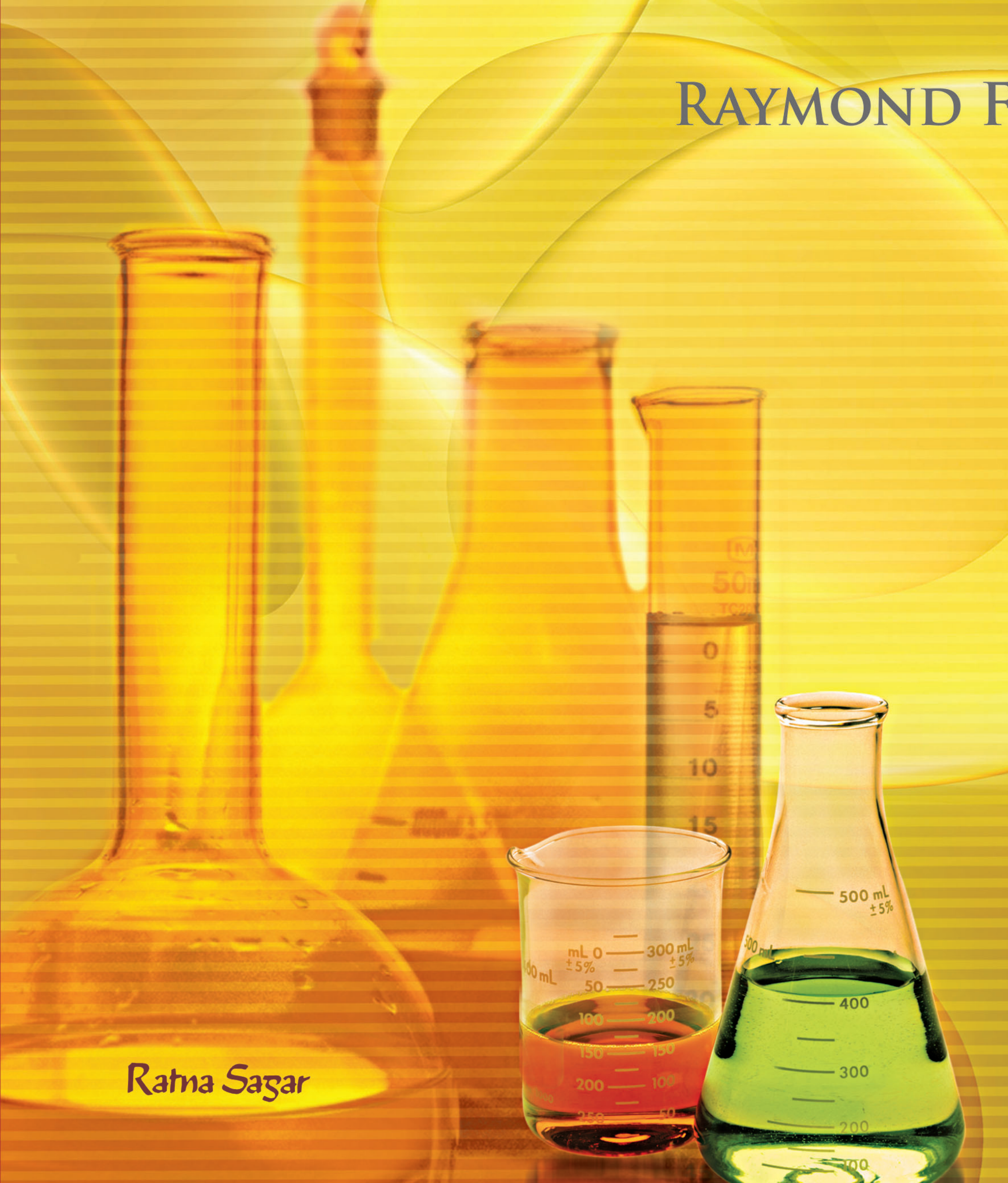
AS PER THE LATEST ICSE SYLLABUS

Revised and Updated

LIVING SCIENCE
CHEMISTRY

RAYMOND FERNANDES



CLASS 10

Ratna Sagar

Based on the latest syllabus prescribed by the
Council for the **I**ndian **S**chool **C**ertificate **E**xaminations

