3 Classification of Elements and Periodicity in Properties

The elements of the periodic table can be found everywhere; in fact, they comprise everything. When fired from artillery shells, white phosphorus, for example, flares in spectacular bursts with a yellow flame and produces dense white smoke. It is used to hide troop movements and to illuminate a battlefield.

Topic Notes

- *Periodic Table and Properties*
- *Properties of Elements and their Periodic Trends*

PERIODIC TABLE AND PROPERTIES

1

TOPIC 1

WHY DO WE NEED TO CLASSIFY ELEMENTS?

The matter is made up of elements. There are now 118 elements known. The newly found elements are all man-made. Attempts to synthesize new elements are still underway. It is quite difficult to examine the chemistry of all of these elements and their numerous compounds individually due to the enormous number. A method of classifying the elements or to organising them methodically will rationalise existing chemical facts regarding elements and will help to accurately predict new ones for future research.

Arranging the elements in such a way that similar elements are placed together while dissimilar elements are separated from one another is called the classification of elements. "Periodic table is the arrangement of the known elements according to their properties in a tabular form"

Table: Different periodic classification and its Discoveries

Periodic Classification	Discovered by	Description	Reasons for failure
Law of Triads	Johann Dobereiner	Each Triad's middle element had an atomic weight that was very close to the arithmetic mean of the other two elements. In addition, the middle element's characteristics were in the middle of the other two elements. E.g., Li Na, K; Ca, Sr, Ba; Cl, Br, I	Because it appeared to work only for a few elements, it was dismissed as coincidence.
Law of Octaves	John Alexander Newlands	Arranged the elements in ascending order of atomic weights, noting that every eighth element possessed characteristics comparable to the first element like the eighth note of musical scale. *Eg.* Li, Be, B, C, N, O, F. Na, Mg, Al, Si, P, S, Cl, K and Ca	Only for elements up to calcium Newlands' Law of Octaves appeared to be valid.
Mendeleev's Periodic Table	Dmitri Mendeleev and Lothar Meyer (Working independently)	Similarities in physical and chemical properties arise at regular intervals when elements are arranged in ascending order of their atomic weights.	Some of the elements did not fit in with his scheme of classification if the order of atomic weight was strictly followed. For example, iodine with a lower atomic weight than that of tellurium (Group VI) was placed in Group VII along with fluorine, chlorine, and bromine because of similarities in properties.

Mendeleev's Periodic Table

The Periodic Law was first published by Mendeleev according to which "The physical and chemical properties of the elements are a periodic function of their atomic weights." Mendeleev organized atoms in a table's horizontal rows and vertical columns in order of increasing atomic weights, so that elements with comparable properties were grouped in the same vertical column in such a way that the elements with similar properties occupied the same vertical column or group. He realized that some of the elements did not fit in with his scheme of classification if the order of atomic weight was strictly followed. He ignored the order of atomic weights, thinking that the atomic measurements might be incorrect, and placed the elements with similar properties together. For example, iodine with a lower atomic weight than that of tellurium (Group VI) was placed in Group VII along with fluorine, chlorine and bromine because of similarities in properties.

He proposed that some of the elements were still unidentified and, as a result, created some gaps in the table by grouping elements with similar properties together. Gallium and germanium, for example, were unknown at the time Mendeleev published in his periodic table. He termed these elements Eka-Aluminium and Eka-Silicon because he left a space under aluminium and a gap under silicon.

Periodic Table of Elements
based on Mendeleev's Periodic Law

0	I	II	III	IV	V	VI	VII	VIII	VIII	VIII
	H 1.01									
He 4.00	Li 6.94	Be 9.01	B 10.8	C 12.0	N 14.0	O 16.0	F 19.0			
Ne 20.2	Na 23.0	Mg 24.3	Al 27.0	Si 28.1	P 31.0	S 32.1	Cl 35.5			
Ar 40.0	K 39.1	Ca 40.1 Zn 65.4	Sc 45.0 Ga 69.7	Ti 47.9 Ge 72.6	V 50.9 As 74.9	Cr 52.0 Se 79.0	Mn 54.9 Br 79.9	Fe 55.9	Co 58.9	Ni 58.7
Kr 83.8	Rb 85.5 Ag 108	Sr 87.6 Cd 112	Y 88.9 In 115	Zr 91.2 Sn 119	Nb 92.9 Sb 122	Mo 95.9 Te 128	Tc (99) I 127	Ru 101	Rh 103	Pd 106
Xe 131	Cs 133 Au 197	Ba 137 Hg 201	La 139 Tl 204	Hf 179 Pb 207	Ta 181 Bi 209	W 184 Po (210)	Re 180 At (210)	Os 194	Ir 192	Pt 195
Rn (222)	Fr (223)	Ra (226)	Ac (227)	Th 232	Pa (231)	U 238				

Dobereiner's triads · Known to Mendeleev · Lanthanide series · Actinide series · Known to Ancients

Mendeleev's Periodic Table published in 1905

Example: 1.1: Case Based:

Following the discovery of a significant number of elements, it became necessary to classify and organise them in a logical sequence based on their periodic properties. Johann Wolfgang Dobereiner attempted to group elements with similar properties in 1817. He discovered Dobereiner's triads, which are groups of three elements with comparable physical and chemical properties. John Newlands grouped all known elements in increasing atomic mass order in 1865 and discovered that the properties of every eighth element are comparable to the qualities of the first element.

(A) According to Newlands' law of octaves, the characteristics of magnesium are similar to those of:

(a) Beryllium (b) Lithium
(c) Sodium (d) Potassium

(B) A and B are two elements with similar qualities that follow Newlands' octave law. How many elements exist between A and B?

(a) 9 (b) 8
(c) 6 (d) 7

(C) Give an example of Dobereiner's triad.

(D) On what basis are the elements in Dobereiner's triad arranged?

(E) Assertion (A): Newlands' law of octave shows that Li and Na have the same chemical properties.

Reason (R): Newlands Law of octave depends on ascending order of atomic weight.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).
(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).

(c) (A) is true but (R) is false.
(d) (A) is false but (R) is true.

Ans. (A) *(a) Beryllium*

Explanation: According to Newlands' law of octaves, every eighth element possessed characteristics comparable to the first element. Thus, the characteristics of magnesium are similar to those of Beryllium.

(B) *(c) 6*

Explanation: According to Newlands' octave rule, every eighth element possessed characteristics comparable to the first element. Thus, 6 elements exist between A and B.

(C) Li, Na, K

Explanation: Each triad's middle element had an atomic weight that was almost midway between the other two. In addition, the middle element's characteristics were in the middle of the other two elements.

(D) Atomic mass

Explanation: Each triad's middle element had an atomic weight that was almost midway between the other two. In addition, the middle element's characteristics were in the middle of the other two elements.

(E) *(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).*

Explanation: Newlands in 1885 that, if the chemical elements are arranged according to increasing atomic weight, those with similar physical and chemical properties occur after each interval of seven elements.

Modern Periodic Law and the Present Form of the Periodic Table

Henry Moseley noticed patterns in the elements' characteristic spectra. The plot of $\sqrt{v}$ (where v is the frequency of X-rays emitted) versus the atomic number (Z) yielded a straight line, whereas the plot of $\sqrt{v}$ vs atomic mass yielded a curved line. As a result, he established that an element's atomic number is more elementary than its atomic mass. As a result, Mendeleev's Periodic Law was amended correspondingly.

Periodic Tabl e of Elements

- Alkali metals
- Alkaline-earth metals
- Transition metals
- Other metals
- Other non-metals
- Halogens
- Noble gases
- Rare-earth elements (21, 39, 57–71) and lanthanoid elements (57–71 only)
- Actinoid elements

period \ group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li												5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

Lanthanoid series 6	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinoid series 7	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Modern Periodic Table

Modern Periodic Law can be stated as:

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

In a neutral atom, the atomic number is equal to the nuclear charge (number of protons) or the number of electrons. The periodic variation in electronic configurations, which control the physical and chemical properties of elements and their compounds, is what periodic law is all about.

Periods are the horizontal rows (which Mendeleev named series) and groups are the vertical columns.

Important

➥ *Repetition of the similar properties of the elements placed in a certain group and separated by a certain definite gap of atomic number is known as periodicity.*

Groups or families are made up of elements that have similar external electronic configurations in their atoms and are grouped in vertical columns. The groups are numbered from 1-18 on the recommendation of the International Union of Pure and Applied Chemistry (IUPAC), replacing the earlier notation of groups.

IA–VIII, VIII, IB– VIIB and 0

Important

➥ *There are seven periods in total. The principal quantum number (n) of the elements corresponds to the period number. There are two elements in the first period. The number of elements in the subsequent periods are 8, 8, 18, 18 and 32 elements, respectively. The seventh period has a maximum possible number of elements (32) (based on quantum numbers).*

Example 1.2: What is the basic theme of organisation of elements in the periodic table? **[NCERT]**

Ans. The primary subject of element organisation in the periodic table is to simplify and systematise the study of the properties of all elements and millions of their compounds. Because the properties of components are now researched in groups rather than individually, the study has become simpler.

Example 1.3: **Case Based:**

The physical and chemical properties of elements are periodic functions of their atomic numbers, according to the modern periodic law.

The atomic number of a neutral atom equals the nuclear charge (number of protons) or the number of electrons. Periodic variation in electronic configurations, which determine the physical and chemical properties of elements and their compounds, is the focus of Periodic Law.

Periods are horizontal rows (called series by Mendeleev) while groups are vertical columns.

Groups or families are made up of elements in vertical columns that have comparable external electronic arrangements in their atoms.

(A) Who discovered the Modern periodic law?

(a) Mendeleev (b) Bohr
(c) Rutherford (d) Moseley

(B) Modern periodic law is based on:

(a) atomic mass (b) atomic numbers
(c) no. of porotons (d) no. of neutrons

(C) Define periods in a periodic table.

(D) Define groups in a periodic table.

(E) Assertion (A): Mendeleev's periodic Law is based on atomic number.

Reason (R): Periods are horizontal rows while groups are vertical columns.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).
(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).
(c) (A) is true but (R) is false.
(d) (A) is false but (R) is true.

Ans. **(A)** *(d) Moseley*

Explanation: Modern periodic law was discovered by Henry Moseley. He noticed patterns in the elements' characteristic spectra. The plot of $\sqrt{v}$ (where *v* is the frequency of X-rays emitted) versus the atomic number (*Z*) yielded a straight line, whereas the plot of $\sqrt{v}$ vs atomic mass yielded a curved line. As a result, he established that an element's atomic number is more elementary than its atomic mass. As a result, Mendeleev's Periodic Law was amended correspondingly.

(B) *(b) Atomic Numbers*

Explanation: According to the modern periodic law, the physical and chemical properties of elements are periodic functions of their atomic numbers.

(C) Periods are the horizontal rows of the periodic table. The periodic table has a total of seven periods.

(D) In the periodic table, a group is defined as a vertical column. A group is a series of elements with the same outermost electronic configurations in terms of atom electronic structure. In the long form of the periodic table, there are 18 groups.

(E) *(d) (A) is false but (R) is true.*

Explanation: Mendeleev's periodic law is based on atomic weight.

TOPIC 2

NOMENCLATURE OF ELEMENTS WITH ATOMIC NUMBERS >100

Until a new element's discovery is proven and its name is formally recognized, IUPAC recommends that a systematic nomenclature be derived directly from the element's atomic number using the numerical roots for 0 and digits 1 – 9

Digit	Root	Abbreviation
0	Nil	n
1	Un	u

Digit	Root	Abbreviation
2	Bi	b
3	Tri	t
4	Quad	q
5	Pent	p

Digit	Root	Abbreviation
6	Hex	h
7	Sept	s
8	Oct	o
9	Enn	e

Notation for IUPAC Nomenclature of Elements

Atomic Number	Name according to IUPAC Nomenclature	Symbol	IUPAC Official Name	IUPAC Symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassium	Hs

There have been discoveries of elements with atomic numbers up to 118 thus far. The roots are combined in the order of the atomic number's digits and "ium" is added at the end.

Electronic Configurations of Elements and the Periodic table

A combination of four quantum numbers characterizes an electron in an atom, with the principal quantum number (n) defining the major energy level known as the **shell**. The electronic configuration of an atom refers to the arrangement of electrons in different orbitals. The filling of electrons into a different subshell, also known as orbitals (s, p, d, f) in an atom. There is a direct link between the electronic configurations of the elements and the periodic table in its lengthy form. The quantum numbers of the last orbital filled determine an element's position in the periodic table.

Electronic Configurations in Periods

The value of n for the outermost or valence shell is shown by the period. The filling of the next higher main energy level ($n = 1$, $n = 2$, etc.) is related to each succeeding period in the periodic table.

Each period has twice as many elements as there are atomic orbitals accessible in the energy level being filled.

