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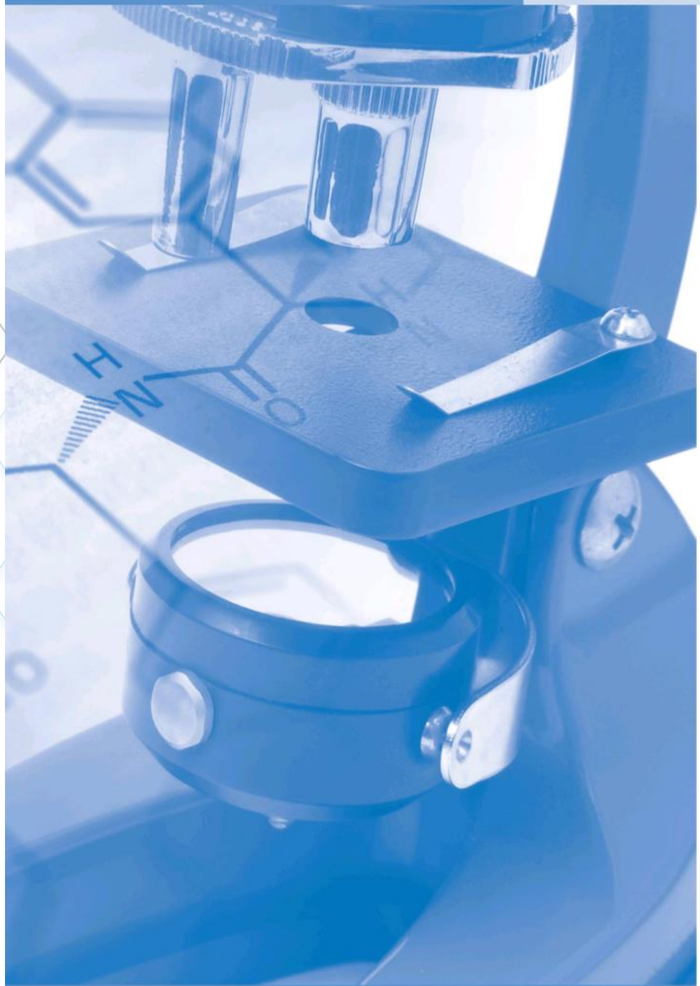
Periodic Table



Remember

Before beginning this chapter, you should be able to:

- Understand earlier versions of classification of elements with emphasis on Mendeleeff's periodic table
- Describe modern periodic table

Key Ideas

After completing this chapter, you should be able to:

- understand the historical aspects of the development of periodic table
- study different periodic laws and the arrangement of elements into groups and periods in the periodic table
- study the trends of atomic size, ionization energy, electron affinity, electropositivity, oxidising and reducing property etc among the elements in the periodic table
- learn comparison of elements based on their periodic properties.

Chemistry involves the study of various elements and their compounds which are quite essential since they invariably find a number of applications in various fields of industry. The discovery of elements is a continuous process with new elements being added to the existing ones. There are 116 authorized elements at present and few more elements have been discovered but yet to be authorized. The study of 100 odd elements and the innumerable compounds formed by them is a tedious task. Therefore, classification of elements has been necessitated for a systematic study. This idea of classification of elements ultimately culminated into the arrangement of elements into a periodic table.

The periodic table of elements can be considered as a single document which reflects the behaviour of elements in a simple and logical way. That means, just by knowing the position of an element in the periodic table, it is possible to predict the nature and behaviour of the elements. In fact, the organization of the elements into a periodic table was a major breakthrough in the history of chemistry.

Historical Aspects

Classification of elements into metals, non-metals and metalloids could not serve the purpose because the elements within a category also show characteristic differences in properties. One of the earliest attempts of classification was made by Dobereiner on the basis of atomic weights. According to this classification, the elements are grouped into triads in which the atomic weight of the middle element is the approximate average of the atomic weights of the other two elements. This classification had a major drawback in that only few elements could be grouped into triads which reflected proper trends in properties.

Another similar attempt was made by Newland who proposed the famous law of octaves. According to Newland's law of octaves, the properties of every eighth element are a kind of repetition of the first. This classification could hold good only up to atomic weight 40 and failed beyond that. The first systematic classification of all the then known elements was made by Mendeleeff by taking atomic weight as the fundamental characteristic property.

Mendeleeff's Periodic Table

According to Mendeleeff, the properties of elements were related to their atomic weights. On this basis, he formulated his periodic law. Mendeleeff's periodic law states that the physical and chemical properties of the elements are the periodic functions of their atomic weights. Therefore, the elements have been arranged in the increasing order of their atomic weights. He compiled a periodic table by arranging the 63 elements known then in eight vertical columns called groups and seven horizontal rows known as periods. Later when the noble gases were discovered, Mendeleeff gave a modified periodic table by including the noble gases in zero group. He categorized the table into eight (columns) groups and seven periods (rows).The groups range from zero to eight and these except the zero and eighth group were further subdivided into two subgroups A and B. He numbered the periods from 1 to 7 of which the 7th period is incomplete.

Merits of Mendeleeff's Periodic Table

(i) Mendeleeff left some gaps for some undiscovered elements in the periodic table. The properties of these elements could be predicted based on the properties of other elements present in the same group. He named these missing elements as Eka boron, Eka aluminium and Eka silicon which were later discovered and named as scandium, gallium and germanium, respectively.

(ii) Mendeleeff corrected the doubtful atomic weights of some elements like indium, beryllium and uranium.

Limitations of Mendeleeff's Periodic Table

(i) In certain parts of the periodic table, element with a higher atomic weight was placed before that of the lower atomic weight in order to maintain the gradation in properties. These are known as anomalous pairs. Examples of anomalous pairs are the following:

(i) Co	Ni	(ii) Te	I	(iii) Ar	K
58•9	58•6	127•6	126•9	39•94	39•10

(ii) Isotopes of an element have different atomic weights. But there is no position for the isotopes in Mendeleeff's periodic table.

(iii) Three groups of three transition elements are placed in group VIII of the periodic table. They should have been given separate positions.

(iv) Transition elements were placed with other elements under the same group.

(v) Coinage metals were placed along with alkali metals.

Periodic System of the Elements in Groups and Periods

Series	Groups of Elements								
	0	I	II	III	IV	V	VI	VII	VIII
1	—	Hydrogen H 1.008	—	—	—	—	—	—	
2	Helium He 4.0	Lithium Li 7.03	Beryllium Be 9.1	Boron B 11.0	Carbon C 12.0	Nitrogen N 14.04	Oxygen O 16.00	Fluorine F 19.0	
3	Neon Ne 19.9	Sodium Na 23.05	Magnesium Mg 24.3	Aluminium Al 27.0	Silicon Si 28.4	Phosphorus P 31.0	Sulphur S 32.06	Chlorine Cl 35.45	
4	Argon Ar 38	Potassium K 39.1	Calcium Ca 40.1	Scandium Sc 44.1	Titanium Ti 48.1	Vanadium V 51.4	Chromium Cr 52.1	Manganese Mn 55.0	Iron Fe 55.9 Cobalt Co 59 Nickel Ni 59 (Cu)
5		Copper Cu 63.6	Zinc Zn 65.4	Gallium Ga 70.0	Germanium Ge 72.3	Arsenic As 75	Selenium Se 79	Bromine Br 79.95	
6	Krypton Kr 81.8	Rubidium Rb 85.4	Strontium Sr 87.6	Yttrium Y 89.0	Zirconium Zr 90.6	Niobium Nb 94.0	Molybdenum Mo 96.0	—	Ruthenium Ru 101.7 Rhodium Rh 103.0 Palladium Pd 106.5 (Ag)
7		Silver Ag 107.9	Cadmium Cd 112.4	Indium In 114.0	Tin Sn 119.0	Antimony Sb 120.0	Tellurium Te 127	Iodine I 127	
8	Xenon Xe 128	Caesium Cs 132.9	Barium Ba 137.4	Lanthanum La 139	Cerium Ce 140	—	—	—	— — —
9		—	—	—	—	—	—	—	

Series	Groups of Elements								
	0	I	II	III	IV	V	VI	VII	VIII
10	—	—	—	Ytterbium Yb 173	—	Tantalum Ta 183	Tungsten W 184	—	Osmium Os 191 Iridium Ir 193 Platinum Pt 194.9 (Au)
11		Gold Au 197.2	Mercury Hg 200.0	Thallium TI 204.1	Lead Pb 206.9	Bismuth Bi 208	—	—	
12	—	—	Radium Rd 224	—	Thorium Th 232	—	Uranium U 239		
Higher Saline Oxides									
	R	R_2O	RO	R_2O_3	RO_2	R_2O_5	RO_3	R_2O_7	RO_4
Higher Gaseous Hydrogen Compounds									
					RH_4	RH_3	RH_2	RH	

Owing to these limitations and the later discovery of atomic structure, Mendeleeff's periodic table has no practical significance at present. Nevertheless, its significance lies in the fact that the concept of periodicity in properties and the resultant arrangement of elements into a periodic table originated from Mendeleeff's classification of elements.

