Revision Notes

Class 10 Science

Chapter 5 - Periodic Classification of Elements

Introduction:

- In the beginning, scientists had classified elements into two broad categories as **metals and non-metals**. Some elements exhibited properties because of which they could neither be classified as metals nor non-metals and hence were called **metalloids**. This classification was not sufficient for scientific study. Over the years, many chemists attempted to make a rational and systematic classification. It was based on the physical and chemical properties of each element. These results were then tabulated in the form of a table.
- **Periodic table** – The table giving the arrangement of the known elements according to their properties so that similar elements fall within the same vertical column and dissimilar elements are separated.

Dobereiner's Triads:

- In 1817, a German chemist named Johann Wolfgang Dobereiner arranged the elements with **similar properties** into groups.
- He identified many such groups which had **three elements** in them. Hence, these groups were termed as **triads**.
- In a triad, the elements were arranged in increasing order of their atomic masses. When the arithmetic mean of the atomic mass of the first and the third element was taken, it came out to be approximately the same as that of the second or the middle element. This is a distinctive feature that can be used to identify if elements form a triad.
- This classification was also not sufficient as not many triads could be identified. Only three were identified as:

Triads	Atomic masses	Arithmetic mean
I		
Lithium	6.9	$\frac{6.9+39.1}{2}=23$
Sodium	23.0	
Potassium	39.1	
II		
Calcium	40.1	$\frac{40.1+137.3}{2}=88.7$
Strontium	87.6	
Barium	137.3	
III		
Chlorine	35.5	$\frac{35.5+126.9}{2}=81.2$
Bromine	79.9	
Iodine	126.9	

Newlands' Law of Octaves:

- In 1866, an English scientist, John Newlands arranged many of the then known elements in the increasing order of their atomic masses. So, he started with the element having the lowest atomic mass (hydrogen) and ended at Thorium which was the 56^{th} element. He noticed that the **eighth element was similar in properties to the first element**.
- It was concluded that there exists some systematic relationship between the order of atomic masses and the repetition of properties of elements. This relationship, when represented in a tabular form, presented a periodic repetition of the properties of the elements. Hence, the term **periodicity** was introduced.
- This had a resemblance to the eight musical notes in both Western as well as Indian music.

Indian sa Western (do)	**re (re)**	**ga (mi)**	**ma (fa)**	**pa (so)**	**da (la)**	**ni (ti)**
H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe
Co and Ni	Cu	Zn	Y	In	As	Se
Br	Rb	Sr	Ce and La	Zr	-	-

- It was found that the Law of Octaves was applicable only up to Calcium because after Calcium, every eighth element did not possess properties similar to that of the first. John Newlands had assumed that only 56 elements existed in nature and no more elements would be discovered in the future. But several new elements were discovered later on, whose properties did not fit into the Law of Octaves. So, this was not sufficient either.

Mendeleev's Periodic Table:

- In 1869, a Russian chemist, Dmitri Ivanovich Mendeleev classified the then known 63 elements based on their physical and chemical properties in the increasing order of the atomic masses in the form of a table.
- He had observed that properties of the elements recur cyclically when they were arranged in the order of their increasing atomic masses. This observation led to the conclusion that the physical and chemical properties of the elements are periodic functions of their atomic masses. This came to be known as the **law of chemical periodicity**.
- Periodic Table is the tabulation of all the known elements in a tablular format based on this law. It contains eight vertical columns called **'groups'** and seven horizontal rows called **'periods'**. Each of the eight groups has two sub-groups A and B. The properties of elements of a sub-group resemble each other more markedly than the properties of those between the elements of the two sub-groups.