(1) The **first period** ($n = 1$) has two elements [H($1s^1$) and He ($1s^2$)]

(2) **Second period** ($n = 2$) involves the filling of $2s$ and $2p$ orbitals. Thus, there are eight elements ranging from lithium ([He]$2s^1$) to neon ([He]$2s^2$ $2p^6$).

(3) **Third period** ($n = 3$) involves the filling of $3s$ and $3p$ orbitals. Thus there are eight elements ranging from sodium ($z = 11$) to argon ($z = 18$).

(4) **The fourth period** ($n = 4$) involves the filling of $4s$ and $4p$ orbitals. In between the filling of $4s$ and $4p$, $3d$ orbitals are also filled. Thus, all nine orbitals are filled and therefore, there are eighteen elements in fourth period from potassium ($Z = 19$) to krypton ($Z = 36$).

(5) **The fifth period** ($n = 5$) involves the filling of $5s$ and $5p$ orbitals. In between the filling of $5s$ and $5p$, $4d$ orbitals are also filled. Thus, all nine orbitals are filled and therefore, there are eighteen elements in fifth period from rubidium ($Z = 37$) to xenon ($Z = 54$).

(6) **The sixth period** ($n = 6$) contains 32 elements, and consecutive electrons enter $6s$, $4f$, $5d$, and $6p$ orbitals. The filling of the $4f$ orbitals begins with cerium (Z=58) and finishes with lutetium (Z=71) resulting in the lanthanoid series.

(7) **The seventh period** (n=7) is characterized by the sequential filling of the $7s$, $5f$, $6d$, and $7p$ orbitals, and it contains the majority of man-made radioactive elements. The period terminates at the element with atomic number 118, which belongs to the noble gas family. Filling up the $5f$ orbitals after actinium ($Z = 89$) produces the actinoid series, which is a $5f$–inner transition series.

(8) To maintain the periodic table's structure and to maintain the principles of classification by

grouping elements with comparable properties in a single group, the 4f and 5f-inner transition series of elements are placed separately at the bottom of periodic table.

Groupwise Electronic Configurations

Electronic configurations of valence shells, number of electrons in outer orbitals, and characteristics are all comparable across elements in the same vertical column or group. This resemblance originates from the fact that the outermost orbitals of these elements have the same amount and distribution of electrons.

Depending on the sort of atomic orbitals that are being filled with electrons, elements are divided into 4 groups:

s-block, p-block, d-block and f-block.

Exceptions include:

(1) Helium is an s-block element, but its location in the p-block alongside another group of 18 elements is justified since it possesses a filled valence shell ($1s^2$) and hence shows characteristics similar to other noble gases.

(2) Because hydrogen only has one s-electron, it belongs to group 1 (alkali metals). It may also gain an electron to become a noble gas and so act like a group 17 (halogen family) element. Hydrogen is placed at the top of the periodic table separately because it's a special case.

Example 1.4: Write the atomic number of the element present in the third period and seventeenth group of the periodic table.

Ans. Chlorine. The atomic number is 17.

OBJECTIVE Type Questions

[1 mark]

Multiple Choice Questions

1. Who gave the first periodic classification?

(a) Dobereiner (b) Newland
(c) Moseley (d) Mendeleev

Ans. *(a) Dobereiner*

Explanation: Johann Dobereiner (1780-1849), a German chemist, classified various groups of three elements as triads in 1829. Lithium, sodium, and potassium were one such trinity. Triads were formed using both physical and chemical features.

2. The last element of the p-block in the 6^{th} period is represented by the outermost electronic configuration.

(a) $7s^2\ 7p^6$
(b) $5f^{14}\ 6d^{10}\ 7s^2\ 7p^0$
(c) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$
(d) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^4$ [NCERT Exemplar]

Ans. *(c) $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$*

Explanation: The last element of the p-block in the 6^{th} period is represented by $4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$.

3. The two exceptions to the classification of the periodic table into groups are:

(a) Hydrogen and Lithium
(b) Hydrogen and Helium
(c) Helium and Nitrogen
(d) Nitrogen and Sodium

Ans. *(b) Hydrogen and Helium*

Explanation: Helium is a s-block element, but its location in the p-block alongside the group of 18 elements is justified since it possesses a filled valence shell ($1s^2$) and hence shows characteristics similar to other noble gases.

Hydrogen only has one s-electron, it should belong to group 1 (alkali metals). It may also so act like a group 17 (halogen family) element which can gain an electron to become a noble gas and thus hydrogen is placed at the top of the periodic table.

4. In the long form of the periodic table, elements are arranged in the increasing order of:

(a) Atomic mass
(b) Atomic number
(c) Several nucleons
(d) Principal Quantum number

Ans. *(b) Atomic number*

Explanation: Modern periodic law can be stated as: the physical and chemical properties of the elements are periodic functions of their atomic numbers.

5. The number of elements in the sixth period is:

(a) 18 (b) 8
(c) 32 (d) 18

Ans. *(c) 32*

Explanation: The formula to find the number of elements in a given period is:

$$\text{if } n \text{ is even, number of elements} = \frac{(n+2)^2}{2}.$$

$$\text{if } n \text{ is odd, number of elements} = \frac{(n+1)^2}{2}$$

So. number of elements in 6^{th} period $= \frac{(6+2)^2}{2}$

$= \frac{64}{2} = 32$

where n is the period number.

Caution

➥ *n is not the principal quantum number in this case.*

6. The elements in which electrons are progressively filled in 4f-orbitals are called:
   (a) actinoids (b) transition elements
   (c) lanthanoids (d) halogens
   [NCERT Exemplar]

Ans. *(c) lanthanoids*

Explanation: Filling up of the 4f-orbitals begins with cerium (Z=58) and ends at lutetium (Z=71) 4f-inner transition series is called the lanthanoid series.

While actinoids are filled in the 5f-orbitals. Halogens are filled in group 17 whereas transition metals are mainly in the *d*-block.

7. Which of the following is not an actinoid?
   (a) Curium (Z = 96) (b) Californium (Z = 98)
   (c) Uranium (Z = 92) (d) Terbium (Z = 62)
   [NCERT Exemplar]

Ans. *(d) Terbium (Z=62)*

Explanation: Actinoids are elements with Z = 90 – 103. Therefore, Terbium (Z=65) is not an actinoid. Terbium (Z=65) is a lanthanoid, whose electronic configuration is $[Xe]4f^9 5d^0 6s^2$.

8. The period number in the long form of the periodic table is equal to:
   (a) magnetic quantum number of any element of the period.
   (b) atomic number of any element of the period.
   (c) maximum principal quantum number of any element of the period.
   (d) maximum Azimuthal quantum number of any element of the period.
   [NCERT Exemplar]

Ans. *(c) maximum principal quantum number of any element of the period.*

Explanation: Because each period begins with the filling of electrons in a new principal quantum number, the period number in the periodic table's long-form refers to the maximum principal quantum number of any element in the period.

The period number is equal to the maximum n of any element (where n is the principal quantum number).

9. Classification is necessary and significant because it helps:
   (a) in the systematic and easy study of the properties of elements
   (b) to know the type of different compounds that different elements can form
   (c) to correlate the properties of elements
   (d) all of the above [Diksha]

Ans. *(d) all of the above*

Explanation: Classification helps in the easy segregation of elements as well as the valency of the element can be determined which helps to know the type of compound which be formed.

Assertion-Reason (A-R)

In the following question no. (10-14) a statement of assertion followed by a statement reason is given Choose the correct answer out of the following choices.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).
(b) Both (A) and (R) are true but R is not the correct explanation of (A).
(c) (A) is true but (R) is false.
(d) (A) is false but (R) is true.

10. Assertion (A): Second period consists of 8 elements.
    Reason (R): Number of elements in each period is four times the number of atomic orbitals available in the energy level that is being filled.

Ans. *(c) (A) is true but (R) is false.*

Explanation: The number of elements in each period is twice the number of atomic orbitals available in the energy level that is being filled.

11. Assertion (A): According to Mendeléev, the properties of elements are a periodic function of their atomic masses.
    Reason (R): Atomic number is equal to the number of protons.

Ans. *(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).*

Explanation: Mendeleev arranged elements in horizontal rows and vertical columns of a table in order of their increasing atomic weights in such a way that the elements with similar properties occupied the same vertical column or group.

12. Assertion (A): The three elements in a triad have the same atomic mass gaps.
    Reason (R): The properties of the elements in a trio are

comparable.

Ans. *(d) (A) is false but (R) is true.*

Explanation: In a triad, the atomic mass of the middle element is equal to the mean of the atomic masses of the first and third elements.

13. Assertion (A): The element ununbium has an atomic number of 112.

Reason (R): In Latin terms, the names for numerals 1 and 2 are *un*-and *bi*-respectively.

Ans. *(a) Both (A) and (R) are true and (R) is the correct explanation of (A).*

Explanation: For 112, the root will be bi, and the symbol will be b. As a result, it will be known as ununbium, and its symbol will be *Uub*. Finally, we may deduce that the name and symbol for an element with the atomic number 112 are *ununbium* and *Uub*.

14. Assertion (A): Mendeleev's periodic law states that the properties of elements are a periodic function of their atomic numbers.

Reason (R): Mendeleev's periodic law could not account for the phenomenon of anomalous pairs.

Ans. *(d) (A) is false but (R) is true.*

Explanation: According to Mendeleev's periodic law, the atomic weight is the basis for all attributes. This means that the assumption that the properties of elements are a periodic function of their atomic numbers is inaccurate. It failed because they did not totally correspond to the sequence of atomic mass and were unable to identify hydrogen in the periodic table. The increase in atomic mass was discovered to be uneven when travelling from one element to another. They failed due to anomalous pairs, in which atoms are placed in the periodic table according to their atomic masses, with the element with the lowest atomic weight coming first. These are known as anomalous pairs, as they defy the rule.

CASE BASED Questions (CBQs)

[4 & 5 marks]

Read the following passages and answer the questions that follow:

15. *Elements in the contemporary periodic table are listed in increasing atomic number order, which is related to the electronic configuration. The elements in the periodic table have been separated into four blocks s, p, d and f based on the type of orbitals receiving the last electron, namely, the contemporary periodic table is divided into seven periods and eighteen groups. Each period starts with the formation of a new energy shell. According to the Aufbau concept, each of the seven periods (1 to 7) has 2, 8, 8, 18, 18, 32 and 32 elements. To keep the periodic table from becoming excessively long, the lanthanoids and actinoids series of f-block elements are placed towards the bottom of the main body of the periodic table.*

(A) How many elements are there in the f-block?

(B) Which concept is used to determine the number of elements in each period?

(C) What is the difference between Lanthanoids and Actinoids in terms of position in the periodic table?

Ans. (A) Aufbau concept is used to determine the number of elements, each of the seven periods has elements.

(B) There is a distinction between Lanthanides and Actinides. The filling of 4f-orbitals is done in lanthanoids, while the filling of 5f-orbitals is done in actinoids.

(C) *(c) d-block*

Explanation: The position of an element in the periodic table is determined by the quantum numbers of the last orbital filled. Lanthanum has the atomic number 57. As the valence shell of lanthanum is 5d, lanthanum belongs to the *d*-block of the periodic table.

16. *The modern table is based on Mendeleev's table, except the modern table arranges the elements by increasing atomic number instead of atomic mass. Atomic number is the number of protons in an atom, and this number is unique for each element. The modern table has more elements than Mendeleev's table because many elements have been discovered since Mendeleev's time.*

Rows of the modern periodic table are called periods, as they are in Mendeleev's table. From left to right across a period, each element has one more proton than the element before it. Some periods in the modern periodic table are longer than others. Columns of the modern table are called groups, as they are in Mendeleev's table. However, the modern table *has many more groups—18 compared with just 8 in Mendeleev's table. Elements in the same group have similar properties. Periodic table is divided into 4 blocks viz. s-block, p-block, d-block and f-block.*

(A) Which of the following is a f-block element?
  (a) Sodium (b) Argon
  (c) Samarium (d) Zinc

(B) s-block elements are:
  (a) Groups I-A and III-B elements
  (b) Groups I-A and II-A elements
  (c) Groups I-A and I-B elements
  (d) Groups I-A and II-B elements

(C) The number of groups and periods respectively in modern periodic table is:
  (a) Seven, eighteen
  (b) Eighteen, seven
  (c) Seven, seven
  (d) Eighteen, eighteen

(D) The element with atomic number 57-belongs to:
  (a) s-block (b) *p*-block
  (c) *d*-block (d) *f*-block

(E) Assertion (A): Helium is placed in group 18 along with *p*-block elements.

Reason (R): It shows properties similar to *p*-block elements.

  (a) Both (A) and (R) are true and (R) is the correct explanation of (A).
  (b) Both (A) and (R) are true but (R) is not the correct explanation of (A).
  (c) (A) is true but (R) is false.
  (d) (A) is false but (R) is true.