After the discovery of fundamental particles, it has been established that electrons only take part in chemical reactions. Therefore, Moseley considered atomic number as the fundamental characteristic of elements.

Modern Periodic Table—The Long Form of the Periodic Table

In this, the atomic number, i.e., the total nuclear charge is used as the basis of the periodic table instead of the atomic weight. This observation led to the modern periodic law. The modern periodic law states that the physical and chemical properties of elements are the periodic functions of their atomic numbers or electronic configurations."

Thus, according to the modern periodic law, if the elements are arranged in the order of their increasing atomic numbers, the elements with similar properties are repeated after certain regular intervals.

The periodic table consists of 7 horizontal rows called periods and 18 vertical columns called groups. The groups are numbered from I to VIII and the zero group. The groups from I to VII are further divided into A and B. VIII group consists of three vertical columns. All the noble gases are kept in the zero group.

Periods	Description
1.	This is the very short period consisting of only two elements, hydrogen and helium. The elements have only one shell.
2.	This is a short period. These elements have two shells. It has eight elements, i.e., from Li to Ne. These are called bridge elements.
3.	This is called a short period. It has 8 elements ranging from Na to Ar. These are called the typical elements.

Periods	Description
4.	This is called a long period. It has 18 elements, i.e., from K to Kr. It consists of first transition series elements, i.e., from Sc to Zn.
5.	This is the second long period. It has 18 elements, i.e., from Rb to Xe. It consists of second transition series elements, i.e., from Y to Cd.
6.	This is the very long period. It has 32 elements, i.e., from Cs to Rn. It has the third transition series elements, i.e., from La to Hg and also the 14 inner transition elements, i.e., from Ce to Lu.
7.	Incomplete period. The elements are radioactive in nature. The elements succeeding uranium are all artificial elements. These are called transuranic elements. This period includes the elements of fourth transition series which is incomplete and the 14 inner transition series elements, i.e., from Th to Lr.

Periodic Table

Division of Periodic Table into Blocks

Depending on the subshell into which the differentiating electron enters, the elements are categorized into four blocks. The newly added electron in an atom of an element is called the differentiating electron.

Blocks	Description	Groups
s-block	The differentiating electron enters the valence s-subshell.	Group I A → alkali metals Group II A → alkaline earth metals Left part of the periodic table. It consists of only metals.
p-block	The differentiating electron enters the valence p-subshell	III A IV A V A VI A VII A (halogens) Zero group It consists of metals, non-metals and metalloids and noble gases. p-block constitutes right part of the periodic table. s- and p-block elements except zero group are called representative elements.
d-block	The differentiating electron enters the penultimate d-subshell	III B, IV B,V B,VI B,VII B,VIII, I B, II B These are called transition elements. It consists of only metals placed in between left and right parts of the periodic table.
f-block	The differentiating electron enters the antepenultimate f-subshell	III B These are the inner transition elements kept separately at the bottom of the periodic table. It consists of only metals.

Merits of the Long Form of the Periodic Table

(i) The classification is based on the atomic number of elements, which is more fundamental than the atomic weight.

(ii) Elements belonging to the same group have same number of electrons in their valence shells and show similar chemical characteristics.

(iii) Variation of chemical properties along a period is correlated with gradual filling of electrons in a particular shell in the period.

1
H
1.008

1 ← Atomic number, Z
H ← Element symbol
1.008 ← Relative atomic mass, A_r

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.008																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.01	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.54	30 Zn 65.41	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.91	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98.91	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.40	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.90	54 Xe 131.30
55 Cs 132.91	56 Ba 137.34	La–Lu	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.37	82 Pb 207.19	83 Bi 208.98	84 Po 210	85 At 210	86 Rn 222
87 Fr 223	88 Ra 226.03	Ac–Lr	104 Rf [261]	105 Db [262]	106 Sg [266]	107 Bh [264]	108 Hs [277]	109 Mt [268]	110 Ds [271]	111 Rg [272]	112 Uub [285]						

Lanthanoids	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm 146.92	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.92	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
Actinoids	89 Ac 227.03	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu 239.05	95 Am 241.06	96 Cm 244.07	97 Bk 249.08	98 Cf 252.08	99 Es 252.09	100 Fm 257.10	101 Md 258.10	102 No 259	103 Lr 262

(iv) Placement of inert gases at the end of each period is logical since it represents completion of corresponding valence shells. Inertness of these elements also could be explained on the basis of electronic configuration.

(v) Certain types of elements like
- **(a)** active metals,
- **(b)** transition metals,
- **(c)** noble gases and
- **(d)** inner transition elements (lanthanides and actinides) have their specific locations in this periodic table.

(vi) This arrangement of the elements is easy to remember and to reproduce since it is based on electronic configuration.

Defects of the Long Form of the Periodic Table

(i) The position of hydrogen in the periodic table is controversial because hydrogen with one electron in its valence shell shows similarities with both alkali metals and halogens. Therefore, placing hydrogen in IA group is not completely justifiable.

(ii) The elements lanthanides and actinides could not be placed in the main body of the modern periodic table.

(iii) This periodic table does not reflect the exact distribution of electrons of some of the transition and inner transition elements.

The properties of the elements, however, can be predicted based on their position in the periodic table.

EXAMPLE

What are the basic differences between Mendeleeff's periodic table and modern periodic table?

SOLUTION

(i) Mendeleeff's periodic table was based on atomic weight whereas modern periodic table was based on atomic number.

(ii) Mendeleef's periodic table consists of 8 groups and 7 periods whereas modern periodic table consists of 18 groups and 7 periods.

EXAMPLE

In the modern periodic table, magnesium is surrounded by elements with atomic numbers 4, 11, 13 and 20. Identify the elements. Which of these have chemical properties resembling magnesium?

SOLUTION

The element $_{12}Mg$ is present in the centre of the elements having atomic numbers 4, 11, 13 and 20. The atomic number 4 and 20 resemble magnesium because they belong to same group having same number of valence electrons.

	Be = 4 (2, 2)	
Na = 11 2, 8, 1	Mg = 12 (2, 8, 2)	Al = 13 (2, 8, 3)
	Ca = 20 (2, 8, 8, 2)	

EXAMPLE

How was the problem of placement of isotopes in Mendeleeff's periodic table overcome in modern periodic table?

SOLUTION

Modern periodic table was based on the atomic number. Since isotopes have similar atomic numbers, the problem of placement of isotopes was solved.

Periodicity

The regular gradation in the properties of elements and their repetition at regular intervals is called periodicity. The properties which show such regular trend are called periodic properties.

Periodic Properties

(i) Atomic Size

Atomic size is defined in terms of atomic radius which is the distance between the nucleus and the valence shell. However, it is not possible to measure the atomic radius directly as an individual atom cannot be isolated.

There are two different ways of measuring the atomic radius:

(a) **Covalent radius**: The atomic size of the non-metals is determined by taking half the distance between the nuclei of two like atoms bonded together by a single covalent bond.

(b) **Metallic radius**: It is determined by taking half the distance between the nuclei of two adjacent metal atoms in a metallic crystal.

Variation of Atomic Size

(a) **Period**: The atomic size generally decreases from left to right along the period. This is because along a period the electrons are added to the same energy level, which leads to an increase in the effective nuclear charge on the outermost shell, thereby decreasing the size of the atom.

Exception: In the noble gases, due to interelectronic repulsions, the size is larger than the corresponding halogen.

(b) **Group**: On going down the group, a new shell is being added from one element to the other. As the electron is added to a shell farther from the nucleus, the effective nuclear force of attraction on the valence shell decreases thereby increasing the size of the atom.

Exception: Atomic sizes of the 5th period elements do not show much increase from the corresponding 4th period elements of same groups. This is because of the addition of 14 electrons in the antepenultimate "f" subshell of lanthanide elements. The contraction in atomic size among lanthanide elements is called lanthanide contraction.