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CHEMISTRY
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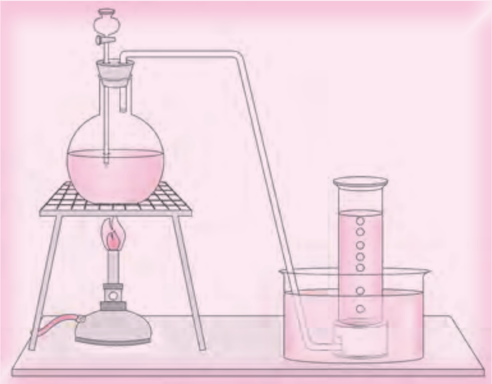

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Preface

Living Science Chemistry for Class X covers the latest syllabus for Chemistry as prescribed by the Council for the Indian School Certificate Examinations (ICSE), New Delhi. The sequencing of the chapters follows the distinct directives of the Board.

I have always maintained that Chemistry as a subject is an acquired taste. It needs to be delivered in a manner that will be appreciated and evict a positive response. From a teacher's point of view, the text needs to touch upon all the essentials that the student has to be informed about, yet not be excessively verbose and pedantic. There should also be a proper plan of instructions for each lesson to facilitate comprehensive and logical teaching. Keeping this in mind, each chapter begins with OBJECTIVES that highlight the expectations of the particular chapter and serve as a lesson plan for individual teachers.

From the student's point of view, the text has to be simple enough for them to read, understand and prepare for the examination. If chapters are only in point form, the essence of a subject is completely lost. It becomes purely an exercise of 'learning by rote', which contradicts the concept of holistic education of the twenty-first century.

The DO I KNOW? section will help students to assess their learning before moving on to the process of actually attempting the questions. The SUMMARY section provides the synopsis, which on a systematic study will furnish all relevant answers to exam-based questions.

All relevant questions of previous ICSE examinations have been incorporated apart from other varieties of questions. For those who would like to go beyond the question-answer syndrome, the book has enough information to arouse their imagination.

I have always believed that no chapter in Chemistry is a watertight unit and a link between the chapters needs to be made. Hence I have designed a MIND MAP, like a Revision sheet at the end of the book.

I hope the book will be an effective tool to acquire basic knowledge of a subject so central to our very existence. With this confidence I present this book, which will be of worth for both the teachers and the students alike.

I wish to thank the publishers for their support in this endeavour.

Suggestions for any improvement will be gratefully acknowledged.

Author

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1

THE PERIODIC CLASSIFICATION

> *"Chemistry creates its objects, and this creative faculty is similar to that of art itself, (and) essentially distinguishes it from the natural and historical Sciences."*
>
> — MARCELIN BERTHELOT

OBJECTIVES

After having studied this chapter the students will familiarize themselves with

- Early attempts of classifications
- The layout of the Modern Periodic Table
- The periodicity of properties
  - atomic radius
  - ionization potential
  - electron affinity
  - electronegativity
  - metallic characteristic
  - non-metallic characteristic
- Relation between atomic numbers and atomic mass
- An introduction to alkali metals
- An introduction to halogens

SYLLABUS AND SCOPE

- *Periodic properties and their variations in groups and periods*

  Definitions of following periodic properties and trends in these properties in groups and periods should be studied:
  - atomic size
  - metallic character
  - non-metallic character
  - ionization potential
  - electron affinity
  - electronegativity
- *Periodicity on the basis of atomic number for elements*

  Relation between atomic number for light elements (proton number) and atomic mass for light elements; the Modern Periodic Table up to Period 3 (students to be exposed to the complete Modern Periodic Table but no questions will be asked on elements beyond Period 3—Argon); periodicity and other related properties to be described in terms of shells (not orbitals); special reference to the alkali metals and halogen groups.

HISTORICAL BACKGROUND

Knowledge of a large number of elements made it absolutely necessary to classify them. A systematic classification would help in better understanding and dealing with the vast number of elements.

1. A classification of elements may facilitate their study.
2. A classification may lead to correlating the properties of elements with some fundamental property or characteristic of an element.
3. A classification may further reveal the relationship between one element and the other.

SOME EARLY ATTEMPTS OF CLASSIFICATION

As the number of discovered elements increased, chemists began to find patterns in their properties. Some early attempts were made in systematic classification. We have read about them in previous classes. Let us discuss some of them here in brief.

Döbereiner's Triad

Law

In a triad the atomic weight of the middle element is almost the arithmetic mean of the other two.

Merits

The law of triad recognized the relationship between atomic weight and properties of elements. For example, lithium, sodium and potassium have similar chemical properties and form a triad.