Ans. (A) *(c) Samarium*

Explanation: Samarium is a f-block element with the atomic number 62 and the periodic table symbol Sm.

(B) *(b) Groups I-A and II-A elements*

Explanation: s-block elements are elements from groups I-A (*alkali metals*) and II-A (*alkaline earth metals*).

(C) *(b) Eighteen, seven*

Explanation: There are seven periods and eighteen groups in the modern periodic table.

(D) The elements of the f-block are listed individually at the bottom of the periodic table because they resemble each other but do not resemble any other group elements.

(D) *(c) (A) is true but (R) is false.*

Explanation: He ($1s^2$) should be placed along with s-block elements because of its electronic configuration but it has a completely filled valence shell and as a result, it exhibits properties of noble gases, thus it is placed along with noble gases (ns^2, np^6).

(E) There are 28 elements belonging to the f-block element and have been grouped in two horizontal rows

VERY SHORT ANSWER Type Questions (VSA)

[1 mark]

17. Define periodicity.

Ans. The recurrence of similar properties of elements after definite intervals when elements are arranged in increasing order of their atomic numbers.

18. What is modern periodic law? Define it.

Ans. According to modern periodic law, "The physical and chemical properties of elements are periodic functions of their atomic numbers".

19. Who is responsible for the development of the modern periodic table?

Ans. Henry Moseley and Dmitri Mendeleev.

20. Write the electronic configuration of Cr(Z=24). Justify your answer. [Diksha]

Ans. The expected configuration of Cr is [Ar]$3d^4\ 4s^2$ but the actual electronic configuration of Cr ($Z = 24$) is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$/[Ar]$3d^5 4s^1$.

This configuration is exceptional as half-filled orbitals have extra stability.

21. Identify the group and valency of the element having an atomic number 119. Also, predict the outermost electronic configuration and write the general formula of its oxide. [NCERT Exemplar]

Ans. The group is one and the valency is also one. The outermost electronic configuration is $8s^1$ and the formula of oxide is M_2O.

22. Why is Hydrogen considered an exceptional element in the periodic table?

Ans. Because hydrogen only has one s-electron, it belongs to group 1 (alkali metals). It may also gain an electron to become a noble gas and so act like a group 17 (halogen family) element. Thus, hydrogen is considered as an exceptional element and is placed at the top of the periodic table.

SHORT ANSWER Type-I Questions (SA-I)

[2 marks]

23. An element has an outer shell electronic configuration $4s^2\,4p^3$. Find:

(a) The atomic number of the element placed next below it.

(b) The atomic number of the next noble gas.

[Delhi Gov. QB 2022]

Ans. (a) Arsenic with atomic number has an electronic configuration $[Ar]\,3d^{10}\,4s^2\,4p^3$

Then the atomic number of the element placed next below it will be 51 which is Sb.

(b) Noble gas after arsenic is krypton which has atomic number 36.

24. What did Dobereiner's law of triads state? What are the limitations of the triads?

Ans. Each triad's middle element had an atomic weight that was almost midway between the other two. In addition, the middle element's characteristics were in the middle of the other two elements.

Limitation: Because it appeared to work only for a few elements, it was dismissed as a fluke.

25. (A) What is the basic difference in approach between Mendeleev's Periodic Law and the Modern Periodic Law?

(B) Why do elements in the same group have similar physical and chemical properties?

Ans. (A) The basic difference in approach between Mendeleev's Periodic Law and Modern Periodic Law is the change in the basis of the classification of elements from atomic weight to atomic number.

(B) The elements in a group have the same valence shell electronic configuration and hence have similar physical and chemical properties.

Related Theory

➥ *Elements in the same group have similar chemical properties and can be classified as metals, metalloids, and non-metals, or as main group elements, transition elements, and inner transition elements.*

SHORT ANSWER Type-II Questions (SA-II)

[3 marks]

26. Determine the group and valency of the element with the atomic number 119. Predict the outermost electronic configuration as well as the overall formula for its oxide.

Ans. The current configuration of the periodic table's long form can hold a maximum of 118 elements. Following that, according to the Aufbau principle, the 8s-orbital should be filled. As a result, the outer electronic configuration of element 119 with atomic number 119 is $[Og]\,8s^1$. Because it possesses one electron in the outermost s-orbital, its valency is 1, and it should be classified as a group 1 metal, along with alkali metals. Its oxide will have the general formula M_2O, where M is the element.

1st group, Valency: 1

Outermost electronic configuration = $8s^1$

Oxide formula = M_2O

LONG ANSWER Type Questions (LA)

[4 & 5 marks]

27. What are the general characteristics of the modern periodic table's lengthy form?

Ans. The following are some of the general properties of the long version of the periodic table:

(1) Groups are 18 vertical columns.

(2) From left to right, their groupings are numbered 1-18.

(3) The term "periods" refers to seven horizontal rows.

(4) The primary group components are those found in groups 1, 2 and 13-17.

(5) Transition elements are the elements found in groups 3-12.

(6) The elements are ordered in ascending order of atomic number.

(7) The inert gases, such as He, Ne, Ar and so on.

(8) Except for the 0 and VIII groups, all of the groups in the contemporary periodic table are separated into subgroups A and B.

Related Theory

➥ *Lanthanoids and actinoids have been given a separate place at the bottom of the periodic table due to their analogous behaviour.*

28. Explain the characteristics of each of the seven periods.

Ans. The first period, with only two elements hydrogen and helium, is thought to be the shortest.

Both the second and third periods include eight elements, ranging from lithium to neon in the second and sodium to argon in the third.

Both the fourth and fifth periods include 18 elements, ranging from potassium to krypton in the fourth and rubidium to xenon in the fifth.

The sixth period, which begins with caesium and ends with radon, has 32 elements and is claimed to be the longest.

The seventh period also contains 32 elements, it begins with francium, which has an atomic number of 87 and continues until atomic number 118(Og), which are wholly synthetic and are called transuranium elements.

29. Write the drawbacks in Mendeleev's periodic table that led to its modification.

[NCERT Exemplar]

Ans. Mendeleev's periodic table has the following drawbacks:

(1) **Hydrogen's position:** Hydrogen location is not fixed. It does, however, resemble the elements of the group 1 (alkali metals) as well as the elements of the group 17 (halogens). As a result, the position of hydrogen in the periodic table is incorrect.

(2) **Anomalous pairs:** The ascending order of atomic masses were not followed in certain pairs of elements. Mendeleev arranged the elements in these circumstances based on similarities in their properties rather than the increasing order of their atomic weights. Argon (Ar) is placed before potassium, for example. Likewise, cobalt comes before nickel while tellurium (Te) comes before iodine. These positions were not supported by evidence.

(3) **Isotope position:** Isotopes are atoms of the same element with differing atomic weights but the same atomic number. As a result, according to Mendeleev's classification, these should be classified differently based on their atomic masses. For example, hydrogen isotopes with different atomic weights should be stored in three different locations. On the other hand, Isotopes, do not have their spot in the periodic table.

PROPERTIES OF ELEMENTS AND THEIR PERIODIC TRENDS

2

TOPIC 1

PERIODIC TRENDS IN PROPERTIES OF ELEMENTS

Periodic Trends in Physical Properties

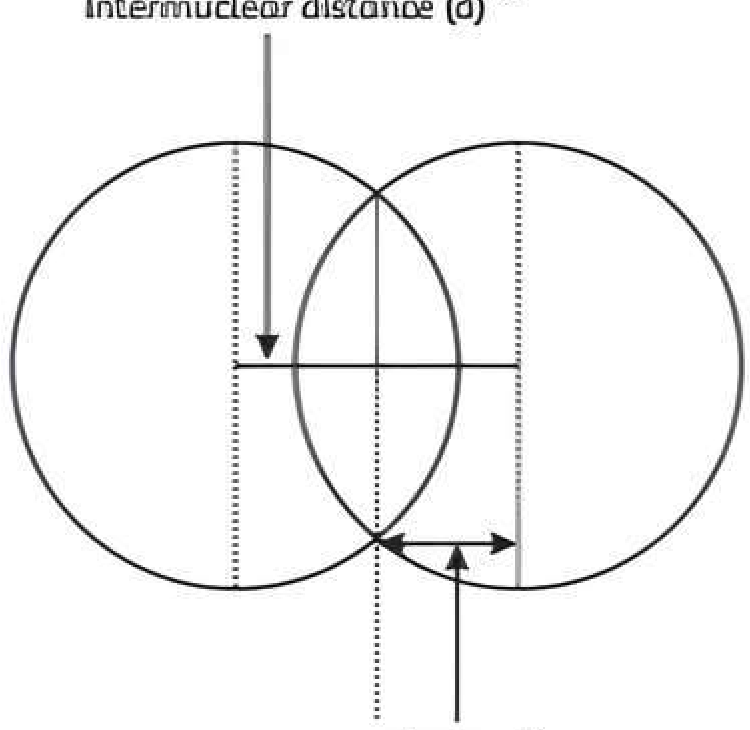
We observe a common trend in physical and chemical properties as we move across a period from left to right or down the group. This trend in properties is known as periodic properties. Some important periodic properties like atomic size, metallic character, non-metallic character, ionisation potential, electron affinity and electronegativity are discussed here:

Atomic radius

As we know atoms are tiny spherical bodies so they have atomic radii. Atomic radius in uncomplicated words can be defined as the distance between the centre of its nucleus to the outermost shell. It is difficult to determine the exact radius of an atom. Thus, atomic radius is generally estimated by knowing the distance between the atoms in the combined form. The atomic radius is of the following types:

(1) Covalent Radius: A covalent radius is a radius between the elements that form a covalent bond. A covalent bond is the bond-forming by two non-metallic elements. Half of the internuclear distance between the nuclei of the two bonded atoms of the same element in a molecule is referred to as covalent radius.

$$\text{Covalent radius} = \frac{1}{2} \times d$$

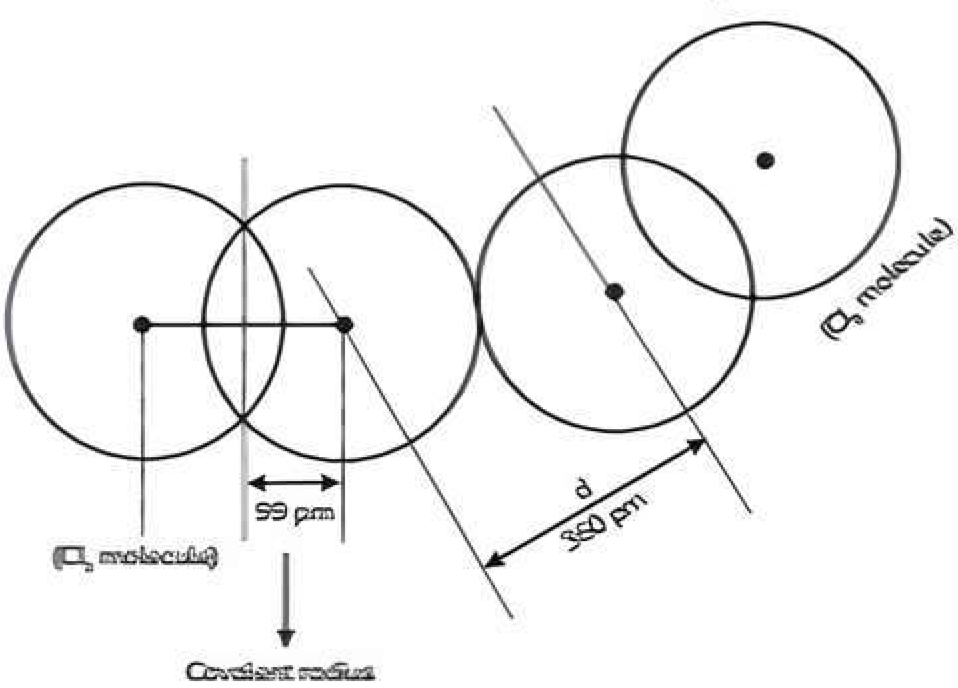


Covalent radius

(2) Metallic radius: The atomic radius of a metal is known as metallic radius. It is defined as half of the distance between two consecutive nuclei of adjacent metal atoms in solid state.

For example, the metallic radius of iron is 126 pm as the separation between two adjoining iron atoms in solid iron is 252 pm.

(3) van der Waals radius: Van der Waals radius is described as the half of the distance between the nuclei of two non-bonded neighbouring atoms of adjoining molecules in a stable state. Here we can see in the diagram, the chlorine atom's van der Waals radius is 180 pm.