(ii) Ionization Energy

Ionization energy of an element is the minimum amount of energy required to remove the outermost electron from an isolated neutral gaseous atom of that element in its lowest energy state to form a cation. This is called the first ionization potential.

The minimum amount of energy required to remove the electron from a gaseous atom is called its **first ionization potential**. The **second ionization potential** is the energy required to remove the electron from the outermost shell of a unipositively charged gaseous ion. Similarly, we can define the third and the fourth ionization energies or ionization potentials. If the first ionization energy is considered as I_1 and the second as I_2 and so on then, $I_1 < I_2 < I_3 < I_4 < \ldots\ldots$ Theoretically, an atom can have as many ionization potentials as there are electrons in it.

This successive increase in the values of ionization potential is due to increase in effective nuclear charge from the neutral atom to the respective ions.

Variation of Ionization Potential

(a) **Period**: Ionization energy generally increases from left to right along a period. This is because along a period, the atomic size decreases and the valence shell is closer to the nucleus. Thus more energy is required to remove an electron from the valence shell.

(b) **Group**: The ionization energy decreases from top to bottom in a group because the atomic size increases down the group, thus resulting in decrease in the effective nuclear force of attraction.

Graphical Representation of Variation of IP with Atomic Numbers

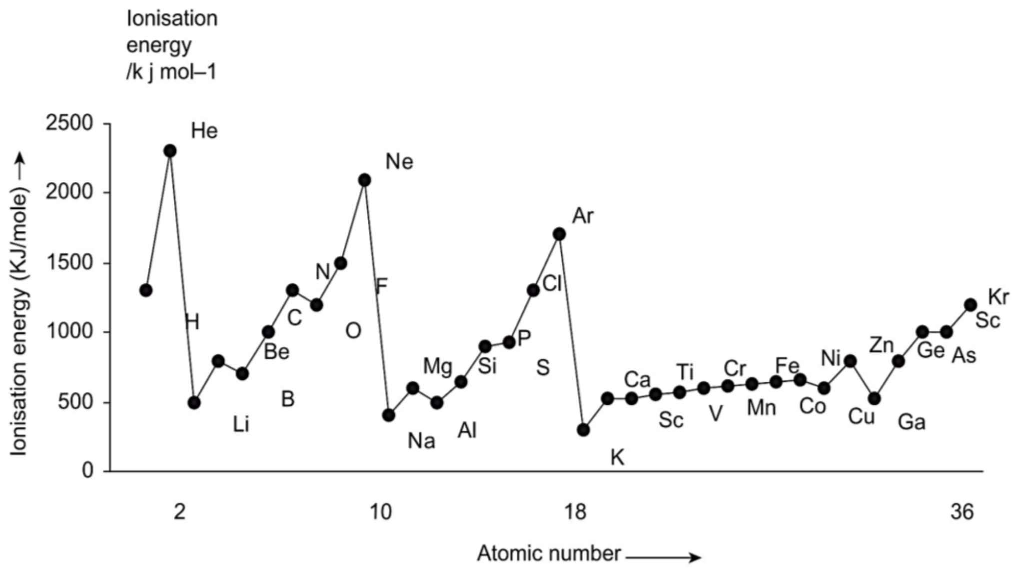


Figure 4.1 The ionization energies of the elements (hydrogen to krypton)

Conclusions from Fig. 4.1:

(a) The noble gases occupy the highest positions indicating that these have exceptionally high ionization potential values.

(b) Variation of ionization potential from nitrogen to oxygen and phosphorus to sulphur does not follow the regular trend.

(c) The variation of IP in transition metals is more regular and marginal in comparison to the representative elements.

(iii) Electron Affinity

The amount of energy released when an electron is added to an isolated neutral gaseous atom in its lowest energy state to produce an anion is called the electron affinity.

Variation of Electron Affinity

(a) **Period**: As the atomic size decreases and nuclear charge increases from left to right across the period, the electron affinity values generally increase.

(b) Group: The electron affinity gradually decreases from top to bottom due to the steady increase in atomic radius and the resultant decrease of effective nuclear charge.

Graphical representation of electron affinity with atomic numbers

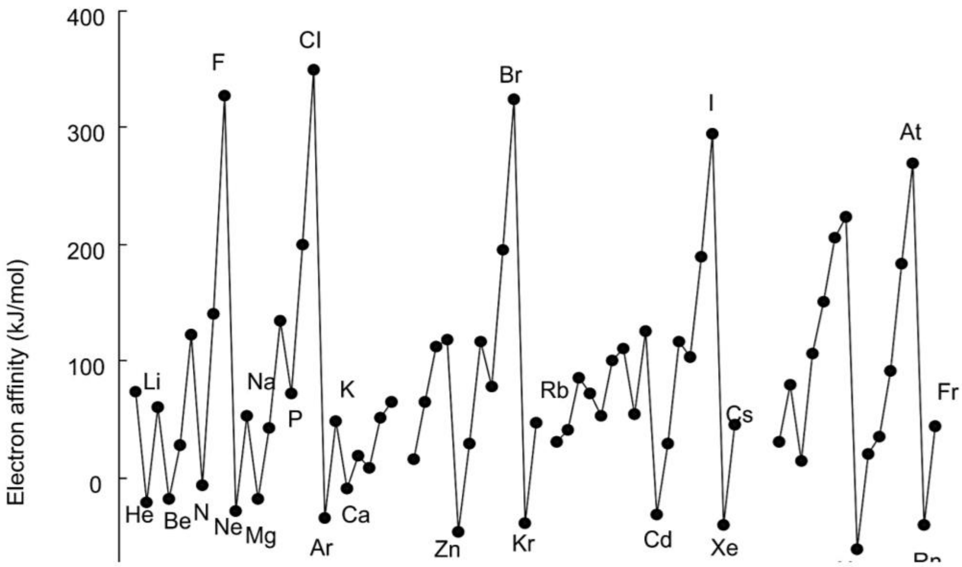


Figure 4.2 Atomic number

Conclusions drawn from Fig. 4.2:

(a) The noble gases show zero electron affinity values indicating that the addition of an electron to those atoms is not possible.

(b) The elements of IIA group show negative values for electron affinity which indicates that the addition of electron is associated with the absorption of energy.

(c) Halogens occupy the highest positions in the curve which shows maximum tendency to form negative ions.

(iv) Electronegativity

Electronegativity of an element is defined as the tendency of an atom to attract the shared pair of electrons towards itself in a bonded molecule.

Since electronegativity is a relative value, measurement of absolute value is not possible. One way of measurement of electronegativity is to take an average of ionization potential and electron affinity. In the other method, fluorine is taken as standard and the electronegativity values of other elements are calculated with respect to fluorine.

Variation of Electronegativity

(a) Period: Electronegativity increases along a period from left to right. This is due to the increase in nuclear charge and decrease in atomic size.

(b) Group: Electronegativity decreases from top to bottom in a group due to an increase in the atomic size.

(v) Electropositive Character

Electropositive character of an element is its ability to lose one or more electrons to form a positively charged ion.

The electropositive character of an element depends on its ionization potential. An element in which the atom has greater ability to lose an electron means that it has low ionization potential and high electropositivity. Such an element acts as a good reducing agent.

An element which has a high ionization potential has low electropositivity and hence cannot give out electron easily. Such an element acts as a good oxidizing agent.

Variation of Electropositive Character

(a) Period: The electropositive character decreases (the oxidizing capacity increases and the reducing capacity decreases) from left to right along a period due to the increase in ionization potential.

(b) Group: Due to the decrease of ionization potential and electron affinity down the group the electropositivity increases, oxidizing capacity decreases and reducing capacity increases.

(vi) Metallic and Non-metallic Property

Metals can generally give away electrons easily and form positive ions while non-metals generally accept electrons and form negative ions.

Variation of Metallic and Non-metallic Properties

(a) Period: As the electronegativity increases from left to right along a period, the metallic character decreases and the non-metallic character increases.

(b) Group: As the electronegativity decreases from top to bottom in a group the metallic character increases and non-metallic character decreases.

Thus the right side of the periodic table (p-block) comprises of non-metals, metalloids and metals. The left side of the table is completely occupied by metals. The middle portion (d-block) as well as the bottom portion (f-block) is comprised of metals.

This kind of distribution shows that though the periodic table is comprised of metals, non-metals and metalloids, 75% of the elements are metals. Since the metals have certain characteristic properties which clearly distinguish them from the rest of the elements, the study of metals gained significance. The metals also differ among themselves in their properties and hence they are placed in the respective positions in the periodic table.