Demerits

Not many elements could be grouped into triads.

Newland's Law of Octaves

Law

If elements are arranged in an increasing order of atomic weights, the eighth element starting from a given one will have similar properties to the one started with.

Table 1.1 MENDELEEV'S PERIODIC TABLE OF 1872 (BASED ON ATOMIC MASS)

GROUP	I		II		III		IV		V		VI		VII		VIII		
Oxide+ Hydride+	R_2O RH		RO RH_2		R_2O_3 RH_3		RO_2 RH_4		R_2O_5 RH_3		RO_3 RH_2		R_2O_7 RH		RO_4		
Periods ↓	A	B	A	B	A	B	A	B	A	B	A	B	A	B	Transition Series		
1	H 1.008																
2	Li 6.939		Be 9.012		B 10.81		C 12.011		N 14.007		I 15.999		F 18.998				
3	Na 22.99		Mg 24.31		Al 29.988		Si 28.09		P 30.974		S 32.06		Cl 35.453				
4 First Series	K 39.102		Ca 40.08			* 44.96		Ti 47.90		V 50.94		Cr 50.20		Mn 54.94	Fe 55.85	Co 58.93	Ni 58.71
Second Series		Cu 63.54		Zn 65.37	* 69.72		* 72.59		As 74.92		Se 78.96		Br 79.09				
5 First Series	Rb 85.47		Sr 87.62			Y 88.91		Zr 91.22		Nb 92.91		Mo 95.94		* 99	Ru 101.07	Rh 102.91	Pd 106.4
Second Series		Ag 107.87		Cd 112.40	In 114.82		Sn 118.69		Sb 121.75		Te 127.60		I 126.90				
6 First Series	Cs 132.90		Ba 137.34			La 138.91		Hf 178.49		Ta 180.95		W 183.85			Os 190.2	Ir 192.2	Pt 195.09
Second Series		Au 196.97		Hg 200.59	Tl 204.37		Pb 207.19		Bi 208.98								

+The letter R was used in the formula of oxides and hydrides to represent any of the elements in the group.

*Elements not discovered in 1872, but Mendeleev left gap for unknown elements which were discovered subsequently.

Merits

Atomic weights were made the basis of classification.

Periodicity of properties was recognized for the first time.

Demerits

The law of octaves failed because firstly the existence of transition elements was not realized and secondly Newland was very rigid in his views and was not prepared to accept any exceptions to his law.

Mendeleev's Periodic Law

Law

The physical and chemical properties of elements are the periodic functions of their atomic weights.

Merits

Correction of atomic weights were marked and undiscovered elements were predicted.

Demerits

1. Position of hydrogen was not defined.
2. Lighter elements were placed after heavier elements, contrary to the law.
3. The position of rare earth elements and isotopes was ambiguous.
4. Chemically similar elements were placed in different groups and vice versa.
5. The Periodic Table did not reflect the electronic configuration of elements.

The Modern Periodic Law

Law

The physical and chemical properties of elements are the periodic functions of their atomic numbers.

Merits

It is the most appropriate classification till date.

Demerits

The position of hydrogen is still not certain.

THE LAYOUT OF THE MODERN PERIODIC TABLE

1. The Periodic Table has 18 vertical columns called groups and 7 horizontal rows called periods.
2. Elements in the same group have similar properties.
3. Each group is sub-divided into 2 parts:
   a. Normal Elements
   b. Transition Elements

4.

Group IA	Alkali metals	[H, Li, Na, K, Rb, Cs, Fr]
Group IIA	Alkaline earth metals	[Be, Mg, Ca, Sr, Ba, Ra]
Group VIIA	Halogens	[F, Cl, Br, I, At]
Group Zero	Noble gases	[He, Ne, Ar, Kr, Xe, Rn]

5.

Period 1	Shortest period	2 elements
Periods 2 and 3	Short periods	8 elements
Periods 4 and 5	Long periods	18 elements
Period 6	Longest period	32 elements
Period 7	Incomplete period	26 elements

6. The inner transition elements:

Lanthanide Series

Position	Group III B, Period 6
Number of elements	14

Actinide Series

Position	Group III B, Period 7
Number of elements	14

7. All elements in the same period will have the same number of shells but their valence electrons will increase gradually on going from left to right across the period.

