	Unit 1	Unit 2	Unit 3	Unit 4	Unit 5		
	7 %	15 %	20 %	8 %	10 %		
						40 %	100 %

XVIII. Main Author of the Module

Dr. Chrispin Kowenje is a lecturer in the Department of Chemistry Maseno University Kenya. Kowenje joined Egerton University, Njoro Kenya for both his BSc. degree in Chemistry and Mathematical statistics as well as for his MSc. in Chemistry. The MSc. work (1998), which inspired his interest in the inorganic chemistry, was on Soil factors that affect fertilizers dissolution in soils. From 1998-2002, Dr. Kowenje was an assistant lecturer of Chemistry at Maseno University Kenya. In 2002, he then proceeded to the State University of New York (SUNY) at Binghamton for his Doctoral studies. His PhD work was in the chemistry of cations exchanged zeolite. Currently, (since 2006), Dr. Kowenje is interested in the interactions of zeolites with various environmental contaminants.

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Accompanying files

1. Reading 1 (PDF Compulsory reading for unit 1)
2. Reading 2 (PDF Compulsory reading for unit 2)
3. Reading 3 (PDF Compulsory reading for unit 3)
4. Reading 4 (PDF Compulsory reading for units 4 & 5)
5. FICN 01 – HCCl_3 Molecule – doc.
6. FICN 02 – CCl_4 Molecule – doc
7. FICN 03 – group 13 elements – pdf
8. FICN 04 – group 1 & 17 electronegativity trends – word doc.
9. FICN 05 – ionization trends – word doc.
10. FICN 06 – general chemistry of non-metals -pdf
11. FICN 07 – group 14 chemistry - pdf
12. FICN 08 – electron sharing in bonds – doc