Achievements of Mendeleev's Periodic Table:

Some of the important contributions of Mendeleev's periodic table are as follows:

1) **Systematic Study of Elements** – The table provided the arrangements of elements showing similar properties into groups. This was very useful in studying and remembering the properties of a large number of elements in a systematic way.
2) **Prediction of New Elements** – Mendeleev had predicted new elements and had left three blanks for these undiscovered elements. He was able to predict their properties more or less accurately. He named them eka-boron, eka-aluminium and eka-silicon.
3) **Correction of Atomic Masses -** Based on the elements' positions in the periodic table, Mendeleev was able to correct their atomic masses. The atomic mass of beryllium was corrected from 13.5 to 9.0.

Limitations of Mendeleev's classification:

Although Mendeleev's periodic table has many advantages, it could not explain certain things, which are considered as its limitations. They are as follows:

1) **Assigning a position to Hydrogen:** Hydrogen has a electronic configuration as that of alkali metals and combines with halogens, oxygen, sulphur to form compounds, like HCL, H_2O, H_2S and at the same time it exists as a diatomic molecule like halogens. So, Mendeleev was not able to assign a proper position for hydrogen.
2) **Assigning position to isotopes:** The isotopes have similar properties but differ in their atomic masses. Mendeleev's classification would place them in different groups due to their different atomic masses, but isotopes were not placed so as their properties were similar.
3) **Anomalous pairing of some elements:** Mendeleev did not follow the increasing atomic masses but grouped some elements based on similar properties. Argon with an atomic mass of 39.9 was placed before potassium with atomic mass of 39.1. Also some elements with similar properties like copper (Cu) and mercury (Hg) were placed separately, and some very dissimilar elements were placed in one group. Copper was placed in group I, the elements of which had no similarities with copper.

The Modern Periodic Table:

- In the year 1913, an English physicist named **Henry Mosely** found that the **atomic number** of an element, which was denoted by **symbol 'Z'** was a more basic property to group them instead of their atomic masses. Thus Mendeleev's periodic table was modified for the same. The elements were now grouped based on the increasing atomic number.
- This came to be known as the **Modern Periodic Law** and it states, 'properties of the elements are a periodic function of their atomic number'. Hence the new classification of the elements based on this came into existence and was termed as '**Modern Periodic Table'**.
- With this system of grouping it was easy to predict the properties of the elements when they were arranged in the order of increasing atomic numbers. It is to be noted that the **periodicity** of the elements is based on the **electronic configuration** or the number of protons in the nucleus.

Group → / Period ↓	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	4 Be											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	* 71 Lu	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	** 103 Lr	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
*			57 La	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		
**			89 Ac	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		

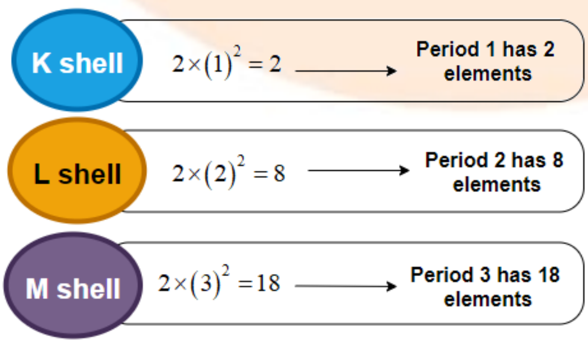
Salient features of the Modern Periodic Table:
The table has **18 vertical columns** that are known by the name of **groups** and **7 horizontal rows** that are named as **periods**.

1) **Periods:**

- There are 7 periods in this table. The periods have the same elements that have the **same valence shell** or the energy shell. Example - Na, Mg, Al, Si, P, S, Cl are placed in the same shell as they have the electronic shells as K, L and M.

- In a period, the number of electrons present in the **energy shells increase by 1** on moving from left to right within a period. Example - Na - 1, Mg - 2, Al - 3, and so on.

- The **number of elements** present in a period can be determined by the formula $2n^2$, where n is the number of the shell from the nucleus. Example:



- The **first period** consists of two elements only namely, hydrogen and helium as they have only 1 valence shell. Example - hydrogen (Z = 1 or shell as K = 1), helium (Z = 2 or shell as K = 2)

- The **second period** has 8 elements with 2 shells and it starts with lithium (Z = 3 or shells as K = 2, L = 1) and ends with neon (Z = 10 or shells as K = 2, L = 8).