The number of valence electrons of normal elements corresponds to the group number, e.g. nitrogen belongs to Group VA and therefore, it will have 5 electrons in the outer shell.

Transition elements or B group elements will always have 2 electrons in their outer shell. It is for this reason that they are also referred to as transition metals.

8. All elements in the same group will have the same number of valence electrons but the number of shells will increase on going from top to bottom in the group. The number of shells corresponds to the period number, e.g. potassium belongs to the 4th Period and therefore, it will have 4 shells.

DID YOU KNOW?

Gold derives its symbol from the Latin word, aurum, which means 'shining dawn'.

Helium derives its name from the Greek word, helios, which means sun. It was discovered by spectroscopy during a solar eclipse.

TYPES OF ELEMENTS

The elements can be classified as:

1. Representative or Normal elements
2. Transition elements
3. Inner transition elements
4. Inert gases

Normal elements are also called 'A' group elements. These elements have their last shell being filled.

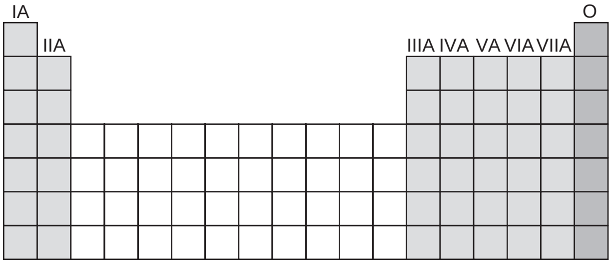


Fig. 1.1 Position of normal elements in the periodic table.

The different families of normal elements are as follows:

Group IA	Group 1	Alkali metals	H, Li, Na, Rb, Cs, Fr
Group IIA	Group 2	Alkaline earth metals	Be, Mg, Ca, Sr, Ba, Ra
Group III A	Group 13	Boron family	B, Al, Ga, In, Tl
Group IV A	Group 14	Carbon family	C, Si, Ge, Sn, Pb
Group V A	Group 15	Nitrogen family	N, P, As, Sb, Bi
Group VI A	Group 16	Oxygen family	O, S, Se, Te, Po
Group VII A	Group 17	Halogens	F, Cl, Br, I, At
Zero		Noble gases	He, Ne, Ar, Kr, Xe, Rn

The characteristics of these elements vary from being strongly metallic to strongly non-metallic. Although it is inappropriate to generalize we can safely say that the group of alkali metals and alkaline earth metals are strongly metallic as the name suggests and the group of halogens is strongly non-metallic.

The **transition elements** are placed in between the normal elements. They are also referred to as 'B' group elements. Transition elements are so called because their position in the periodic table is between the metals and the non-metals. The properties of these elements are also transitional between metals and non-metals. The transition elements have two electrons in the last shell and an incomplete second last shell. Since the last shell has two electrons, the transition elements have metallic properties. These elements occupy group 3 to group 12 of the periodic table.

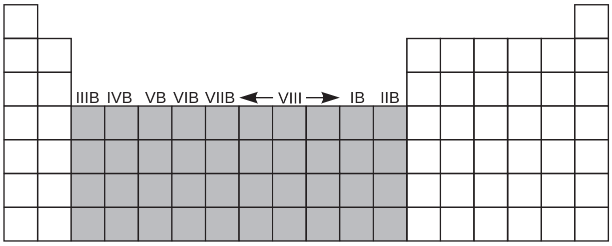


Fig. 1.2 Position of transition elements in the periodic table.

Properties of transition elements:

- All transition elements are metals.
- Most transition elements display variable valency.
- The size of the atomic radii decreases from left to right across a row in the transition elements.
- The melting and boiling points of transition metals are usually very high, exception being Zn (420 °C), Cadmium (321 °C) and Hg (liquid, hence melting point is –38 °C).
- Many transition metals are sufficient electropositive and react with mineral acids to liberate hydrogen. However, some metals remain unreactive, e.g. Ru (Ruthenium), Rh (Rhodium), Pd (Palladium), Os (Osmium), Ir (Iridium), Pt (Platinum) and Au (Gold).
- Many ionic and covalent compounds of transition elements are coloured. This is in contrast with compounds of the normal elements that are generally white.

Inner transition elements are placed within the transition elements. There are two series of fourteen elements each. These series are called the Lanthanide series and the Actinide series.

All lanthanides are placed in group III B, period 6 of the periodic table. Actinides occupy group III B, period 7 of the periodic table.