For example, alkali metals which possess one electron in their valence shells are grouped together in group IA of the modern periodic table. Similarly, the alkaline earth metals which possess two electrons in their valence shells are placed in Group IIA of the periodic table. These two categories of metals belong to "s-" block of the periodic table.

The other two categories are transition metals and inner transition metals. These metals have certain specific properties which are not exhibited by any other groups of elements. Some of the specific properties of transition metals which clearly distinguish them from the other metals are paramagnetic nature of the atoms and ions, formation of coloured compounds and exhibition of variable oxidation states in their compounds.

Formation of Coloured Compounds

All the compounds of alkali and alkaline earth metals are colourless. However, many compounds of transition metals have their characteristic colours due to the presence of incomplete d-subshells.

Examples

$CuSO_4\ 5H_2O$	Blue
$K_2Cr_2O_7$	Orange
$KMnO_4$	Violet
$CuCl_2$	Bluish green
$FeSO_4$	Light green

These elements also do not show a regular trend in the common periodic properties studied separately. Therefore, these are placed in the "d-" block of the periodic table and are studied separately.

Comparative Study of Alkali Metals and Alkaline Earth Metals

Alkali metals and alkaline earth metals belong to "s-" block of the periodic table and they show regular gradation in the properties.

Melting and boiling points of alkali metals are lower than those of alkaline earth metals as well as transition metals to their right.

Reaction with Water

Alkali metals react vigorously with water and liberate hydrogen gas. Reaction of lithium is less vigorous:

$$\underset{\text{(alkali metal)}}{M} + 2H_2O \rightarrow MOH + H_2$$

Alkaline earth metals also react with water to liberate hydrogen gas:

$$\underset{\text{(alkaline earth metal)}}{M} + 2H_2O \rightarrow M(OH)_2 + H_2$$

The reaction is less vigorous in case of alkaline earth metals when compared to alkali metals.

Reaction with Oxygen

Alkali and alkaline earth metals burn in the presence of O_2 to give their corresponding oxides:

$$4M + O_2 \rightarrow 2M_2O$$
(alkali metals)

$$2M + O_2 \rightarrow 2MO$$
(alkaline earth metals)

Sodium (alkali metal) and barium (alkaline earth metal) also give their stable peroxides along with the oxides:

$$2Na + O_2 \rightarrow Na_2O_2$$

$$Ba + O_2 \rightarrow BaO_2$$

Reaction with Hydrogen

Alkali metals and alkaline earth metals react with H_2 and form their stable hydrides:

$$2M + H_2 \rightarrow 2MH$$
(alkali metals)

$$M + H_2 \rightarrow MH_2$$
(alkaline earth metals)

Beryllium hydride is unstable.

All the hydrides of alkali and alkaline earth metals get readily hydrolysed by water and liberate hydrogen gas.

$$2MH + 2H_2O \rightarrow 2MOH + 2H_2\uparrow$$
(alkali metals)

$$MH_2 + 2H_2O \rightarrow M(OH)_2 + 2H_2\uparrow$$
(alkaline earth metals)

Reaction with Chlorine

All the alkali metals and alkaline earth metals react with chlorine to give their corresponding chlorides:

$$2M + Cl_2 \rightarrow 2MCl$$
(alkali metals)

$$M + Cl_2 \rightarrow MCl_2$$
(alkaline earth metals)

EXAMPLE

(a) Name two elements you would expect to show chemical reactions similar to sodium.

What is the basis of your choice?

(b) Arrange the elements present in that group in the increasing order of reactivity.

Give reason.

SOLUTION

(a) Lithium and potassium show same chemical properties as sodium, because all the three elements possess same number of valence electrons and belong to same group that is IA group.

(b) Increasing order of reactivity is Li < Na < K < Rb < Cs. This is due to decrease in ionization potential values from top to bottom in a group.

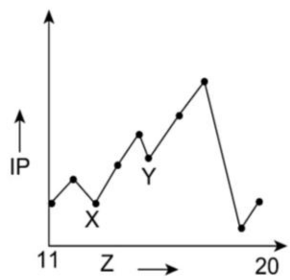
EXAMPLE

Teacher explained periodicity of ionization potential in a period and group with reasons. She then drew a curve of IP vs atomic number for elements with Z = 11 to Z = 20. On the basis of the curve, she asked the following questions to students.

(a) Identify the elements occupying peaks and bottommost points.

(b) Also identify 'X', and 'Y' marked in the curve.

(c) In what respect 'X' and 'Y' deviate from their preceeding elements?



When students answered these correctly, she was satisfied and told – "Now, I will give you a brain teaser to get me answer by next chemistry class". Students enthusiastically said – "yes"! Then teacher wrote the following question on board __

(d) "Why do 'X' and 'Y' show deviations from their preceeding elements in the trend of IP values"? What explanation would you give if you were there in the class?

SOLUTION

(a) Peak is occupied by noble gas, Ar. Bottom most point is occupied by alkali metal, K.

(b) 'X' is aluminium and 'Y' is sulphur

(c) Generally, with increase in atomic number, IP value should also increase. However, from magnesium to aluminium, IP value decreases. Similarly, from phosphorus to sulphur, IP value decreases.

(d) In magnesium, there is a pair of electrons in 3s orbital. Removal of electron is slightly more difficult. Moreover, since the electron to be removed belongs to 's' subshell having greater penetrating power, the IP value is relatively higher. In case of aluminium, there is an unpaired 'p' electron, the removal a which is easier. Hence IP value is less.

EXAMPLE

Ionization energy is the energy absorbed whereas electron affinity is the energy released. Explain.

SOLUTION

When an electron is present in the outermost shell of an atom, it is bound by nuclear force of attraction. In order to remove this electron, some amount of energy should be supplied to overcome the nuclear force of attraction. During the addition of an electron to a neutral atom, the electron approaches to a neutral atom due to nuclear force of attraction. This process involves work done towards the force of attraction and hence energy is released. Thus, ionization potential is the energy absorbed and electron affinity is the energy released.

EXAMPLE

(a) One mole of atoms of an element 'X' absorb 800 kJ of energy for the formation of unipositive ions. For the convertion of one mole of 'X' atoms into tripositive ions, 6800 kJ of energy is required. Second and third ionization energies are in 2 : 3 ratio. Calculate the 2nd and 3rd ionization energies.

(b) First and second ionization energies of $Be_{(g)}$ are 900 $kJmol^{-1}$ and 1750 $kJmol^{-1}$ respectively. Calculate approximate percentage of $Be^{2+}{}_{(g)}$ ions, if 1 g of $Be_{(g)}$ absorbs 150 kJ of energy [Be atomic weight is 9].

SOLUTION

(a) $x \rightarrow x^+$, 800 kJ/mol (IP_1)

$x \rightarrow x^{+3}$, 6800 kJ

$\therefore IP_1 + IP_2 + IP_3 = 6800 \text{ kJ} \Rightarrow IP_2 + IP_3 = 6800 - 800 = 6000$

$x^+ \rightarrow x^{+2}\ IP_2;\ x^{+2} \rightarrow x^{+3}\ IP_3$

$\because IP_2 : IP_3 = 2 : 3 \Rightarrow \dfrac{IP_2}{IP_3} = \dfrac{2}{3}$

$\Rightarrow IP_2 = \frac{2}{3} . IP_3$ or $IP_3 = \frac{3IP_2}{2}$

$\Rightarrow IP_2 + IP_3 = 6000$

$\Rightarrow IP_2 + \frac{3}{2} . IP_2 = 6000 \Rightarrow 2\ IP_2 + 2\ IP_3 = 6000 \times 2$

$IP_2 = \frac{6000}{5} \times 2 = 2400$ kJ/mol; $IP_3 = \frac{6000}{5} \times 3 = 3600$ kJ/mol

(b) 1 mole (9 g) of beryllium requires 900 kJ to dislodge the 1st electron.

$\therefore$ 1 g of Be requires $= \frac{1 \times 900}{9} = 100$ kj

$\therefore$ Difference in the amount of energy absorbed = 150 – 100 = 50 kJ

To dislodge the 2nd electron from 1 mole of Beryllium ion, 1750 kJ is required.

50 kJ is sufficient for $\frac{50}{1750} \times 9 = 0.25 = 25\%$

Percentage of $Be^{2+}{}_{(g)}$ ions is 25%.