- The **third period** has 8 elements with 3 shells and it starts with sodium (Z = 11 or shells as K = 2, L = 8, M = 1) and ends with argon (Z = 18 or shells as K = 2, L = 8, M = 8).

- Similarly the **fourth period** has 18 elements with 4 shells and starts with potassium (Z = 19) and ending with krypton (Z = 36).

- The **fifth period** having 18 elements with 5 shell starts with rubidium (Z = 37) and ends with xenon (Z = 54).

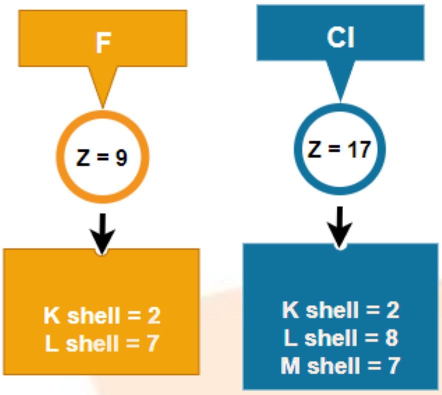
- The **sixth period** with 32 elements has 6 shells and it starts with caesium (Z = 55) ending with radon (Z = 86).

- The **seventh and last period** is incomplete with 19 elements starts francium (Z = 87) and goes on till oganesson (Z = 118).

2) **Groups:**

- There are 18 groups in the periodic table. The group consists of elements that have the same number of electrons in the valence shell or outermost shell of the atom.

- The valence shell predominantly decides the physical or chemical properties of the elements, so the elements of the same group have the same properties due to the same number of valence electrons. Example - fluorine and chlorine have valence electrons as 7 and they belong to group 17.



- The number of shells increases by one unit as the elements move down in the periodic table in a group.

3) Blocks:

- The periodic table is also divided into **4 blocks** that is based on the subshell of the valence electrons. They are:

- **s-Block elements**: All the elements of group 1 and 2 are included in this block and their general electronic configuration is ns^{1-2} Example - Hydrogen (H), Sodium (Na), etc from group 1 and Magnesium (Mg), Calcium (Ca), etc from group 2.

- **p-Block elements**: This includes the elements from group 13 to 18. They have an electronic configuration as ns^2np^{1-6}.

- **d-block elements**: This includes group 3 to 12 elements. They have a general electronic configuration as $(n-1)d^{1-10}ns^{1-2}$.

- **f-block elements**: This block has sets of elements, lanthanides and the actinides. They have the electronic configuration of $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$. The lanthanides starts from Lanthanum (La) - Lutetium (Lu) and the actinides starts from Actinium (Ac) - Lawrencium (Lr).

Position of elements in the periodic table:

- The **position** of the various elements are decided on the basis of their **valence shells** and the **number of electrons** present in their valence shells. Example - Sodium (Z - 11, 2,8,1), so it has 3 shells, so it is placed in period 3 and since it has 1 valence electron in outermost shell, it is placed in group 1.

- The position of an element in the periodic table determines its **chemical nature**. Based on the position of the elements in the periodic table, they can be classified as follows:

  1) **Noble gases:** These are a group of elements placed in group 18, which are tasteless, odourless monoatomic gases that have very low chemical reactivity. There are 6 such gases, namely, Helium (He), Neon (Ne), Argon (Ar), Krypton (Kr), Xenon (Xe), Radon (Rn). They are also referred to as inert gases and due to their inertness, they are suitable to be used where reactions are not required. Example - He is used by deep-sea divers in the breathing gas to prevent toxicity of oxygen, nitrogen and carbon dioxide.

  2) **Normal elements:** All the elements that are placed in groups 1 to 7 are included in this.