Properties of inner transition elements

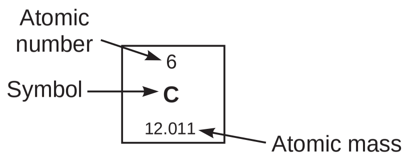
- The inner transition elements have their last three shells partly filled while the remaining inner shells are completely filled.

Table 1.2 LONG FORM PERIODIC TABLE

GROUP NUMBER

PERIOD NUMBER

IUPAC (1984)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	IA	IIA	IIIB	IVB	VB	VIB	VIIB				IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
								VIII										
1	1 **H** 1.0079																	2 **He** 4.0026
2	3 **Li** 6.941	4 **Be** 9.01218											5 **B** 10.811	6 **C** 12.011	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.998403	10 **Ne** 20.179
3	11 **Na** 22.9897	12 **Mg** 24.305											13 **Al** 26.98154	14 **Si** 28.0855	15 **P** 30.9734	16 **S** 32.066	17 **Cl** 35.453	18 **Ar** 39.948
4	19 **K** 39.0983	20 **Ca** 40.08	21 **Sc** 44.9559	22 **Ti** 47.88	23 **V** 50.9415	24 **Cr** 51.996	25 **Mn** 54.9380	26 **Fe** 55.847	27 **Co** 58.9332	28 **Ni** 58.69	29 **Cu** 63.546	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.9216	34 **Se** 78.96	35 **Br** 79.909	36 **Kr** 83.80
5	37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.9059	40 **Zr** 91.22	41 **Nb** 92.9064	42 **Mo** 95.94	43 **Tc** (99)	44 **Ru** 101.07	45 **Rh** 102.905	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.69	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.904	54 **Xe** 131.29
6	55 **Cs** 132.9054	56 **Ba** 137.34	57 **La*** 138.91	72 **Hf** 178.49	73 **Ta** 180.9479	74 **W** 183.85	75 **Re** 186.206	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.9665	80 **Hg** 200.59	81 **Tl** 204.383	82 **Pb** 207.2	83 **Bi** 208.9804	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** 226.0254	89 **Ac**** 227.028	104 **Rf** 261.11	105 **Db** 262.1	106 **Sg** 263	107 **Bh** 262.12	108 **Hs** (265)	109 **Mt** (266)	110 **Ds** (269)	111 **Rg** (272.2)	112 **Cn** (285)	113 **Uut** (284)	114 **Fl** (289)	115 **Uup** (288)	116 **Lv** (293)	117 † **Uus**	118 **Uuo** (294)



*LANTHANIDES	58 **Ce** 140.12	59 **Pr** 140.9077	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.9254	66 **Dy** 162.50	67 **Ho** 164.9304	68 **Er** 167.26	69 **Tm** 168.9342	70 **Yb** 173.04	71 **Lu** 174.967
**ACTINIDES	90 **Th** 232.0381	91 **Pa** 231.0359	92 **U** 238.0289	93 **Np** 237.0482	94 **Pu** (242)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (254)	101 **Md** (258)	102 **No** (259)	103 **Lr** (260)

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† The element 117 is not known but is included to show its expected position.

- Elements within each of these series are very similar in chemical properties.
- The inner transition elements display metallic characteristics.
- The metals are all soft and silvery white.
- They are electropositive and therefore reactive.

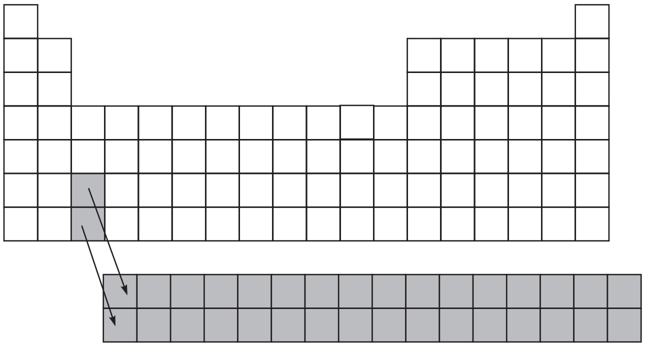

Fig. 1.3 Position of inner transition elements in the periodic table.

PERIODICITY IN PROPERTIES OF ELEMENTS ACROSS THE PERIODIC TABLE

Elements show periodicity in their properties because of their **valence shell configuration**. All elements showing periodicity in properties have the same number of electrons in the last or valence shell. Thus, the valence shell configurations of elements recur at regular intervals of atomic number and it is due to this fact that there is recurrence of chemical properties.