Chlorine's van der Waals radius

Where, d is the internuclear distance between the two non-bonded neighbouring atoms

$$\text{Van der Waals radius} = \frac{1}{2} \times 360 \text{ pm} = 180 \text{ pm}$$

Important

➡ *The decreasing order of atomic radius between Covalent radius, Metallic radius and van der Waals radius is as follows: Van der Waals radius > Metallic radius > Covalent radius*

➡ *We see exceptional behaviour in the noble gas. The atomic radii of inert gases steadily increase in comparison to the halogen group, because in noble gases, atomic radii refer to the van der Waals radius, but in other elements, it relates to the covalent radius.*

For calculating atomic radius, both covalent and metallic radius are considered depending on whether the element is a non-metal or metal. Below are a few elements with their atomic radii. These trends could be explained on the basis of nuclear charge and energy level]

Table: Atomic radii/(pm) across the periods

Atom (Period II)	Atomic Radius	Atom (Period III)	Atomic Radius
Li	152	Na	186
Be	111	Mg	160
B	88	Al	143
C	77	Si	117
N	74	P	110
O	66	S	104
F	64	Cl	99

Table: Atomic radii (pm) down the groups

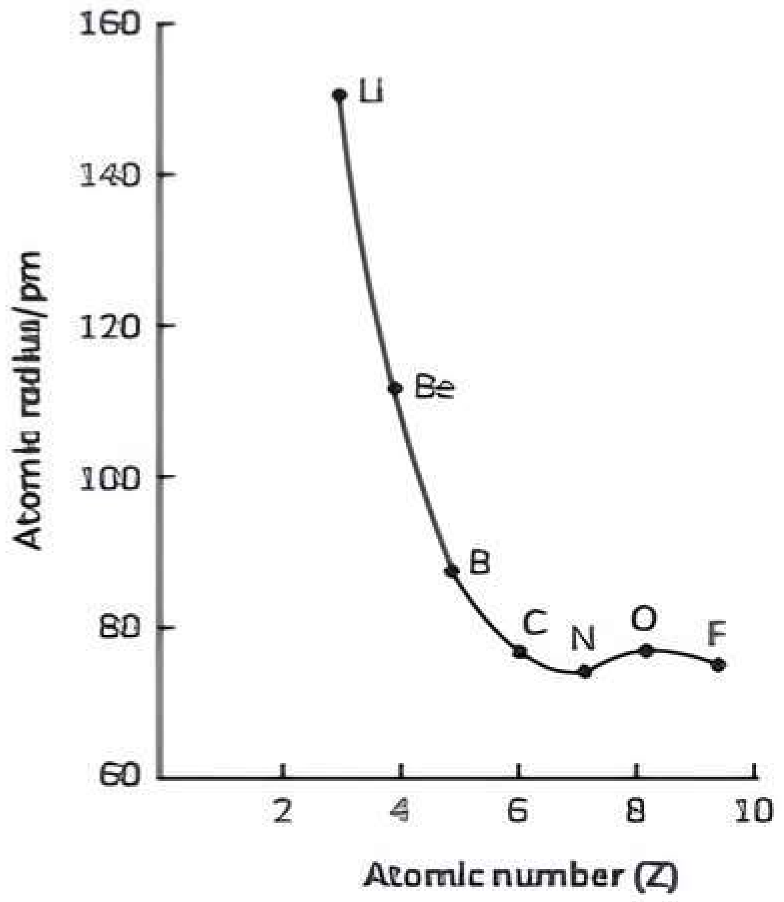
Atom (Group I)	Atomic Radius	Atom (Group 17)	Atomic Radius
Li	152	F	64
Na	186	Cl	99
K	231	Br	114
Rb	244	I	133
Cs	262	At	140

From the above tables, we can notice the trends in atomic radii across the period and down the group, the atomic size normally reduces across the period while there is a rise in atomic size down the group.

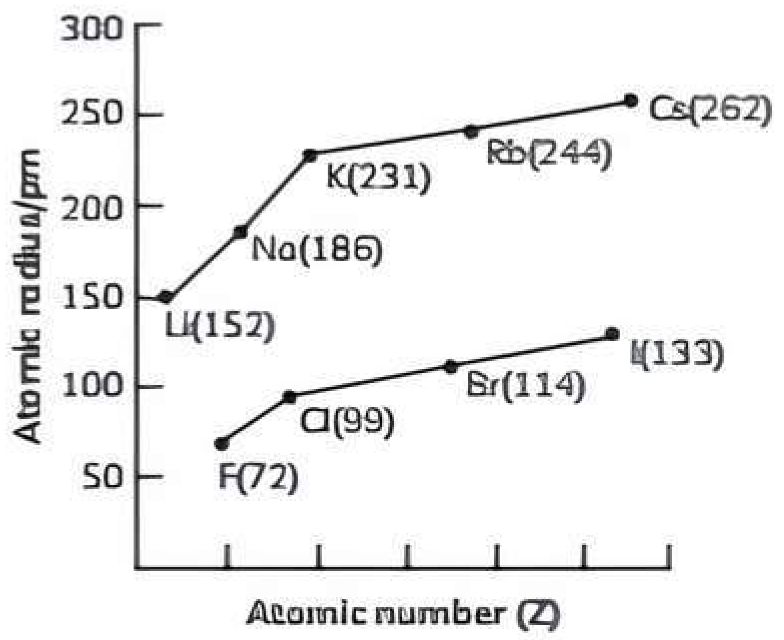
Across the period the atomic radii decrease due to an increase in the the effective nuclear charge. This is because across the period, the addition of electrons takes place in the same valence shell that gradually increases the force of attraction of the electrons to the nucleus. However, we see that down the group, there is an increase in the atomic radius. In a group, there is a constant increase in the principal quantum number (n) due to the addition of a new shell. By adding a new shell, the distance between the valence electrons and the nucleus increases which results in a rise in the atomic size.

Important

➥ *The atomic radii of noble gases are not considered as their non-bonded radii are very large. For comparing the radii of noble gases with other elements, one should consider van der Waals radii of other elements and not their covalent radii.*

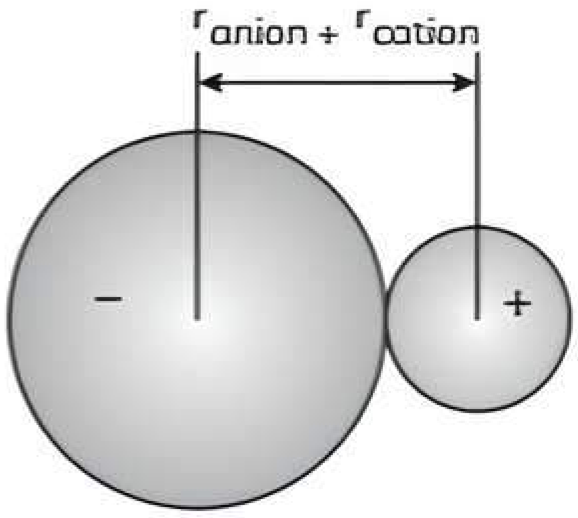


Variation of atomic radii with atomic number across the second period



Variation of atomic radii with an atomic number for alkali metals and halogens

(4) Ionic radius: Ionic radius is the distance from the nucleus of an ion up to which it has an influence on its electron cloud. Ions are formed when an atom loses or gains electrons. When an atom loses an electron it forms a cation and when it gains an electron it becomes an anion. The ionic radius can be described as the distance between the nucleus of an ion and the outermost shell of the ion. The atomic size of a cation will be smaller than that of the parent atom. This is because as it loose electrons, it's effective nuclear charge increases. An anion is relatively larger in size than its parent atom. This is because when an atom gains electrons the total number of electrons increases which tends to create more repulsion between electrons and thus overshadows the net effective nuclear charge.



Ionic Radius

Isoelectronic species

An isoelectronic series is a series of elements and ions that have equal numbers of electrons. In a series of isoelectronic species, as the nuclear charge increases, the force of attraction by the nucleus on the electrons also increases. As a result of this, ionic radius decreases. Higher the magnitude of its negative charge the more will be its ionic radii and higher the magnitude of its positive charge the smaller will be its ionic radii.

Example 2.2: Which of the following species will have the largest and the smallest size? Mg, Mg^{2+}, Al^{3+}. [NCERT]

Ans. The atomic radii decreases across the period. Cations are relatively smaller than their parent atoms in size. Thus, Mg^{2+} is smaller than Mg. Among all the isoelectronic species, the atom with the largest positive nuclear charge will have the smallest radius. Thus, Al^{3+} is smaller than Mg^{2+}. Hence, in the above question, the largest species is Mg and the smallest is Al^{3+}.

Related Theory

➥ *Isoelectronic species are known as atoms or ions that have the same number of electrons. In the case of such species as the magnitude of nuclear charge increases the ionic radius decreases.*

Ionisation Energy

The amount of energy required to remove the most loosely bound electron from an isolated gaseous atom is referred to as Ionisation enthalpy. The first ionisation enthalpy change for the reaction can be shown as

$$A_{(g)} + IE \rightarrow A^{+}_{(g)} + e^{-}$$

Similarly, the second and third ionisation enthalpy is the energy required to remove the second and third most loosely bound electron respectively. It's SI unit is kJ/mol and always has a positive value. The ionisation energy of a multivalent atom increases with a consecutive removal of an electron. This is due to an increase in effective nuclear charge on the valence electron that makes it difficult to ionize a cation. Therefore, $IE_1 < IE_2 < IE_3$

Factors governing the ionisation energy

(1) Effective nuclear charge
(2) Atomic size
(3) Penetration effect of electrons
(4) Screening effect of inner electrons
(5) Electronic configuration

(1) Effective nuclear charge

It is the net nuclear attraction towards the valence shell electrons. The greater is the effective nuclear charge, the more tightly the outer electrons are held by the nucleus and thus more will be the ionisation enthalpy.

(2) Atomic size

As we know the atomic radii vary inversely to the effective nuclear charge. Thus, ionisation enthalpy will also vary inversely to atomic radii. Therefore, the smaller is the atomic radii of an element, the greater will be its ionisation enthalpy.

(3) Penetration Effect

Penetration means the proximity of an electron in an orbital to the nucleus. For each shell and subshell, it can be observed as the relative density of electrons near the nucleus of an atom. Now if we look into the radial probability distribution functions, we see that the electron density of *s*-orbitals is closer than that of *p*-and *d*-orbitals.The order of penetration power will be $2s > 2p > 3s > 3p > 4s > 3d$. Thus, the more is the penetration effect, the more will be the ionisation enthalpy.

(4) Screening effect of inner electrons

The electrons in the inner shells operate as a screen or shield between the nucleus and the electrons in the outermost shell. This phenomenon is referred to as the shielding or screening effect. Due to this, the outermost electrons experience a low effective nuclear charge and not the actual nuclear charge. The effective nuclear charge can be given as:

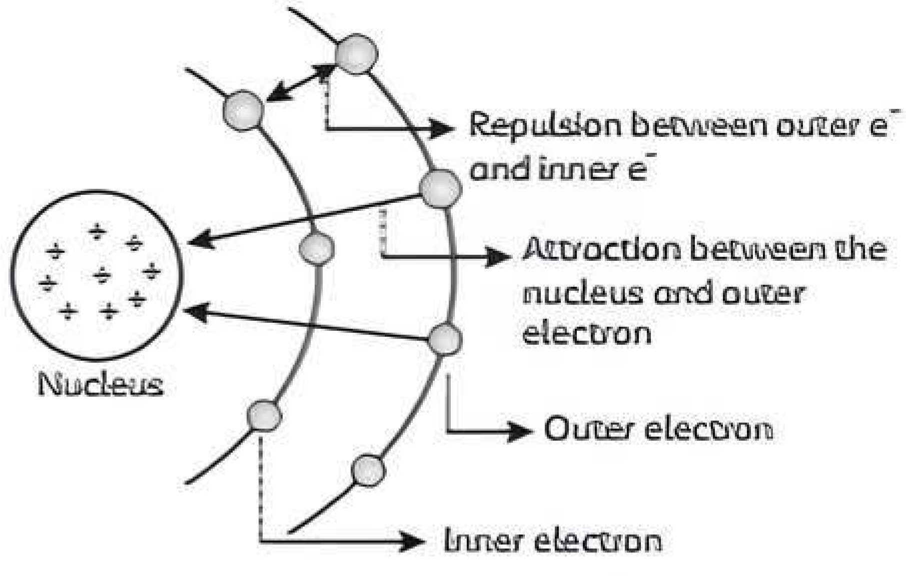
$Z_{effective} = Z - \sigma$

$Z_{effective}$ = effective nuclear charge

Z = actual nuclear charge

σ = screening constant

Thus, the larger is the number of electrons within the inner shells, the larger will be the screening effect and the smaller is the force of attraction and thus, the lesser will be the ionisation enthalpy.