EXAMPLE

The ionic size of Cl^- is greater than that of K^+ ion, though they are isoelectronic. Explain.

SOLUTION

Nuclear forces of attraction on valence electrons is higher in K^+ ion, than in that in Cl^-. This is due to the greater nuclear charge of K^+ ion. So ionic size of Cl^- ion is greater than that of K^+ ion.

EXAMPLE

A sharp change in the atomic radius is observed from lithium to potassium but a gradual change in the atomic radius is observed from potassium to caesium. Explain.

SOLUTION

A sharp increase in the atomic radius is observed from lithium to potassium due to increase in the number of shells. But a gradual change is observed from potassium to caesium due to filling up of d electrons, in the pennultanate shell which increases the effective nuclear charge. Hence increase in size is gradual, not sharp.

When students answered these correctly, she was satisfied and told – "Now, I will give you a brain teaser to get me answer by next chemistry class". Students enthusiastically said – "yes"! Then teacher wrote the following question on board __

(d) "Why do 'X' and 'Y' show deviations from their preceeding elements in the trend of IP values"? What explanation would you give if you were there in the class?

SOLUTION

(a) Peak is occupied by noble gas, Ar. Bottom most point is occupied by alkali metal, K.

(b) 'X' is aluminium and 'Y' is sulphur

(c) Generally, with increase in atomic number, IP value should also increase. However, from magnesium to aluminium, IP value decreases. Similarly, from phosphorus to sulphur, IP value decreases.

(d) In magnesium, there is a pair of electrons in 3s orbital. Removal of electron is slightly more difficult. Moreover, since the electron to be removed belongs to 's' subshell having greater penetrating power, the IP value is relatively higher. In case of aluminium, there is an unpaired 'p' electron, the removal a which is easier. Hence IP value is less.

EXAMPLE

Ionization energy is the energy absorbed whereas electron affinity is the energy released. Explain.

SOLUTION

When an electron is present in the outermost shell of an atom, it is bound by nuclear force of attraction. In order to remove this electron, some amount of energy should be supplied to overcome the nuclear force of attraction. During the addition of an electron to a neutral atom, the electron approaches to a neutral atom due to nuclear force of attraction. This process involves work done towards the force of attraction and hence energy is released. Thus, ionization potential is the energy absorbed and electron affinity is the energy released.

EXAMPLE

(a) One mole of atoms of an element 'X' absorb 800 kJ of energy for the formation of unipositive ions. For the convertion of one mole of 'X' atoms into tripositive ions, 6800 kJ of energy is required. Second and third ionization energies are in 2 : 3 ratio. Calculate the 2^{nd} and 3^{rd} ionization energies.

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PRACTICE QUESTIONS

TEST YOUR CONCEPTS

Very Short Answer Type Questions

1. Why did Mendeleeff arrange the elements in the increasing order of their atomic weight in his periodic table?
2. How was the problem of placement of isotopes in Mendeleeff's periodic table overcome in the modern periodic table?
3. Why did Newland's law of octaves fail?
4. Zero group elements are known as _____ gases.
5. Why are the "s-" and "p-" blocks elements called representative elements?
6. General electronic configuration of s-block elements is _____.
7. Describe the trend of oxidizing and reducing property along a period and a group of the periodic table.
8. Name the elements whose atomic weight was corrected by Mendeleeff.
9. Describe the trend of electron affinity along a period.
10. Describe the trend of atomic size along a period and a group in the periodic table.
11. Which periods are called the short periods and which are called the long periods? Why?
12. Describe the trend of ionization potential along a period and a group in the periodic table.
13. Why are the chemical properties of the elements belonging to the same group similar?
14. Why is it that hydrogen can be placed in both I A group and VII A group of the modern periodic table?
15. An element, the atomic number with 36 belongs to _______ group and _______ period.
16. What do you mean by representative elements? Give an example.
17. Which group contains the noble gases and what are the special characteristics of these gases?
18. General electronic configuration of chalcogens is ____________.
19. Describe the trend of ionization potential down the group in the periodic table.
20. Describe how the metallic and non-metallic characters vary along a period and group of the periodic table.
21. The elements present in ______ and ______ periods can exhibit diagonal relationship.
22. Ionization potential value of Cs is equal to the electron affinity value of its corresponding ______.
23. ______ block elements can exhibit variable number of oxidation states.
24. What is a covalent radius?
25. In the long form of periodic table ________ group contains the maximum number of elements.
26. Anion size is always _____ than the parent atom.
27. How does the electron affinity vary in the first transition series?
28. In the f-block elements differentiating electron enters into f-orbital of ______ shell.
29. _____ group elements are the strongest reducing agents.
30. What is metallic radius? Give the implication of measuring a metallic radius.

Short Answer Type Questions

31. What are the merits of Mendeleeff's periodic table?
32. What are the achievements and the limitations of Newland's classification?
33. What are the advantages of the classification of elements?
34. Why are the lanthanides and the actinides given a separate position in the modern periodic table?
35. Are all the elements present in III B of same type? Explain.

36. Describe how effective nuclear charge changes along a period and a group in a periodic table.

37. Why is the modern periodic table also referred to as Bohr's table?

38. Why are the values of the first, the second, the third ionization energies different from each other?

39. How does electron affinity vary in halogens?

40. What is the basic difference between Mendeleeff's periodic table and modern periodic table?

41. Among all the periods which one is incomplete and what are the main important properties exhibited by all the elements present in this period?

42. What is the difference between electronegativity and electropositivity?

43. Mention different categories of elements in the long form of the periodic table.

44. What is the difference between transition and inner transition elements?

45. How does filling of valence electrons takes place in p-block, d-block and f-block elements?

Essay Type Questions

46. What are the merits and limitations of the long form of the periodic table?

47. Describe briefly Mendeleeff 's periodic table along with its merits.

48. What are the limitations of Mendeleeff 's periodic table?

49. Explain briefly the periods of the modern periodic table.

50. Explain briefly about the groups of the modern periodic table and also classify the elements depending on electron filling in s-, p-, d- and f-subshells along with the position in the periodic table.

CONCEPT APPLICATION

Level 1

***Direction for questions from 1 to 7:* State whether the following statements are true or false.**

1. Modern periodic table is based on atomic number.
2. In a period, the first element has the smallest size.
3. "s" and "p" block elements except zero group are known as representative elements.
4. The inert gas present in the second long period is Xe.
5. Among the isoelectronic ions Na^+, Mg^{2+} and Al^{3+}, Na^+ ion have the largest size.
6. Alkali metals act as strong reducing agents by undergoing reduction.
7. Electronegativity is the property of an atom in a bonded molecule.

***Direction for questions from 8 to 14:* Fill in the blanks.**

8. In the periodic table, vertical columns of elements are called _____ and horizontal rows of elements are called _____.
9. In the long form of the periodic table, ________ group elements can release maximum amount of energy by the addition of electron into its valence shell.
10. The ascending order of the first ionization potential of C, N, O and F is _____.
11. The total number of inner transition elements present in the 7th period are _________.
12. Amongst the species Br^-, Br, Br^+, the smallest one in size is _____.

13. In most of the lanthanides, the penultimate shell contains _________ electrons in ______orbitals.

14. Metals exhibit _____ oxidation states in their compounds.

***Direction for question 15:* Match the entries in column A with the appropriate ones in column B.**

15.

Column A		Column B
A. s-block	( )	a. Radioactive metal
B. Fluorine	( )	b. Highest first IP value
C. Chlorine	( )	c. Lanthanides
D. d-block	( )	d. Highly reactive metals
E. 4f-series	( )	e. Heavy metals
F. Helium	( )	f. Always form −1 oxidation state
G. Francium	( )	g. Highest electron affinity value
H. Noble gases	( )	h. Zero group elements

***Direction for questions from 16 to 45:* For each of the questions, four choices have been provided.**

Select the correct alternative.

16. ***Assertion*** (A): First ionization energy of beryllium is greater than that of boron.

Reason (R): Boron has larger size than beryllium.