Factors that influence the trend in the periodicity of properties of elements are:

- Effective nuclear charge
- Number of orbits
- Screening effect

The properties that will be discussed here are:

- Atomic radius
- Ionization potential
- Electron affinity
- Electronegativity
- Metallic characteristic
- Non-metallic character

Atomic radius

An atom is made of a nucleus surrounded by an electron cloud which does not have a sharply defined boundary. This arbitrary or rather diffused nature of an atom makes it difficult to conceptualize the idea of an atomic radius. However, atomic radius is usually considered as the distance from the centre of the nucleus to the outermost shell, i.e. to a point where the electron density is effectively zero.

Table 1.3 ATOMIC RADII (IN PICOMETERS) OF FIRST TWENTY ELEMENTS

H 37							**He** 40
Li 152	**Be** 111	**B** 88	**C** 77	**N** 74	**O** 66	**F** 64	**Ne** 70
Na 186	**Mg** 160	**Al** 143	**Si** 117	**P** 110	**S** 104	**Cl** 99	**Ar** 94
K 227	**Ca** 197						

Trend

Across the period from left to right: The atomic radius decreases across the period from left to right.

Down the group: The atomic radius increases as one moves from top to bottom.

Reason

Across the period

Dominating factor: *Effective nuclear charge*

Along a period the effective nuclear charge increases. This is due to the fact that the number of electrons increases (in the same shell), increasing the number of protons in the nucleus. This pulls the valence shell of electrons in the atom towards itself, thus decreasing the atomic radius.

Down the group

Dominating factor: *Number of orbits*

As we move down a group the number of orbits keeps on increasing along with the number of electrons and subsequently the number of protons. The space required to accommodate the extra orbits takes prevalence and therefore the atomic size increases.

DID YOU KNOW?

The only two elements named after women scientists are Curium after Marie Curie and Meitnerium after Lise Meitner.

Ionization potential

It is the energy required to remove an electron from an isolated gaseous atom in its ground state.

Ionization energy is measured in electron volts (eV) which is related to kilocalories and kilojoules as follows:

$$1\ \text{eV/mol} = 23.06\ \text{kcal/mol} = 96.49\ \text{kJ/mol}$$

The energy required to remove the first electron from an isolated gaseous atom is called the **first ionization potential (IP_1)**.

The energy required to remove the second electron is called the **second ionization potential (IP_2)**.

$$IP_2 > IP_1$$

This is because as electrons are removed from the valence shell the effective nuclear charge as distributed over the remaining electrons increases, causing the second ionization potential to be greater.

Table 1.4 FIRST IONIZATION POTENTIAL (IN kJ/mol) OF FIRST TWENTY ELEMENTS

H 1311							**He** 2372
Li 520	**Be** 899	**B** 801	**C** 1086	**N** 1403	**O** 1410	**F** 1681	**Ne** 2080
Na 496	**Mg** 737	**Al** 577	**Si** 786	**P** 1012	**S** 999	**Cl** 1255	**Ar** 1521
K 419	**Ca** 590						

Trend

Across the period from left to right: The ionization potential increases from left to right, being the smallest for the alkali metal atoms and the largest for the noble gas atoms.

Along the group: The ionization potential decreases.

Reason

Across the period

Dominating factor: *Effective nuclear charge*

Along a period the effective nuclear charge increases. This causes the atomic radius to decrease thus getting the valence shell closer to the nucleus. This makes it difficult to remove electrons from the valence shell and therefore the ionization potential increases along a period from left to right.

Down the group

Dominating factor: *Number of orbits and the intermediate screening effect of electrons*

As we move down a group the number of orbits keeps on increasing along with the number of electrons. The distance from the nucleus coupled with the interference of the electron between the nucleus and the valence shell renders the valence electrons weakly bound to the nucleus. Thus, less energy is required to remove electrons from the valence shell. As a result the ionization potential decreases down a group.

Electron affinity

It is the amount of energy released when an isolated gaseous atom accepts an electron to form the gaseous negative ion. Electron affinity is expressed in terms of electron volts (eV) or kilojoule per mole (kJ/mol).