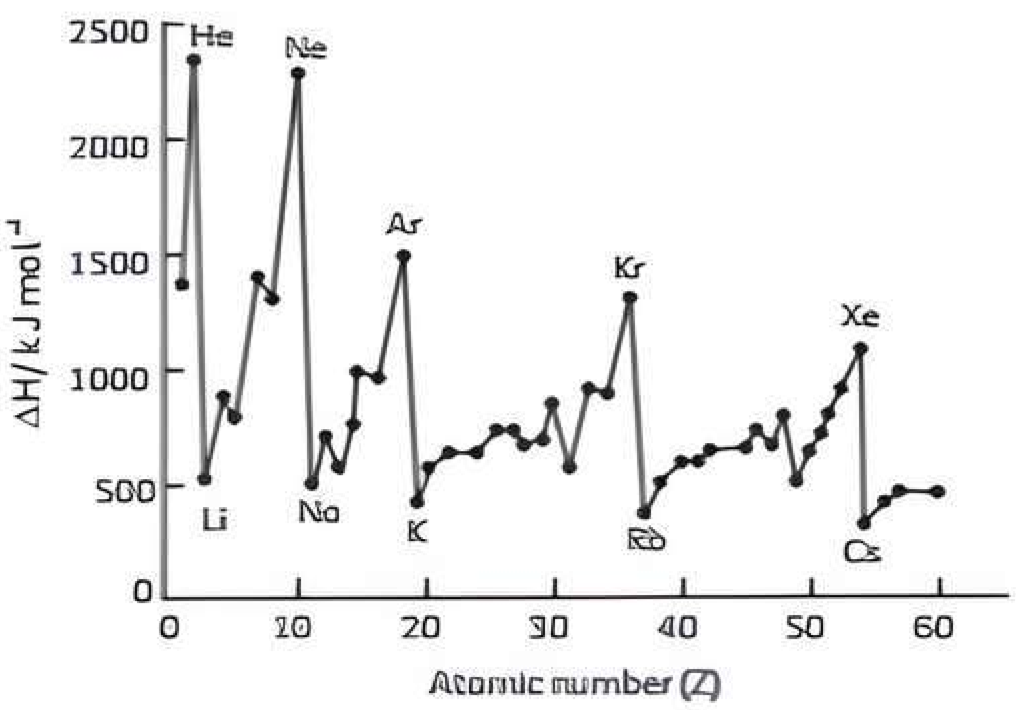
Shielding or Screening effect

(5) Electronic configuration

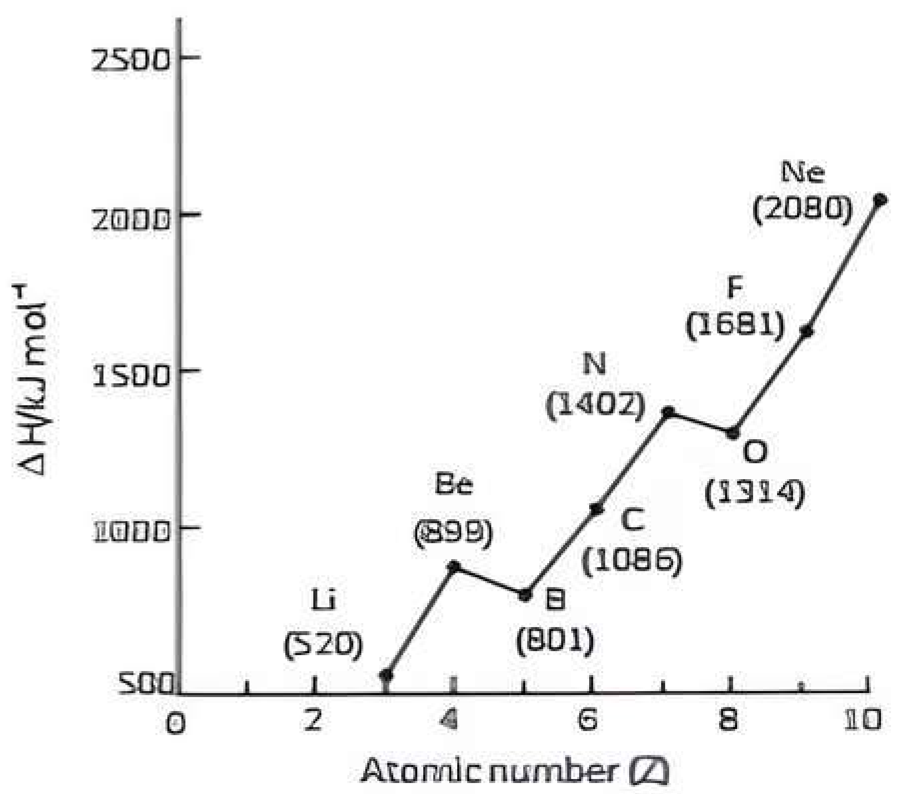
Elements having half-filled and fully-filled orbitals are stable. So if we will try to remove an electron from these orbitals, then it will make them less stable. Hence, more amount of energy is required to remove an electron from these orbitals. Thus, these elements will have higher ionisation energy.

Variation of Ionisation enthalpy in the Periodic Table

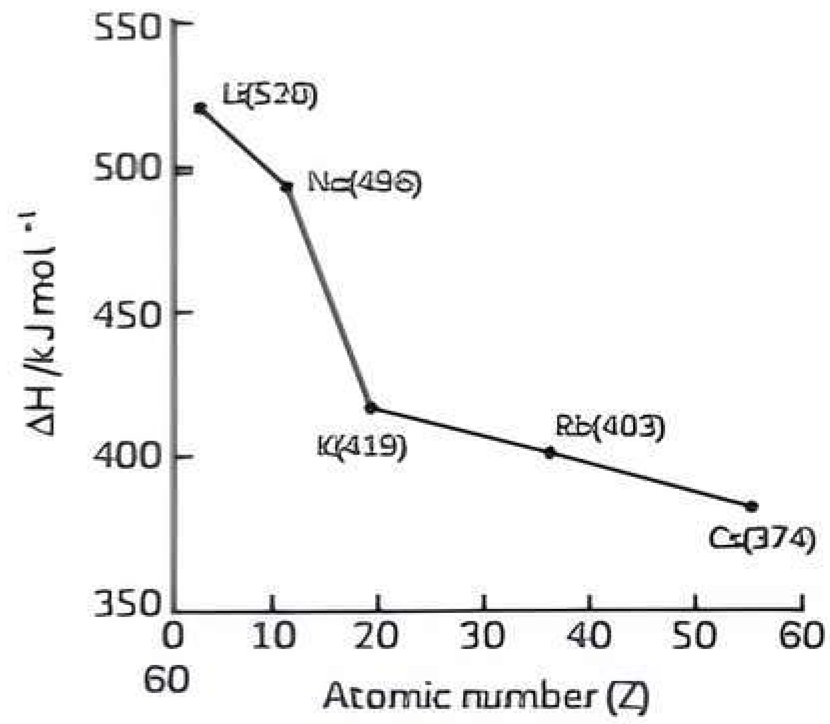
Ionisation enthalpy depends upon the shielding effect, nuclear charge, penetration effect and electronic configuration. On moving left to right in a period, the ionisation enthalpy increases as the effective nuclear charge increases. Thus in a period, noble gases have maximum ionisation enthalpy. However, on moving down the group, it's value decreases. This is because the number of shells increases down the group. The outermost electrons will be far away from the nucleus, and the effective nuclear charge will be less.



Variation of first ionization enthalpies



First ionisation enthalpies of elements of the second period as a function of atomic number



First ionisation enthalpy of alkali metals as a function of atomic number.

Important

- *Beryllium has more IE_1 than Boron as it has a fully filled s-orbital, as more energy is required to remove an electron from half or fully-filled orbitals. The ionisation enthalpy of the inert gases is unusually higher because of the stable configuration.*

The first ionisation enthalpy IE_1 of N (nitrogen) is greater than that of O (oxygen) because of stable half-filled 2*p*-orbitals in nitrogen.

^{24}Cr and ^{29}Cu have higher second ionisation, IE_2 than other transition elements of the 3*d*-series because of the half-filled and fully-filled stable *d*-orbitals respectively.

Electron gain enthalpy (EGE or ΔH_{eg})

Electron Gain Enthalpy is the enthalpy change accompanying a process when an electron is added to a neutral isolated gaseous atom to form the negative ion. During the addition of an electron, energy can either be released or absorbed. Its SI unit is kJ/mol. It can be positive (*i.e.*, endothermic process) or negative (*i.e.*, exothermic process) depending upon the nature of an element. For example, halogens have a strong tendency to accept electrons in order to attain noble gas configuration, so they have maximum negative electron gain enthalpy (because the energy is being released in the process) in their respective periods of the periodic table, whereas noble gases have large positive electron gain enthalpies because the electron has to enter the next higher principal quantum level leading to a very unstable electronic configuration.

Successive Electron Gain Enthalpies

The magnitude of the first electron gain enthalpy is generally negative. However, the second and further electron gain enthalpies are generally positive. For example:

$O_{(g)} + e^- \rightarrow O^-_{(g)}$ $\quad \Delta H_{eg} = -141 \text{ kJ mol}^{-1}$

$O^-_{(g)} + e^- \rightarrow O^{2-}_{(g)}$ $\quad \Delta H_{eg} = 780 \text{ kJ mol}^{-1}$

Factors governing electron gain enthalpy and variation in periodic table

(1) Atomic size

(2) Nuclear charge

(3) Electronic configuration

(1) Atomic size

As the size of the atom increases, the distance between the nucleus and the last shell which receives the incoming electrons increases. This decreases the force of attraction between the nucleus and the incoming electron. Hence, the electron gain enthalpy becomes less negative.

(2) Nuclear charge

As the nuclear charge increases, the force of attraction between the nucleus and the incoming electron increases. Hence, the enthalpy becomes more negative.

(3) Electronic configuration

Elements with exactly half-filled or completely filled orbitals are very stable. You have to supply energy to add an electron. Hence, their electron gain enthalpy has large positive values.

Variation of Electron Gain Enthalpy in the periodic table

As we move across a period from left to right the atomic size decreases and the nuclear charge increases. Both these factors tend to increase the attraction by the nucleus for the incoming electron. Hence, electron gain enthalpy becomes more and more negative in a period from left to right. As we move down a group, both the atomic size and the nuclear charge increases. But the effect of the increase in atomic size is much more pronounced than the nuclear charge. With the increase in atomic size, the attraction of the nucleus for the incoming electron decreases. Hence, the electron gain enthalpy becomes less negative. We can measure electron gain enthalpy by the Born-Haber cycle.

Those which have excess ΔH_{eg} are good oxidizing agents.

Group 1	$\Delta_{eg}H$	Group 16	$\Delta_{eg}H$	Group 17	$\Delta_{eg}H$	Group 18	$\Delta_{eg}H$
H	– 73					He	+ 48
Li	– 60	O	– 141	F	– 328	Ne	+ 116
Na	– 53	S	– 200	Cl	– 349	Ar	+ 96
K	– 48	Se	– 195	Br	– 325	Kr	+ 96
Rb	– 47	Te	– 190	I	– 295	Xe	+ 77
Cs	– 46	Po	– 174	At	– 270	Rn	– 68

Election Gain Enthalpies (kJ mol^{-1})

Important

➥ *The electron gain enthalpies of some of the elements of 2nd period i.e. N, O and F are less negative than the corresponding elements of the third period i.e. P, S and Cl. This is because the elements of the second period have the smallest atomic size among the elements in their respective group. As a result, there are considerable electron-electron repulsions within the atom itself and hence the additional electron is not accepted with the same ease as is the case with the remaining elements in the same group. Thus, chlorine has the maximum negative electron gain enthalpy in the periodic table.*

Example 2.3: Which of the following will have the most negative electron gain enthalpy and which will have the least negative? Explain your answer. [NCERT]

Ans. When we proceed in a period from left to right, the more negative the electron gain enthalpy becomes and the less negative, it becomes as we proceed down the group. When in the 2p-orbital, an electron enters causes greater repulsion as compared to when an electron enters to the 3p-orbitals. As a result, chlorine has the largest negative electron gain enthalpy, while phosphorus has the lowest.

Electronegativity

The tendency of an atom in a molecule to attract the shared pair of electrons towards itself is known as electronegativity. It is not a measurable quantity like ionisation enthalpy and electron gain enthalpy, etc. Although various numerical scales of electronegativity, such as the Pauling, Mulliken-Jaffe, and Allred Rochow scales, have been devised. Among all scales, the Pauling scale is the most broadly used. In 1922, an American scientist Linus Pauling assigned arbitrarily a value of 4.0 to fluorine, the elements are taken into consideration to have the finest tendency to draw the shared pair of electrons. Approximate values for the electronegativity of some elements are given beneath within the table.

Table: Electronegativity values on the Pauling scale

Element	Z	Pauling Electronegativity
C	6	2.55
N	7	3.04
O	8	3.44
Br	35	2.96
Pt	78	2.28
Au	79	2.54
Tl	81	1.62

The order of some elements based on Electronegativity (on the Pauling scale) is:

$$\underset{(4.0)}{F} > \underset{(3.5)}{O} > \underset{(3.0)}{N} = \underset{(3.0)}{Cl} > \underset{(2.8)}{Br}$$

The electronegativity of any given element is not constant but depends on the following factors:

(1) Atomic size: Electronegativity is inversely proportional to atomic size *i.e.*, as the atomic size increases the electronegativity decreases.