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

17. Which of the following properties is a periodic property?
(a) colour (b) melting point
(c) refractive index (d) atomic size

18. Which one of the following electronic configurations corresponds to the most electropositive character?
(a) $[He]2s^1$ (b) $[Ne]3s^1$
(c) $[Ar]4s^1$ (d) $[Xe]6s^1$

19. The general electronic configuration of representative elements is
(a) $ns^{1-2}\ np^{1-6}$ (b) $ns^{1-2}\ np^{1-5}$
(c) $ns^{1-2}\ (n-1)p^{1-5}$ (d) $ns^{1-2}\ (n-1)p^{1-6}$

20. Predict the formula of a compound formed between a metal "M" which has 1st, 2nd, 3rd IP values as 518, 7314, 9820 kJ mol^{-1}, respectively and a halogen "X."
(a) MX_2 (b) M_2X_3
(c) MX_3 (d) MX

21. In a period, from the left to right the electron affinity increases, but alkaline earth metals have lower electron affinity than alkali metals because
(a) alkaline earth metals have lesser atomic radius than alkali metals.
(b) alkaline earth have higher electronegativity than alkali metals.
(c) alkaline earth metals have completely filled "s-" orbitals.
(d) alkaline earth metals have lesser electronegativity than alkali metals.

22. 5f series elements are known as
(a) lanthanides
(b) representative elements
(c) transition elements
(d) actinides

23. The first ionization energies of Li, Be, B and C are in the order:
(a) Li > Be < B < C
(b) Li < Be > B < C
(c) Li > Be > B > C
(d) Li < Be > B > C

24. Predict the powerful oxidizing agent in the 3rd period:
(a) sulphur (b) sodium
(c) chlorine (d) bromine

25. An element with atomic number "32" belongs to
(a) 4th period, VIA group
(b) 3rd period, IVA group
(c) 4th period, IVA group
(d) 5th period, VA group

26. The element having the electronic configuration $[Kr]\ 4d^{10}\ 4f^{14}\ 5s^2\ 5p^6\ 5d^2\ 6s^2$ belongs to
(a) s-block (b) p-block
(c) d-block (d) f-block

PRACTICE QUESTIONS

27. The number of valence electrons that can be present in the second element of any period is

(a) 1 (b) 2
(c) 5 (d) 7

28. Transition metals can exhibit variable valency. It is because of

(a) the smaller atomic radius
(b) the high screening effect
(c) the very less energy difference between (n−1) d-subshell and ns-subshell
(d) the high nuclear charge

29. Arrange O, F, Cl and N in the descending order of their electronegativity:

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

30. In the sixth period, the orbitals which are completely filled are

(a) 6s, 6p, 5d, 5f (b) 6s, 6p, 6d, 6f
(c) 6s, 5f, 6d, 6p (d) 6s, 4f, 5d, 6p

31. Which of the following sequence of explanation is appropriate for explaining the reason for the periodicity of reducing property in a period or group?

(1) Tendency to undergo oxidation decreases in a period and increases in a group.
(2) In a period ionization energy increases and in a group it decreases.
(3) In a period the atomic size decreases and in a group it increases.
(4) The elements present in the left side of the periodic table have strong reducing property.

(a) 3 2 1 4 (b) 3 4 2 1
(c) 4 3 1 2 (d) 4 2 1 3

32. Which of the following triads does not follow Dobereiner's law of triads?

(a) Li, Na, K (b) Ca, Sr, Ba
(c) Be, Mg, Ca (d) Cu, Ag, Au

33. X belongs to IA or 1st group and 5th period and Y succeeds X in the group. Z succeeds Y in the period. Arrange the following statements in the correct sequence in order to arrange X, Y and Z in the increasing order of their atomic sizes.

(1) Effect of number of valence electrons and number of shells on the atomic size.
(2) Identification of the elements X, Y and Z.
(3) Determination of the number of shells and the number of valence electrons present in X, Y and Z.
(4) Determination of the positions of Y and Z in the periodic table based on the position of X.

(a) 4 2 3 1 (b) 2 4 3 1
(c) 3 1 4 (d) 4 3 1

34. Which among the following is not an anomalous pair in Mendeleeff's periodic table?

(a) Co, Ni (b) Te, I
(c) Ar, K (d) Sc, Ga

35. The energy released when an electron is added to the valence shell of a neutral, gaseous, isolated atom is called __________

(a) electronegativity (b) ionization potential
(c) electron affinity (d) lattice energy

36. Which one of the following electronic configuration corresponds to the element with maximum electropositive character?

(a) $[Kr]5s^1$ (b) $[Ne]3s^1$
(c) $[Ar]4s^1$ (d) $[Xe]6s^1$

37. Which of the following elements has maximum electronegativity?

(a) P (b) S
(c) Al (d) Si

38. Which of the following elements acts as the best reducing agent?

(a) Na (b) Cl
(c) Mg (d) F

39. The elements present in d-block are

(a) metals and non-metals
(b) only metals
(c) only non-metals
(d) metals, metalloids and non-metals

40. The total number of elements present in the 6th period is

(a) 18 (b) 31
(c) 32 (d) 17

41. An element with atomic number "32" belongs to

(a) 4th period, VIA group or 16th group
(b) 3rd period, IVA group or 14th group
(c) 4th period, IVA group or 14th group
(d) 5th period, VA group or 15th group

42. The formulae of oxides formed when barium burns in presence of oxygen are

(a) Ba_2O, BaO (b) BaO, BaO_2
(c) BaO, Ba_2O_2 (d) Ba_2O, BaO_2

43. If the consecutive ionization energies of an element A are 496, 4564, 6918, 9542 kJ/mole respectively, the formula of the oxide formed by A is

(a) A_2O_3 (b) AO
(c) AO_2 (d) A_2O

44. The numerical value of ionization energy of a unipositive ion is approximately equal to

(a) IP value of neutral atom
(b) EA value of neutral atom
(c) IP value of dipositive ion
(d) EA value of dipositive ion

45. Which of the following sets of atomic numbers corresponds to elements belonging to s-, d-, f-, p-blocks respectively?

(a) 35, 37, 29, 70 (b) 35, 29, 70, 37
(c) 37, 29, 70, 35 (d) 37, 29, 35, 70

Level 2

1. Why is the atomic radius of oxygen slightly more than that of nitrogen?

2. Predict the position of an element with atomic number 55 in the periodic table and explain the chemical behaviour of that element.

3. Why is electron affinity of chlorine more than that of fluorine?

4. The ionic size of Cl^- is greater than that of K^+ ion, though they are isoelectronic. Explain.

5. Why do most of the inner transition elements exhibit stable +3 oxidation state?

6. Xenon has an octet configuration in the valence shell but it shows variable valency with the highly electronegative elements like oxygen and fluorine. Comment.

7. Though oxygen and sulphur belong to the same group, sulphur shows many oxidation states, whereas oxygen does not show a higher oxidation state than that of –2. Comment.

8. First and second ionization energies of $Be_{(g)}$ are 900 kJ mol^{-1} and 1750 kJ mol^{-1} approximately. Calculate approximate percentage of $Be^{2+}{}_{(g)}$ ions, if 1 g of $Be_{(g)}$ absorbs 150 kJ of energy [atomic weight of Be is 9].

9. The separation of one lanthanide from another is an extremely difficult task, almost as difficult as the separation of isotopes of one element. Justify.

10. A sharp change in the atomic radius is observed from lithium to potassium but a gradual change in the atomic radius is observed from potassium to caesium. Explain.

11. In a particular transition series, the atomic radius reaches minimum up to group VIII elements, then again increases towards the end of the series. Give appropriate reasons.

12. Ionization potential decreases down the group in representative elements while transition elements of the 6th period have greater ionization potential values compared to the elements of the 5th period. Justify.

13. The densities of transition metals are greater than the alkali and alkaline earth metals. Justify.

14. Successive ionization potentials of an element "X" are IP_1 = 520 kJ mol^{-1}, IP_2 = 7,298 kJ mol^{-1} and IP_3 = 10,815 kJ mol^{-1}. Give the probable formulae of chloride and oxide of "X."

15. An element "A" has atomic number 28. What is the electronic configuration of the elements below it in the periodic table? Predict their chemical properties.

***Directions for questions from 16 to 25:* Application-Based Questions**

16. When a Dobereiner triad is considered, the sum of atomic weights of extreme elements X and Z is 177.6 and difference of Z and Y is five times the number of protons present in a neon atom. Identify X, Y and Z.

17. Iron can form two types of ions. Compare the ionic sizes of the respective ions.

18. Three elements A, B and C have atomic numbers as x, x + 1 and x + 2. The atomic size of C is greater than that of B but atomic size of B is less than that

PRACTICE QUESTIONS

of A. Predict the positions of A, B and C in the periodic table and justify the trend in atomic size.