Table 1.5 ELECTRON AFFINITY (IN kJ/mol) FOR ONE ELECTRON BEING ADDED TO THE OUTERMOST SHELL (THE NEGATIVE SIGN INDICATES RELEASE OF ENERGY)

H –73							**He** 0
Li –60	**Be** ~0	**B** –27	**C** –122	**N** –0.1	**O** –141	**F** –328	**Ne** 0
Na –53	**Mg** ~0	**Al** –44	**Si** –134	**P** –72	**S** –200	**Cl** –349	**Ar** 0
K –48	**Ca** ~0						

Trend

Across the period from left to right: The electron affinity increases.

Down the group: The electron affinity decreases.

Reason

Across the period

Dominating factor: *Effective nuclear charge*

Along a period the effective nuclear charge increases and this decreases the atomic radius. Smaller the atomic radius, greater is the tendency of the atom to attract electrons. This increases the electron affinity along a period from left to right.

Down the group

Dominating factor: *Number of orbits*

As we move down a group the number of orbits keeps on increasing and therefore the atomic size increases. The increase in size and the decrease in the effective nuclear charge both favour the decrease in the electron affinity down a group.

Electronegativity

The tendency of an atom in a molecule to attract a shared pair of electrons towards itself is called electronegativity.

The electronegativity value indicates the nature of a chemical bond in the following way:

1. If the difference in electronegativities of the combining atoms is zero, then the bond formed

is a non-polar covalent bond. This is true for combinations between non-metals.

2. If the difference in electronegativities of the combining atoms is small, then the bond formed is polar covalent.
3. If the difference in electronegativities of the combining atoms is large, then the bond formed is ionic or electrovalent.
4. If the electronegativity difference is 1.7 or less then the bond is covalent. If the electronegativity difference is more than 1.7, then the bond is ionic.

Table 1.6 ELECTRONEGATIVITIES USING PAULING'S SCALE AS DETERMINED BY PAULING

H 2.1								**He**
Li 1.0	**Be** 1.5		**B** 2.0	**C** 2.5	**N** 3.0	**O** 3.5	**F** 4.0	**Ne**
Na 0.9	**Mg** 1.2		**Al** 1.5	**Si** 1.8	**P** 2.1	**S** 2.5	**Cl** 3.0	**Ar**
K 0.8	**Ca** 1.0							

A number of scales were devised to measure the electronegativities of atoms but the most important of them is the **Pauling's scale**. According to Pauling, fluorine is the most electronegative element followed by oxygen.

Trend

Across the period from left to right: The electronegativity increases.

Down the group: The electronegativity decreases.

Reason

Across the period

Dominating factor: *Effective nuclear charge*

Along a period the effective nuclear charge increases thus decreasing the atomic radius. This favours the increase in electronegativity and therefore the electronegativity of elements increases along a period from left to right.

Down the group

Dominating factor: *Number of orbits*

As we move down a group the number of orbits keeps on increasing and therefore the atomic size increases. Therefore, the electronegativity decreases.

Metallic character

Metallic character is defined as the tendency of an atom to lose electrons and form positive ion (cations).

Trend

Across the period from left to right: The metallic character decreases.

Down the group: The metallic character increases.

Reason

Across the period

Dominating factor: *Effective nuclear charge*

The metallic character depends upon atomic size and ionization potential. As we move along a period from left to right, the effective nuclear charge increases thus decreasing the atomic radius. This favours the increase in electronegativity and therefore the tendency to lose electrons is low. This accounts for the decrease in the metallic character along a period.

Down the group

Dominating factor: *Number of orbits*

The metallic character depends upon atomic size and ionization potential. As we move down a group the number of orbits keeps on increasing and therefore the atomic size increases. This means that the electronegativity decreases. These factors enhance the loss of electrons and therefore the metallic character increases down a group.

Non-metallic character

Non-metallic character is the tendency of an element to gain electrons and form a negative ion (anion).

Trend

Across the period from left to right: The non-metallic character increases.

Down the group: The non-metallic character decreases.

Reason

Across the period

Dominating factor: *Effective nuclear charge*

The non-metallic character depends upon atomic size and ionization potential. As we move along a period the effective nuclear charge increases thus decreasing the atomic radius. This favours the increase in electronegativity and therefore the tendency to gain electrons is high. This accounts for the increase in the non-metallic characteristic along a period.

Down the group

Dominating factor: *Number of orbits*

The non-metallic character depends upon atomic size and ionization potential. As we move down a group the number of orbits keeps on increasing and

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