(2) State of hybridisation: *sp*-hybridised carbon is more electronegative than sp^2-hybridised carbon, which in turn is more electronegative than sp^3-hybridised carbon.

Electronegativity of carbon atom:

$$C_2H_6 < C_2H_4 < C_2H_2$$

(3) Oxidation state: The electronegativity increases as the oxidation state of an atom increases.

(4) Effective nuclear charge: Electronegativity is directly proportional to effective nuclear charge *i.e.*, as the effective nuclear charge increases, electronegativity of an atom also increases.

Variation of Electronegativity in the Periodic Table

As we move across a period from left to right the nuclear charge increases and the atomic size decreases, therefore the value of electronegativity increases across a period in the modern periodic table. For example, the electronegativity trend across period 3 in the periodic table is depicted below.

Atom (Period II)	Li	Be	B	C	N	O	F
Electro-negativity	1.0	1.5	2.0	2.5	3.0	3.5	4.0
Atom (Period III)	Na	Mg	Al	Si	P	S	Cl
Electro-negativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0

Electronegativity Values (on Pauling scale) Across the Periods

The value of electronegativity decreases as we move down the group because there is an increase in the atomic number as we move down the group in the modern periodic table. The nuclear charge also increases but the effect of the increase in nuclear charge is overcome by the addition of one shell.

Atom (Group 1)	Electro-negativity value	Atom (Group 17)	Electro-negativity value
Li	1.0	F	4.0
Na	0.9	Cl	3.0
K	0.8	Br	2.8
Rb	0.8	I	2.5
Cs	0.7	At	2.2

Electronegativity value (Pauling scale) down the group

It is a general observation that metals show a lower value of electronegativity as compared to the non-metals. Therefore, metals are electropositive and non-metals are electronegative in nature. The elements in period two differ in properties from their respective group elements due to the small size and higher value of electronegativity.

Example 2.4: Case Based:

Groups 13 (i.e., group IIIA)- 17 (i.e., group VIIA) of the periodic table, as well as elements from the group 18, i.e., the zero group, make up the p-block of the periodic table. The furthest p-orbital is occupied by the final electron in the p-block elements. They have 3 to 8 electrons in the outermost shell. Because the number of p-orbitals is three, the maximum number of electrons that may be compelled in a p-orbital arrangement is six. As a result, the periodic table has six groupings of p-block elements. The p-block contains all three categories of elements: metals, non-metals, and metalloids. The crossing line in the p-block separates the metals from the non-metals in the p-block. Non-metals are on the right of the line, whereas metals are on the left. The metalloids are discovered along the way. The p-block exhibits considerable variability in attributes due to the proximity of a wide spectrum of components. Except for He, the valence shell electrical design of p-block elements is ns^2np^{1-6}.

(A) The electronegativity of O, F and N should be in the sequence of:

(a) N > O > F (b) O > F > N
(c) O > N > F (d) F > O > N

(B) Match Column I to Column II and use the supplied codes to pick the right answer.

	Column I (Atoms)		Column II (Properties)
(A)	He	(i)	Highest electronegativity
(B)	F	(ii)	Most electropositive
(C)	Rb	(iii)	Strongest reducing agent
(D)	Li	(iv)	Highest ionisation energy

Select the correct option:

	(A)	(B)	(C)	(D)
(a)	(iv)	(ii)	(iii)	(i)
(b)	(i)	(iv)	(ii)	(iii)
(c)	(iv)	(i)	(iii)	(ii)
(d)	(iv)	(i)	(ii)	(iii)

(C) What is the increasing order of electronegativity of Si, P, C and N elements?

(D) Which element has the highest Ionisation energy in the halogen group?

(E) Assertion (A): Generally, ionisation enthalpy increases from left to right in a period.

Reason (R): Electron enters in a new shell when moving from left to right in a period.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).

(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).

(c) (A) is true but (R) is false.

(d) (A) is false but (R) is true.

Ans. (A) *(d) F > O > N*

Explanation: Moving from left to right in a period the electronegativity increases. As a result, the electronegativity of O, F and N should be in the following order: F>O>N

(B) *(d) (A)-(iv); (B)-(i); (C)-(ii); (D)-(iii)*

Explanation: The appropriate match is:

Helium (He) $1s^2$	Due to its noble gas nature and tiny size, it has the highest ionisation energy.
Fluorine (F) $1s^2 2s^2 2p^5$	Because of its tiny size and oxidation state of −1. It has the highest electronegativity.
Rubidium (Rb) $[Kr] 5s^1$	Because of its massive atomic size, it is the most electropositive element.
Lithium (Li) $1s^2 2s^1$	Because of its tiny size and positive oxidation state (+1). It is the most powerful reducing agent.

(C) Going across the period electronegativity increases and going down the group electronegativity decreases in general. As a result, the following is the progressive order of electronegativity:

$$Si < P < C < N$$

(D) Fluorine, as it is the smallest element in the group.

(E) *(c) (A) is true but (R) is false.*

Explanation: Electron enters in a new shell when moving down the group.

Periodic Trends in Chemical Properties

The elements within the periodic table display chemical properties with noticeable periodicity. Here we will talk about the periodicity of the valence state indicated by elements, and anomalous behaviour of the elements of the second period (from lithium to fluorine).

Here are the few properties which cause periodic variation in the periodic table:

Valency

The combining ability of the element is referred to as valency. It is related to the electronic configuration of an atom of the element and is usually decided by electrons present in the valence shell.

Valency increases from 1 to 4 before declining to zero (for noble gases), while moving along a period from left to right and valency remains constant when moving down a group. Transition metals show off variable valency due to the fact that they can use electrons from the outer in addition to the penultimate shell.

Valence is the maximum characteristic property of an element and may be understood in the words of several electrons existing within the outermost shell of the atom. The valence of illustrative elements is normally the same as the number of electrons within the outermost orbitals (s and p-block elements) or the same to eight minus the number of outermost electrons as expressed below.

Table: Periodicity of Valence or Oxidation States.

Group	Number of valence electron	Valency
1	1	1
2	2	2
13	3	3
14	4	4

15	5	3.5
16	6	2.6
17	7	1.7
18	8	0.8

Variation along a period

Noticing the number of valence electrons from left to right in a period, it will increase from 1 to 8. But the valency of an element first increases from one to four and then decreases to zero. These trends found within the valence of the element in hydrides and oxides are given beneath within the table.

Table: Periodic trends in the valency of elements as shown by the formulas of their compounds.

Group	Formula of hydride	Formula of oxide
1	LiH, NaH, KH	Li_2O, Na_2O, K_2O
2	CaH_2	MgO, CaO, SrO, BaO
13	B_2H_6, AlH_3	B_2O_3, Al_2O_3, Ga_2O_3, In_2O_3
14	CH_4, SiH_4, GeH_4, SnH_4	CO_2, SiO_2, GeO_2, SnO_2, PbO_2
15	NH_3, PH_3, AsH_3, SbH_3	N_2O_3, N_2O_5, P_4O_6, P_4O_{10}, As_2O_3, As_2O_5, Sb_2O_3, Sb_2O_5, BiO_3
16	H_2O, H_2S, H_2Se, H_2Te	SO_3, SeO_3, TeO_3
17	HF, HCl, HBr, HI	Cl_2O_7

In the 3rd period, elements show variable valences. The gradation found with respect to oxygen which grows up to seven. Nowadays the term oxidation state is more regularly used for valence. To represent the idea of oxidation states, consider two oxygen-containing compounds, OF_2 and K_2O. The three elements (K, O and F) involved within bonding in these compounds have electronegativity in the order of, F > O > K. The electrical configuration of F is $1s^22s^22p^5$, while O's is $1s^22s^22p^4$. Therefore, within the formation of OF_2, the molecule, each of the two atoms makes contributions to one electron in sharing for the formation of a bond with the oxygen atom at the same time as the O atom shares two electrons with F atoms. F being the maximum electronegative element is given an oxidation state of −1 at the same time as O being less electronegative than F is given an oxidation state of +2. In K_2O, oxygen is more electronegative and accepts two electrons from which have the electronic configuration $1s^22s^22p^63s^23p^64s^1$, consequently displaying an oxidation state of −2. K on the other hand, gains an oxidation state of +1 by losing one electron from the $4s^1$ to oxygen. Thus, the charge obtained by an element's atom based on electronegative consideration from other atoms within the molecule can be characterized as its oxidation state in an individual compound.

Variation within a group

When we move down the group, the number of valence electrons stays identical, consequently, all of the elements in a group show off the identical valence.

Anomalous Properties of Second Period Elements

Each group's starting elements, such as group 1 (lithium), group 2 (beryllium), and group 13-17 (boron to fluorine) are different in many respects from the rest of the group members. Although all the members of alkali metals and alkaline earth metals form ionic compounds, lithium and beryllium form covalent compounds. In fact, the properties of lithium and beryllium are more similar to the second element of the following group, i.e., magnesium and aluminium respectively. This similarity is commonly referred to as diagonal relationship. Also, the maximum covalency exhibited by the first members of each group is 4 whereas the other members of each group can expand their valence shell to show a covalency of 6.

The main reason for the different chemical behaviour of the first member of a particular group of elements in the s- and p-blocks compared to the other members of the same group is due to its:

(i) small size, (ii) large charge/ radius ratio, (iii) high electronegativity and (iv) absence of d-orbitals.

Diagonal relationship

II	Li	Be	B	C	N	O	F
III	Na	Mg	Al	Si	P	S	Cl

Diagonal relationship between the elements of the second and third periods

Periodic Trends and Chemical Reactivity

So far looking at the periodic trends in certain essential properties such as atomic/ionic radii, ionisation enthalpy, electron gain enthalpy, and valence, determined that periodicity in those properties and their electronic configuration is going on hand in hand. This means that the chemical and physical properties of elements are exceptionally dependent on their electronic configuration. Trends of some chemical properties are as follows:

Electronegativity and non-metallic character

Reactivity of non-metals will increase with the increase in electronegativity in addition to electron gain enthalpy and reduction in atomic radii. Reactivity of metal will rise with the reduction in IE, electronegativity, and rise in atomic size. Thus the halogens are highly reactive non-metals and their reactivity decreases on moving down the group.

Electropositivity or metallic character

It is the tendency of an element to develop a positive ion by losing valence electrons. The higher the electropositive character, the higher is metallic character. On moving across the period, electropositive character decreases while on moving down the group, it increases.

In their respective periods, halogens are the least electropositive elements, while alkali metals are among, the most electropositive elements.

Example 2.5: **Show by a chemical reaction with water that Na_2O is a basic oxide and Cl_2O_7 is an acidic oxide.** **[NCERT]**

Ans. When Na_2O reacts with water, it generates a strong base, whereas Cl_2O_7 forms a strong acid when it reacts with water.

$$Na_2O + H_2O \rightarrow 2NaOH$$

$$Cl_2O_7 + H_2O \rightarrow 2HClO_4$$

We can test their acidic and basic nature by using litmus paper.

Example 2.6: **Case Based:**

The electron gain enthalpy is the change in enthalpy that happens when in its ground state, the addition of an electron to a gaseous atom transmutes it to a negative ion ($\Delta_{eg}H$). It's a measurement of how easily an atom can attract an electron to form an anion.

$$X_{(g)} + e^- \rightarrow X^-_{(g)};\ \Delta H = \Delta_{eg} H$$

An atom's ground state is its most stable state. The reason being that when the isolated gaseous atom is in the excited state, a lesser amount of energy is released when it gets converted into gaseous anion after accepting an electron. As a result, in the definition of electron gain enthalpy, the phrases ground state, and isolated gaseous atom have been included. Similar to ionisation enthalpy, electron gain enthalpy is expressed in electron volts per atom or per mole.

(A) **Noble gases consist of a positive electron gain enthalpy as a result of:**
- (a) stable configuration
- (b) large size
- (c) high reactivity
- (d) unstable configuration

(B) **O and F have a lower electron gain enthalpy than S and Cl. It's because of:**
- (a) small size
- (b) less repulsion
- (c) large size
- (d) high electronegativity

(C) **Fluorine, chlorine, bromine, and iodine have different electron gain enthalpies (in kJ/mol). Write them.**

(D) **Electron gain enthalpy of Mg is positive. Explain.**

(E) **Assertion (A): The enthalpy of electron gain is influenced by atomic radii.**

Reason (R): The distance between two atoms when they are bound together by a single bond in a covalent molecule is known as covalent radii.

- **(a) Both (A) and (R) are true and (R) is the correct explanation of (A).**
- **(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).**
- **(c) (A) is true but (R) is false.**
- **(d) (A) is false but (R) is true.**

Ans. (A) *(a) stable configuration*

Explanation: Noble gases do not accept an electron because of their stable structure, and hence have a positive electron gain enthalpy.