19. A teacher while explaining the definition of ionization potential gave an example: Ionization potential value of sodium is 140 kJ/mole. That means, 140 kJ of energy is required to remove the lone electrons from "3s" orbitals of one mole of gaseous sodium atoms. Then, a student Rinku stood up immediately and asked: sodium is a solid metal. Why should we consider it in gaseous state? What explanation would have been given by the teacher?

20. Formation of X^- ion from "X" is associated with the release of 144 kJ/mole energy. However, formation of X^{-2} ion from X^- is associated with the absorption of 840 kJ of energy. How do you account for this?

21. A tripositive ion of element "X" has six electrons with l = 0 and five electrons with l = 2.

(a) Predict the group and period to which "X" belongs.

(b) Does the element show variable valencies? Justify.

22. Identify the position of the elements having outer electronic configuration:

(a) $ns^2\ np^5$ for n = 3

(b) $(n-1)d^2\ ns^2$ for n = 4

23. The steady decrease in size of lanthanide elements and their ions is called lanthanide contraction. In lanthanide elements, the differentiating electron enters into antepenultimate "f-" subshell.

(a) How does the entry of differentiating electron into antepenultimate shell affect the atomic and ionic sizes?

(b) Densities of 6th period elements are double to those of 5th period elements. Why?

24. An element "X" has three completely filled shells and forms a stable uninegative ion and the total number of electrons in the third and fourth shells is same. Identify the elements which form unipositive and dipositive ions which are isoelectronic with X^- ion. Compare the reactivities of these two elements with water. Give a reason in support of your answer.

25. Why are the electron affinities of alkaline earth metals positive?

$$M_{(g)} + e^- \rightarrow M^-_{(g)}\ \Delta H = +\ \text{(Electron affinity)}$$

Level 3

1. Generally transition elements form coloured compounds. But Sc^{+3} forms colourless compounds. Justify.

2. The first ionization energy of magnesium is greater than that of sodium, whereas the reverse is true for second ionization energy. Explain.

3. Why do some of the periodic properties like atomic size, electronegativity and ionization potential of transition and inner transition elements not show as much variation as the representative elements?

4. Why are the electron affinities of IIA group and IIB group elements less than zero?

5. The density of elements belonging to the 3rd transition series is double that of the elements belonging to the 2nd transition series provided they are in the same group unlike alkali and alkaline earth metals. Comment.

Directions for questions from 6 to 10: **Application-Based Questions**

6. In the d-block elements as we move along a period the atomic radius decreases in the beginning, reaches minimum and then increases towards the end of the series. Explain with appropriate reasons.

7. Why are the values of electron affinity of the IIA group and IIB group elements less than zero?

8. Why do some of the periodic properties like atomic size, electronegativity and ionization potential of transition and inner transition elements do not show as much variation as the representative elements?

9. Three elements X, Y and Z have atomic numbers 22, 40 and 72, respectively. Comment on the trend of change in ionization potential values from X to Y and Y to Z.

10. There is a sharp decrease in the ionization energies from boron to aluminium but it is almost the same from aluminium to gallium. Justify.

CONCEPT APPLICATION

Level 1

True or false

1. True	2. False	3. True	4. True
5. False	6. True	7. True	

Fill in the blanks

8. groups, periods	9. VII	10. C, O, N, F	11. 14
12. Br^+	13. 8, s, p	14. Positive	

Match the following

15. A : d	B : f	C : g	D : e
E : c	F : b	G : a	H : h

Multiple choice questions

16. b	17. d	18. d	19. c
20. d	21. c	22. d	23. b
24. **c**	25. c	26. c	27. b
28. c	29. b	30. d	

HINTS AND EXPLANATION

Solutions for questions from 31 to 45:

31. (i) In a period the atomic size decreases and in a group it increases.
(ii) In a period ionization energy increases and in a group it decreases.
(iii) Tendency to undergo oxidation decreases in a period and increases in a group.
(iv) The elements present in the left side of the periodic table have strong reducing property.

32. Since the atomic weight of Ag is not equal to the mean of the atomic weights of Cu and Au, (Cu, Ag and Au) triad do not follow Dobereiner's law of triads.

33. (i) Determination of the positions of Y and Z in the periodic table based on the position of X.
(ii) Determination of the number of shells and the number of valence electrons present in X, Y and Z
(iii) Effect of number of valence electrons and number of shells on atomic size.

34. In case of first three pairs, the atomic weights of latter elements are greater than the former elements; so, they are called anomalous pairs. The last pair is not an anomalous pair.

35. Electron affinity is defined as the amount of energy released when an electron is added to the valence shell of gaseous, neutral and isolated atom.

36. Electropositive character increases from top to bottom in a group. Hence the element with electronic configuration [Xe] $6s^1$ has the maximum electropositive character.

37. Electronegativity increases from left to right in a period. Hence, S has the maximum value.

38. Sodium belongs to IA group which are good reducing agents.

39. All d-block elements are metals.

40. The total number of element present in the 6th period is 32 elements including 14 lanthanides.

41. Z = 32 corresponds to germanium which belongs to the IVA group (or) 4th group and 4th period. Electronic configuration is [Ar] $3d^{10}\ 4s^2\ 4p^2$.

42. $2Ba + O_2 \rightarrow 2BaO$ (Barium oxide)

$Ba + O_2 \rightarrow BaO_2$ (Barium peroxide)

43. In element A there is a large difference between I and II ionization energies; so, it can form the stable ion A^+. The oxide of A would be A_2O.

44. The numerical value of ionization energy of a uni-positive ion is approximately equal to the electron affinity of dipositive ion.

45. Electronic configuration of an element with atomic number 37 is [Kr] $5s^1$. 29 is [Ar] $3d^{10}$ $4s^1$, 70 is [Xe] $4f^{14}$ $5d^0$ $6s^2$. 35 is [Ar] $3d^{10}$ $4s^2$ $4p^5$.

Hence, 37, 29, 70 and 35 belong to s, d, f and p, respectively.

Level 2

1. (i) comparing electronic configuration of nitrogen and oxygen
(ii) electronic arrangement in the valence shell of nitrogen and oxygen
(iii) relation between the electronic arrangement and the stability of the atom
(iv) effect of stability on size of the atom

2. (i) electronic configuration of the element and predict group and period
(ii) correlation between the position of the element in the periodic table and its chemical behaviour

3. (i) factors affecting electron affinity
(ii) electronic configuration of chlorine and fluorine
(iii) comparison of the number of valence electrons with respect to the atomic size between chlorine and fluorine

4. (i) nuclear charge is different in two ions
(ii) effect of nuclear charge on the size of the atom

5. (i) electronic configuration of inner transition elements, entry of differentiating electron
(ii) comparison of energy of valence shell and that of penultimate shell
(iii) participation of penultimate shell electrons in bonding

6. (i) Xenon belongs to the 5th period.
(ii) atomic size of xenon in comparison to the other noble gases
(iii) comparison of influence of electronegativity of fluorine and oxygen

7. (i) comparison of electronic configuration of oxygen and sulphur
(ii) presence of vacant "d" subshell and scope for excited state configuration

8. (i) amount of energy absorbed by1 g of Be to form Be^+ and Be^{+2}
(ii) energy required for conversion of 1 g $Be_{(g)}$ into Be^+ ions
(iii) calculation of additional amount of energy
(iv) conversion of Be^+ ions to Be^{+2} ions
(v) 25%

9. (i) differentiating electron of lanthanides
(ii) lanthanide contraction and its effect on atomic and ionic sizes of lanthanide elements
(iii) dependence of chemical properties on the atomic sizes
(iv) properties exploited for separation of different elements

10. (i) factors affecting the atomic radius
(ii) the effect of increase in inner shells on atomic radius

11. (i) electron filling in transition elements
(ii) effect on effective nuclear charge
(iii) relation between effective nuclear charge and atomic radius
(iv) comparison of electronic configuration in (n – 1) d subshell of the transition elements of each group
(v) effect on atomic size

12. (i) factors effecting ionization potential
(ii) comparison of the screening effects of s, p, d and f electrons

13. (i) factors affecting density
(ii) comparison of electron filling in alkali, alkaline earth and transition metals
(iii) effect on atomic size
(iv) comparison of mass with respect to atomic volume among alkali, alkaline earth and transition metals

14. (i) comparing IP values; valency of X

(ii) abrupt change in IP value

(iii) valence electronic configuration of stable ion

(iv) valencies of the corresponding positive and negative radicals

15. (i) electronic configuration of element A

(ii) electronic configuration corresponding to Z = 28

(iii) position of A in the periodic table

(iv) position of the element present below A in the periodic table and its electronic configuration

(v) determination of the properties from its position

***Solutions for questions from 16 to 25:* Application-Based Questions**

16.
$$x + z = 177 \times 6 \qquad 177.6$$
$$z - y = 5 \times 10 \qquad 50.0$$
$$(-)(+) \quad - \qquad \overline{127.6}$$
$$x + y = 127.6$$

As X, Y and Z form Dobereiner triad

$y = \frac{x+z}{2}$, where x, y and z are atomic masses of X, Y and Z, respectively.

$$y = \frac{177.6}{2} = 88.8$$

$x = 127.6 - 88.8 = 38.8$

$z = 177.6 - 38.8 = 138.8$

$x = 38 \times 8 \approx 40$

$X = 40 \Rightarrow$ Ca

$Y = 88 \times 80 \Rightarrow$ Sr

$Z = 138 \times 6 \Rightarrow$ Ba

17. Both Fe^{2+} and Fe^{3+} have the same number of protons but the number of electron in Fe^{3+} is less than that of Fe^{2+}. Hence the effective nuclear charge is more in the case of Fe^{3+} and thereby the ionic size of Fe^{3+} is less than that of Fe^{2+}.