(B) *(a) small size*

Explanation: When the atom's size is small, there is higher repulsion for the approaching electron.

When an electron is introduced to O or F, it moves to a lower (n=2) level and experiences greater repulsion than an electron added to S or Cl which moves to a larger (n=3) level and experiences less repulsion.

(C) Electron gain enthalpies of fluorine, chlorine, bromine, and iodine are −333, −349, −325 and −296 kJ/mol respectively.

The enthalpy of electron gain ($\Delta_{eg}H$) is the enthalpy change for transforming 1 mol of isolated atoms to anions by the addition of electrons. $\Delta_{eg}H$ is negative for all halogens (exothermic). In general, $\Delta_{eg}H$ becomes less negative, when compared to other components of the same group from top to bottom.

However, there is a difference in fluorine from others. Because of its small size, Fluorine has a greater repulsion for extra added electrons as compared to other atoms. Thus, $\Delta_{eg}H$ isn't as exothermic as expected for the F-atom. As a result, the proper values of electron gain enthalpies:

$$\underset{kJ\,mol^{-1}\ (-333)}{F} < \underset{(-349)}{Cl} > \underset{(-325)}{Br} > \underset{(-296)}{I}$$

(D) Metals like sodium, potassium, magnesium etc. lose electrons to obtain the inert gas configuration. Hence, Magnesium atoms will not add electrons easily. Some external energy is needed to add the electrons in their atoms. Hence, electron gain enthalpy for metals will be positive.

(E) *(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).*

Explanation: The enthalpy of electron gain is inversely proportional to atomic size, i.e., with the increase in atomic size value of electron gain enthalpy decreases.

OBJECTIVE Type Questions

[1 mark]

Multiple Choice Questions

1. **Which of the following groups of two elements show a diagonal relationship?**
   (a) Silicon and Boron
   (b) Aluminium and Boron
   (c) Gallium and Boron
   (d) Carbon and Boron

Ans. *(a) Silicon and Boron*

Explanation:

(a) Within the periodic table, Silicon and Boron are placed diagonally. They have similar properties. They both have high melting and boiling point and are non-metallic elements. Both are non-conductor of electricity and are semiconductors. Both subsist in amorphous and have a crystalline structure.

(b) Aluminium and Boron are in the same group. They do not have a diagonal relationship. Even though it has the same properties, it is due to the reason that they are in the same group.

(c) Gallium and Boron belong to the same family. There is no diagonal link between them. Even though they have the same properties, they are in the same group.

(d) Carbon and Boron are in the same period. They do not have any similarities and do not show a diagonal relationship.

Related Theory

- *The diagonal relationship exists between certain elements in the periodic table. These elements are placed diagonally adjacent in the second and third row of the periodic table, among the first twenty elements.*
- *The diagonal elements usually show similarities in their properties, which is exhibited on moving from left to right and down the group in the periodic table.*

2. **First ionisation enthalpies of Na, Mg, Al and Si are in the order:**
   (a) Na<Mg>Al<Si
   (b) Na>Mg>Al>Si
   (c) Na<Mg<Al<Si
   (d) Na>Mg>Al<Si **[NCERT Exemplar]**

Ans. *(a) Na<Mg>Al<Si*

Explanation: As ongoing across a period from left to right, the nuclear charge exceeds the shielding. As a consequence, the ionisation enthalpy will rise and the outermost electrons are held tightly throughout a period. Mg's outermost electron is in the 3s-orbital, which is a stable gas configuration, and the subshell has a greater penetrating impact than Al's outermost electron, which is in the 3*p* subshell.

3. **Which of the following has the largest size?**
   (a) Al (b) Al^{+}
   (c) Al^{2+} (d) Al^{3+} **[Diksha]**

Ans. *(a) Al*

Explanation: The atomic radii of cations are less than that of an atom as the electron is lost by the atom. The more electrons are lost, the more the size of the atom decreases.

(a) Al has the largest size among all as it is the parent atom.

(b) Al^{+} size is less than the parent atom as it loses one electron and forms a cation.

(c) Al^{2+} is even smaller than the parent atom as it lost two electrons.

(d) Al^{3+} is the smallest among all as it lost three electrons.

4. **What is the correct order of radii?**
   (a) $I^{+} < I < I^{-}$
   (b) Na > Mg > Al
   (c) $Na^{+} > Mg^{2+} > Al^{3+}$
   (d) All of the above

Ans. *(d) All of the above*

Explanation: We know that on moving down the group, the atomic radii increase, at the same time on moving along a period from left to right, atomic radii reduce as the effective nuclear charge increases. Moreover, the atomic radii of cation are less than that of an atom as the electron is lost by the atom while the atomic radii of anions are more than that of an atom as the electron gains by the atom. Thus, all the options given above are correct.

(1) As we know that anion's size is greater than the parent atom so I^{-} is the largest among all while the cation is smaller than the parent atom thus, I^{+} has a smaller size among all.

(2) As we know when we move from left to right the size of the atom decrease so here Na has the largest size and Al has the smallest.

(3) In iso-electronic species as the positive charge increases, size of cation decreases. Hence, atomic size order is $Na^{+} > Mg^{2+} > Al^{3+}$.

5. The correct order of electron gain enthalpies of group 17 is:

(a) F<Cl<Br<I (b) F>Cl>Br>I

(c) F<Cl
I (d) F<Cl>Br>I

Ans. *(d) F<Cl>Br>I*

Explanation: More negative is the electron gain enthalpy as we proceed down the group. In the case of fluorine, the repulsion between electrons is greater than in chlorine due to its small size. As a result, chlorine has the largest negative electron gain enthalpy.

6. The formation of the oxide ion, $O^{2-}(g)$, from oxygen atom requires first an exothermic and then an endothermic step as shown below:

$O_{(g)} + e^- \rightarrow O^-_{(g)}$ $[\Delta H_{eg} = -141 kJ\ mol^{-1}]$

Thus, the process of formation of O^{2-} in gas phase is unfavourable even though O^{2-} is isoelectronic with neon. It is because:

(a) oxygen is more electronegative.

(b) addition of electrons in oxygen results in a larger size of the ion.

(c) electron repulsion outweighs the stability gained by achieving noble gas configuration.

(d) O^- ion has a comparatively smaller size than an oxygen atom. [NCERT Exemplar]

Ans. *(c) electron repulsion outweighs the stability gained by achieving noble gas configuration.*

Explanation: When one electron is added to an O atom the energy is released to produce an O^- ion. As a result, the first electron gain enthalpy of oxygen is negative. But when one more electron is added to O^-, it forms O^{2-}. To overcome the strong electronic repulsions, the energy has to be provided. As a result, the enthalpy of O's second electron gain is positive.

Assertion-Reason (A-R)

In the following question no. (7-10), a statement of assertion followed by a statement of reason is given Choose the correct answer out of the follows choices.

(a) Both (A) and (R) are true and (R) is the correct explanation of (A).

(b) Both (A) and (R) are true but (R) is not the correct explanation of (A).

(c) (A) is true but (R) is false.

(d) (A) is false but (R) is true.

7. Assertion (A): Isoelectronic species consist of different radii.

Reason (R): There are a different number of electrons in isoelectronic species.

Ans. *(c) (A) is true but (R) is false.*

Explanation: Isoelectronic species do not have the same radii as they have a dissimilar number of protons and neutrons. Moreover, these are the type of ions or atoms which contain the same number of electrons. For example, Mg^{2+}, O^{2-}, Ne, etc.

8. Assertion (A): Electron gain enthalpy becomes less negative as we go down a group.

Reason (R): The size of the atom increases on going down the group and the added electron would be farther from the nucleus.

[NCERT Exemplar]

Ans. *(a) Both (A) and (R) are true and (R) is the correct explanation of (A).*

Explanation: Because the size of the atom increases as the electron is added farther away from the nucleus of the atom, on moving down the group, the screening effect increases resulting in making electron gain enthalpy less negative.

9. Assertion (A): Boron has a smaller first ionisation enthalpy than beryllium.

Reason (R): The penetration of a 2s electron to the nucleus is more than the 2*p* electron hence 2*p* electron is more shielded by the inner core of electrons than the 2s electrons. [NCERT Exemplar]

Ans. *(a) Both (A) and (R) are true and (R) is the correct explanation of (A).*

Explanation: During ionisation, the electron removed in case of beryllium is from the s-orbital and the electron removed from the boron atom is from the p-orbital, and the penetration of 2*s* electron to the nucleus is more than that of 2*p* electron hence, 2*p* electron of boron is more shielded from the nucleus than the 2*s* electron.

10. Assertion (A): Generally, ionisation enthalpy increases from left to right in a period.

Reason (R): When successive electrons are added to the orbitals in the same principal quantum level, the shielding effect of the inner core of electrons does not increase very much to compensate for the increased attraction of the electron to the nucleus.

[NCERT Exemplar]

Ans. (a) *Both (A) and (R) are true and (R) is the correct explanation of (A).*

Explanation: The minimum amount of energy required for the removal of the loosely bound electron from the isolated gaseous atom is referred to as Ionisation enthalpy. Here because of the rise in the attraction of the nucleus resulting in a rising in ionisation enthalpy while moving along the period from left to right which is accurately explained in the reason. So, it is the correct explanation of the given assertion.

CASE BASED Questions (CBQs)

[**4** & **5** marks]

Read the following passages and answer the questions that follow:

11. *The radius of the isoelectronic species might be dissimilar due to their dissimilar nuclear charges. As already described the size of the cation is usually smaller than its parent atom whereas the size of the anion is usually larger than its parent. The successful lack of electrons from an atom raises the effective nuclear charge whereas the successful achievement of electrons reduces the effective nuclear charge. This is the cause of the cation with an extra positive charge having a smaller radius due to the extranuclear attraction of the electrons. While the anion with the extra negative charge has a bigger radius because, in this situation, The electrons' net repulsion surpasses the nuclear charge, causing the size to expand.*

(A) Which of the following species has the biggest size?

N, N^{3-}, F and F^-

(B) Which of the following species has the smallest size?

N, N^{3-}, F and F^-

(C) Give a reason for the answers of (A) and (B).

Ans. (A) The biggest ion is N^{3-}.

(B) The smallest species is F.

(C) N^{3-} and F^- are anions that indicate that their size is greater than their parent atoms. As we know more the electron gains the size of the ion compared to its parent atom so N^{3-} has the largest size whereas F has the smallest size as its effective nuclear charge is greater than N.

12. *Though elements are categorized into s, p, d and f-block, it is possible to classify them further into metals, non-metals, and metalloids. More than 78 per cent of all recognized elements are metals. These elements are found on the periodic table's left side and in the middle. These elements are malleable and ductile, own lustre, have high densities, and are good conductors of heat and electricity. Metals typically have excessive melting points, and boiling points, and most are solid at room temperature. Non-metals are much fewer in number than metals. There are about 17 non-metals in total. These elements are found on the periodic table's upper right-hand side. Their melting points and boiling points are generally low. At room temperature at non-metals typically are solids or gases. They aren't ductile or malleable in the least. They are poor heat and electrical conductors. The non-metallic property rises from left to right in a period. Metallic character rises down the group.*

(A) Name some metalloids.

(B) Arrange the following items in ascending metallic character order:

Si, Be, Mg, Na, P

The atomic number and position in the periodic table are taken into consideration.

(C) Give a reason for the order obtained in (B).

Ans. (A) Silicon, germanium, arsenic, antimony and tellurium.

(B) P < Si < Be < Mg < Na

(C) We know that on moving down the group metallic character rises whereas on moving along a period from left to right it reduces, thus, the ascending order of the metallic character is: P < Si < Be < Mg < Na.

13. *The quantity of energy liberated when an electron is added to a remote gaseous atom is referred to as electron gain enthalpy. The fundamental electron affinity is usually exothermic. Because it takes energy to add an electron to an anion and overcome the force of repulsion, the second electron* *affinity is endothermic. In general, in a group, electron gain enthalpy reduces as the size rises and, in a period, electron gain enthalpy rises as the atomic radius reduces. However, there are numerous exceptions because of the stable electronic configurations and due to the smaller size of atoms.*

(A) Which of the halogens has the highest electron affinity?

(a) Cl (b) F

(c) Br (d) I

(B) Choose the chalcogen with the maximum negative electron gain enthalpy.

(a) S (b) Se

(c) O (d) Po

(C) Element which has more negative electron gain enthalpy is:

(a) F (b) O

(c) Cl (d) S

[Delhi Gov. QB 2022]

(D) Which of the following statements is incorrect?

(a) Helium has the highest first ionisation enthalpy in the periodic table.

(b) Chlorine has less negative electron gain enthalpy than fluorine.

(c) Mercury and bromine are liquids at room temperature.

(d) In any period, atomic radius of alkali metal is the highest.

Ans. (A) *(a) Cl*

Explanation: Among the species, the highest negative electron gain enthalpy is of Cl the atom. The reason for this is that as you move down the group, the electron gain enthalpy gets somewhat negative. But the introduction of an electron to the 2*p* orbital creates extra repulsion than the introduction of an electron to the 3*p* orbital.

(B) *(c) O*

Explanation: Oxygen has the maximum negative electron gain enthalpy. Because of the smaller atomic size.

(C) *(c) Cl*

Explanation: Chlorine has the highest negative electron gain enthalpy in the periodic table.

(D) *(b) Chlorine has less negative electron gain enthalpy than fluorine*

Explanation: Chlorine has more negative electron gain enthalpy than fluorine. Due to the very small size of fluorine, it's not able to accommodate the newly coming electron and hence has lower negative electron gain enthalpy than fluorine.

14. *The minimal quantity of energy required to dispose of the outermost electron from a remoted gaseous atom is called the ionisation enthalpy (IE) of the element. From top to bottom, the ionisation energy of a group decreases, whereas the ionisation energy of a period increases from left to right. However, there are numerous exceptions because of stable electronic configurations. The energy required to dispose of the 2^{nd} electron from the monovalent cation is referred to as second ionisation enthalpy (IE_2). Similarly, we've 3^{rd}, 4^{th}... ionisation energy.*

(A) Which one would have the highest difference in their first and second ionisation enthalpies?

Na, Mg, Si ***and*** **P**

Explain.

(B) Why solar panels are made up of silicon?

(C) The ionisation enthalpy of Nitrogen is higher than the Oxygen atom. Explain.

Ans. (A) Na (Sodium). Sodium has the greatest difference as after the removal of an electron, gives stable noble gas configuration.

Electronic configuration of Na = $1s^2 2s^2 2p^6 3s^1$

Thus, to remove the second electron from sodium ion, more energy is required which is now harder as it has a stable inert gas configuration. Thus, sodium has a higher difference in its first and second ionisation enthalpy.

(B) Mostly Silicon is utilized in solar panels as a semiconductor because it is a cost-efficient material that offers good energy efficiency. Furthermore, silicon has a higher corrosion resistance, a longer lifespan, ideal thermal expansion properties, excellent photoconductivity, and low toxicity.

(C) Nitrogen (*Z*=7) has anomalously higher first ionisation enthalpy (1402kJ/mol) than oxygen (Z=8), (1314kJ/mol) This arises due to the fact that half-filled orbitals and full-filled degenerate orbitals are extra solid than the incompletely filled degenerate orbitals. Here nitrogen($1s^2 2s^2 2p^3$) has half-filled *p*-orbitals while oxygen ($1s^2 2s^2 2p^4$) has incompletely stuffed.

VERY SHORT ANSWER Type Questions (VSA)

[1 mark]

15. Make a note regarding the positive enthalpy of electron gain.

Ans. The positive electron gain entholpy is defined as the energy required by an isolated gaseous atom to accept an electron into an atom. The elements having a completely filled orbital do not accept electrons which makes them difficult. Thus, by applying some heat energy they tend to accept the electron.

16. Electron gain enthalpy of F is less negative than that of Cl. Explain. [Diksha]

Ans. As F is of smaller size than that of Cl resulting in greater interelectronic repulsions is correspondingly small in $2p$ orbitals of F. So, the arriving electron does not encounter much attraction. Thus, F has less negative electron gain enthalpy.

17. Beryllium has excessive ionisation enthalpy than Boron. Why?

Ans. This is because $Be(1s^2 2s^2)$ has fully filled s-orbital. The energy required to remove an electron will be more due to presence of stable configuration whereas, in the case of $B(1s^2 2s^2 2p^1)$, there is no such stable fully filled configuration. Thus, it has a stable configuration.

18. Give a reason why the location of inert gases is on the extreme right of the periodic table.

Ans. Inert gases are those elements that have a stable electronic configuration which indicates that they are less reactive than other elements in their respective periods. Furthermore, their electrical configuration is also completely filled.

19. Arrange the elements N, P, O and S in the order of:

(A) increasing first ionisation enthalpy.

(B) increasing non-metallic character.

Give a reason for the arrangement assigned. [NCERT Exemplar]

Ans. (A) S < P < O < N

The ionisation enthalpy increases on proceeding from left to right along the period. But N has higher ionisation enthalpy than O, due to stable half-filled configuratition.

(B) P < S < N <O

The non-metallic nature increases as goes from left to right in the period.

SHORT ANSWER Type-I Questions (SA-I)

[2 marks]

20. First member of each group of representative elements (*i.e.*, *s* and *p-block* elements) shows anomalous behaviour. Illustrate with two examples. [NCERT Exemplar]

Ans. The enormous charge/radius ratio, small size, strong electronegativity, and non-availability of the *d-orbitals* in the valence shell all contribute to second-period elements' unusual behaviour. In comparison to subsequent members of the group, the ability to create p_π–p_π multiple bonds to itself (*e.g.*, C = O, C = C, O = O, N = N) is greatly demonstrated by the first member of each group of *p-block* elements and other by second-period elements (*e.g.*, C = O, C = N, N = O) which requires +91.8kJ mol^{-1} of energy.

21. All transition elements are *d-block* elements, but all elements do not transition elements. Explain. [NCERT Exemplar]

Ans. The few elements of *d*-orbital are referred to as transition elements as they show transitional behaviour between s-block and *p*-block elements. Their properties are intermediate between metallic elements of s-block which contain ionic nature and the non-metallic elements of *p*-blocks which consist of covalent nature. But elements like Zn, Cd, Hg do not show transitional character as they consist of fully-filled *d*-orbitals.

22. Write four characteristic properties of *p*-block elements. [NCERT Exemplar]

Ans. (1) *p*-block consists of metals, non-metals, and metalloids.

(2) In most cases, elements make covalent bonds with other elements.

(3) Their ionisation enthalpy is higher than that of s-blocks elements.

(4) Some of them have more than one oxidation number.

23. Write a short note on s-block elements.

Ans. The elements discovered in groups 1 and 2 of the periodic table are known as s-block elements. Elements that have one electron in their valence shell are the alkali metals. They are reactive elements as their ionisation energies are low. Elements that have two electrons in their valence shell are the alkaline earth metals.

24. Nitrogen has positive electron gain enthalpy whereas oxygen has negative. However, oxygen has lower ionisation enthalpy than nitrogen. Explain. [NCERT Exemplar]

Ans. The outermost electronic configuration of nitrogen $2s^2 2p^3$ is very stable because *p*-orbitals is half filled. The addition of an extra electron to any of the $2p$-orbitals require energy as it breaks the stability of N. Whereas oxygen has 4 electrons in $2p$ orbitals and acquires stable configuration i.e., $2p^3$ configuration after removing one electron. So, nitrogen has positive electron gain enthalpy whereas oxygen has negative. however, oxygen has lower ionisation enthalpy than nitrogen.

25. How would you explain the fact that sodium's first ionisation enthalpy is lower than magnesium's while its second Ionisation enthalpy is higher?

Ans. When one electron is removed from a sodium atom, the ion produced gives the configuration of an inert gas, neon. Removing the second electron from a stable noble gas configuration requires more energy. Thus the sodium's first ionisation enthalpy is lower than magnesium while its second ionisation energy is higher.

SHORT ANSWER Type-II Questions (SA-II)

[3 marks]

26. *p*-Block elements form acidic, basic, and amphoteric oxides. Explain each property by giving two examples and also write the reactions of these oxides with water. [NCERT Exemplar]

Ans. Because of the following reasons, the oxides of *p-block* elements display acidic, basic, and amphoteric properties:

(1) Ionisation enthalpy: More will be the ionisation enthalpy, the stronger will be the acid. When the ionisation enthalpy is greater, its oxide will attain acidic nature whereas when it is low it will attain basic nature. And when it has transitional ionisation enthalpy, it will be amphoteric.

(2) Oxidation states: The oxidation state of the elements is directly related to the acidic nature of the elements.

The higher the oxidation state, the higher will be an acidic character of the oxide.

For example, SO_2 is less acidic than SO_3.

(3) Electronegativity: electronegativity of the element is directly proportional to the acidic nature of the element's oxide: the more electronegative the element, the more acidic its oxide will be. For example, B_2O_3 is less acidic than N_2O_3

$$\underset{\text{Boron trioxide}}{B_2O_3} + 3H_2O \rightarrow \underset{\text{Orthoboric acid}}{2H_3BO_3}$$

$$\underset{\text{Dichloro heptoxide}}{Cl_2O_7} + H_2O \rightarrow \underset{\text{Perchloric acid}}{2HClO_4}$$

27. Write the electronic configuration of Cr (Z = 24). Justify your answer.

(A) Electron gain enthalpy of F is less negative than that of Cl. Explain.

(B) Electronegativity of elements increases on moving from left to right in the periodic table. [Diksha]

Ans. The electronic configuration of Cr (Z = 24) is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$/[Ar] $3d^5 4s^1$. This configuration is exceptional as half-filled orbitals have extra stability.

(A) E.G.E of fluorine is less than that of chlorine due to two reasons:

(1) Small size fluorine atom makes the 2p subshell more compact. This results in repulsion among electrons of the valence shell and also with electrons to be added. Due to this F atom has less tendency to accept an electron.

(2) Because of the small size of fluorine, there will be high electron density around the nucleus. This high electron density screens the nucleus. Because of this, effective nuclear charge gets decreased. Thus, the electron is having less attraction during addition. Hence, electron affinity of fluorine gets decreased.

(B) Electronegativity of elements increases on moving from left to right in the periodic table because of decrease in the size of the atom and the increase in nuclear charge.

LONG ANSWER Type Questions (LA)

[4 & 5 marks]

28. Discuss and compare the trend in ionisation enthalpy of the elements of the group 1 with those of group 17 elements. **[NCERT Exemplar]**

Ans. On moving from Li to F, across the second period, within the similar principal quantum level, the introduction of consecutive electrons to orbitals resulting in poor shielding effect by the inner vitality of electrons to compensate for the excessive attraction of the electron to the nucleus. Thus, because of increasing nuclear charge across a period, it exceeds the shielding.

As a result, the ionisation enthalpy rises along a period from left to right, and the outermost electrons are held more securely. Because the outermost electron gets farther away from the nucleus as you travel down a group, the nuclear charge is shielded more by the electrons in the inner levels. In this situation, the rise in shielding exceeds the increasing nuclear charge and down a group, less energy is required to cast off the outermost electron. As a result, the group ionisation enthalpy decreases.

29. The energy of an electron in the hydrogen atom's ground state is 2.18 × 10^{-18} J. Calculate the atomic hydrogen ionisation enthalpy in terms of J mol^{-1}.

Ans. To calculate the ionisation energy of an electron of the hydrogen atom, we will find the difference between the energy of an electron in its ground state which can be written as E_1 and the energy of an electron at infinity which can be written as E_∞ i.e.,

$$E_\infty - E_1.$$

In the ground state, an electron's energy of the hydrogen atom is 2.18 × 10^{-18} J. (Given)

E_1 for the hydrogen atom = – 2.18 × 10^{-18} J.

And at infinity, an electron's energy is 0. This means, E_∞ for the hydrogen atom = 0.

Thus, hydrogen's ionisation enthalpy (in terms of joules):

$$E_\infty - E_1 = 0-(-2.18 \times 10^{-18} J) = 2.18 \times 10^{-18} J.$$

Hence, from a hydrogen atom, the required energy for the removal of an electron in terms of a joule is 2.18 × 10^{-18} J.

But in question, we need the ionisation enthalpy of atomic hydrogen in terms of J mol^{-1} which means we have to find the required energy for removal of 1 mole of electrons.

As we know that 6.022 × 10^{23} particles i.e., Avogadro's number of particles is present in 1 mole.

Hence, the amount of required energy for removal 1 mole of electrons = 2.18 × 10^{-18} × 6.022 × 10^{23} = 13.30 × 10^5 J mol^{-1}.

30. Define ionisation enthalpy. Discuss the factors affecting ionisation enthalpy of the elements and their trends in the periodic table. **[NCERT Exemplar]**

Ans. The quantity of energy required to cast off the loosely bound electron from the remote gaseous atom is referred to as Ionisation enthalpy.

The following factors that influence ionisation enthalpy are:

(1) the atom's size

(2) nuclear power

(3) orbitals that are half-filled and fully-filled

(4) the orbital shape

Its trends in the periodic table:

(1) The ionisation enthalpies are usually in this order (with a few exceptions):

$$(\Delta_i H_1) < (\Delta_i H_2) < (\Delta_i H_3)$$

(2) Moving from top to bottom in a group reduces the ionisation enthalpy.

(3) Moving from left to right in a period raises the ionisation enthalpy.

Two elements must be considered to comprehend these trends:

(1) Electrons are drawn to the nucleus by their attraction.

(2) Electron repulsion is the act of electrons repelling one another.