18. As the atomic numbers of A, B and C differ by 1 unit each they must be belonging to adjacent groups. Along a period the atomic size decreases till the VII A group and then increases till the zero group. As the atomic size of C is greater than B though it differs by one unit in its atomic number it must be a zero group element. Though the atomic number of B is greater than A, it has smaller size than B and this shows that B is a halogen, i.e., VII A group and A is a chalcogen, i.e., VI A group.

19. For any element, ionization potential is considered for gaseous state only. Ionization potential corresponds to the energy required to overcome the force of attraction exerted by the nucleus on the electron without considering any other kinds of attractions. In gaseous state, the intermolecular forces of attractions being negligible, removal of electron from a neutral gaseous atom involves an amount of energy equal to IP only. In solid and liquid states, the intermolecular forces of attraction are stronger and the energy required to remove electron involves not only IP but also the energy required to overcome the other forces. Thus, for any element, IP is considered only for the isolated gaseous atoms irrespective of the state in which it exists.

20. During the addition of an electron to a neutral atom, work is done towards the nuclear force of attraction. Hence, energy is released. After the addition of an electron, the uninegative species formed, repels the 2nd electron. Hence during the addition of the second electron, work is done against the force of repulsion which requires supply of energy.

21. (a) $l = 0$ corresponds to "s" subshell and $l = 2$ corresponds to "d" subshell. Further, 6 electrons in "s" subshell and 5 electrons in "d" subshell means the ion should have the electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$. Since it is a tripositive ion, the atom could have lost two electrons from 4s subshell and one electron from 3d subshell. Therefore, atomic number is 26. Electronic configuration is [Ar] $3d^6\ 4s^2$. It belongs to the VIII group and 4th period.

(b) Yes. Since Fe is a transition element it shows variable oxidation state. In transition elements, the electrons enter into the penultimate shell instead of valence shell. Energy difference between ns and (n – 1)d orbitals is very less. Hence both ns and (n – 1)d electrons are available for chemical bond formation. So, they exhibit variable valency.

22. (a) ns^2, np^5 for n = 3

$1s^2, 2s^2, 2p^6, 3s^2, sp^5$

So, the period is 3rd and the group is 17th (VII A).

(b) $(n-1)\ d^2\ ns^2$ for n = 4

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^2$

So, the period is 4th and the group is 4^{th} (IV B).

23. (a) In lanthanide elements, the newly added electron enters into anti-penultimate shell keeping the configuration of penultimate and valence shells same. The addition of electrons to inner shell should result in increased screening effect. But the "f-" subshell has a very poor screening effect. So, the addition of electrons to the inner "f-" subshell does not result in corresponding increase in atomic size. Therefore, the increasing nuclear charge from cerium to lutetium (71) results in drastic decrease in atomic and ionic sizes.

(b) Between the 5th and 6th periods, the intervening lanthanides increase the nuclear charge and lanthanide contraction results in decrease of atomic size. As a result of this, mass of atom increases and atomic size decreases. This results in drastic increase in density of elements from the 5th to 6th periods.

24. The element "X" must be a halogen since it forms the stable X^- ion. As there are three completely filled shells in "X," the electronic configuration of "X" could be 2, 8, 18, 18, 7 and X^- is 2, 8, 18, 18, 8. Since there are 54 e^-s in X^-, the corresponding elements which form unipositive and dipositive ions should be cesium (55) and barium (56). Cesium reacts more vigorously with water than barium. This is because reactivity of alkali metals is more due to their greater atomic size when compared with alkaline earth metals. Moreover, alkaline earth metals have greater IP values than alkali metals. Hence, they react slowly with water.

25. All the electrons present in alkaline earth metals are paired. Therefore to add one extra electron to these atoms, work has to be done against the force of repulsion. Hence electron affinity of an alkaline earth metal is a positive.

Level 3

1. (i) electronic configuration of Sc
(ii) energy difference among the orbitals of "d-" subshell.
(iii) basic principle behind exhibition of colour
(iv) comparison of electronic configuration of alkali, alkaline earth and transition metals
(v) electronic configuration of Sc+3
(vi) comparison of electronic configuration of scandium ion with other transition metal ions

2. (i) electronic configuration of sodium and magnesium
(ii) comparison of atomic size of sodium and magnesium
(iii) effect of atomic size on first ionization energy
(iv) electronic configuration of stable ions of sodium and magnesium
(v) relation between the stability of ion and the respective ionization energy

3. (i) shell into which the differentiating electron enters
(ii) general electronic configuration of representative elements
(iii) comparison with general electronic configurations of transition and inner transition elements
(iv) comparison of effective nuclear charge and its effect on the periodic properties

4. (i) electronic configuration of II A and II B
(ii) change in energy that takes place due to the addition of electron to the neutral atom of the elements belonging to the Groups IIA and IIB
(iii) relation between change in energy and electron affinity values

5. (i) factors affecting density
(ii) general trend of atomic sizes in groups of transition elements
(iii) effect of inclusion of lanthanide elements on the atomic size of the elements belonging to the 3rd transition series
(iv) factors affecting the density of an atom
(v) change in atomic size of alkali and alkaline earth metals along the group
(vi) comparison of change of densities along the group among alkali, alkaline earth and transition metals

***Solutions for questions from 6 to 10:* Application-Based Questions**

6. In a particular series of d-block elements, electron filling takes place in the penultimate shell. Hence their atomic size decreases up to the 7th group and reaches minimum and then again increases because with the pairing of electrons, repulsive force also increases. The repulsion among the added electrons exceeds the increased force of attraction (due to the increase of nuclear charge) after the 8th group. Hence, atomic radius increases slightly.

7. Electronic configuration of II A and II B group elements is ns^2np^0. Addition of an electron to these elements requires energy to overcome repulsive forces between completely filled ns electrons and electron going to be added. Hence, electron affinity values are less than zero.

8. In case of representative elements, from left to right in a period, the nuclear charge increases with increase in atomic number. Since the number of inner shells and inner electrons remain constant, the increased effective nuclear charge decreases the atomic size proportionately. In case of transition and inner transition elements, the electrons are being added to the penultimate and anti-penultimate shells, respectively. There is slight increase in screening effect which counterbalances the increased nuclear charge to some extent and hence the decrease in atomic size is less pronounced than in representative elements. Since the other properties in turn depend on atomic size, they also show the same trend as atomic size.

9. The respective elements X, Y and Z are Ti, Zr and Hf. They belong to IVB group of 3d, 4d and 5d series of transition elements, i.e., 4th period, 5th period and 6th period. As they belong to same group and the value of ionization potential decreases down the group, the same trend follows from Ti to Zr. However, the IP value of Hf is double to that of Zr. This is because of intervening lanthanide elements which cause lanthanide contraction. As there is no pronounced increase in atomic size due to lanthanide contraction, the increased nuclear charge increases the effective nuclear force of attraction drastically. Thus, there is a drastic increase in IP value.

10. Among IIIA group elements, boron has the highest ionization potential due to smaller size. The sharp decrease in ionization energy from boron to aluminium is due to increase in atomic radius. The atomic radius increases due to decrease in effective nuclear charge. In gallium contraction in atomic size is significant due to the presence of completely filled d-electrons in the penultimate shell. Thus, the effective nuclear charge increases. Hence both aluminium and gallium have almost equal ionization energy values.