  3) **Alkali metals:** The elements in group 1, namely Lithium (Li) - Francium (Fr), except Hydrogen (H) are termed as alkali metals as they tend to form hydroxide with water which are strong alkalis. Thus alkali metals are very reactive and react quickly with water or air. Example - Sodium (Na) reacts violently with oxygen in the air, so it stored in mineral oil.

  4) **Alkaline earth metals:** These include the group 2 elements starting from Beryllium (Be) - Radium (Ra). They are less reactive than the alkali metals that are found as a compound.

  5) **Transition elements:** These include the elements from group 3 to 11. These are so named as they exhibit a transition in their properties from the left to the right, including increase in atomic size, ionization energy, electronegativity.

  6) **Inner transition elements:** These are elements with similar properties, placed at the end of group 3 in period 7 and 8. These are called as the lanthanide series starting from Lanthanum (La) - Lutetium (Lu) in period 6 with 14 elements. The period 7 contains 14 elements starting from Actinium (Ac) - Lawrencium (Lr).

  7) **Halogens:** These include a group of elements in group 17 that are generally non-metals that can exist in solid, liquid and gas form.

They react with the metals to form salts. They are Fluorine (F), Chlorine (Cl), Bromine (Br), Iodine (I), Astatine (At) and Tennessine (Ts).

Properties of the periodic table:

1) Valency:

- The term valency denotes the number of electrons that are gained or lost by an atom in order to complete its outermost shell to have a stable electronic configuration. This valency is the number of electrons present in the valence shell.

- It can be noted that the **valency increases** from left to right in a period, and then decreases.

- The **valency in the group remains the same** through the group, going downward.

- Example - Period 2 elements have atomic number from 3 to 10, so they have 2 shells with increasing number of valence till C and then it decreases. But all the group 2 elements have 2 electrons in their outermost shell, so their valency is 2.

Period 2

Element	Li	Be	B	C	N	O	F	Ne
Atomic number	3	4	5	6	7	8	9	10
Valency	1	2	3	4	3	2	1	0

Increases → Decreases →

Group 2

Element	Be	Mg	Ca	Sr	Ba	Ra
Atomic number	4	12	20	38	56	88
Valency	2	2	2	2	2	2

Same

2) Atomic size:

- The atomic size is determined by the **atomic radius** of the atom and it can be termed as the distance from the centre of the nucleus of the atom to its outermost shell.

- It is seen that the elements are placed across a period from left to right, the **atomic radius decreases**. This is because an increased nuclear charge has the tendency to pull the electrons towards the nucleus, thereby decreasing the radius of the atom, and thus the atomic size too decreases.

- It is not the same in a group. As the elements move down a group, there is an addition of an extra shell, hence their atomic radius and thus **atomic size increases**. Example -

Period 2

Element	Li	Be	B	C	N	O	F	Ne
Atomic radius	167pm	112pm	87pm	67pm	56pm	48pm	42pm	38pm

Decreases →

Group 2

Element	Be	Mg	Ca	Sr	Ba	Ra
Atomic radius	112pm	145pm	194pm	219pm	253pm	N/A

Increases →

3) Ionization energy:

- It is termed as the amount of energy that is required to remove the electrons that are present in the outermost shell of the atom.

- Across the period, since the atomic radius decreases, these electrons in the outermost shell are much closer to the nucleus and therefore it requires more energy to remove them. So, **ionisation energy increases** across a period.

- In the groups, since a shell is being added down the group, so the atomic radius increases which leads to the electrons in the outermost shell being far from the nucleus and therefore it is easy to remove them. Hence in a group, the **ionisation energy decreases** going from top to bottom.

Period 2

Element	Li	Be	B	C	N	O	F	Ne
Ionization energy	5.39172	9.3227	8.298	11,2603	14.5341	13.6181	17.4228	21.5645

Increases →

Group 2

Element	Be	Mg	Ca	Sr	Ba	Ra
Ionization energy	9.3227	7.6462	6.11316	5.6949	5.2117	5.2784

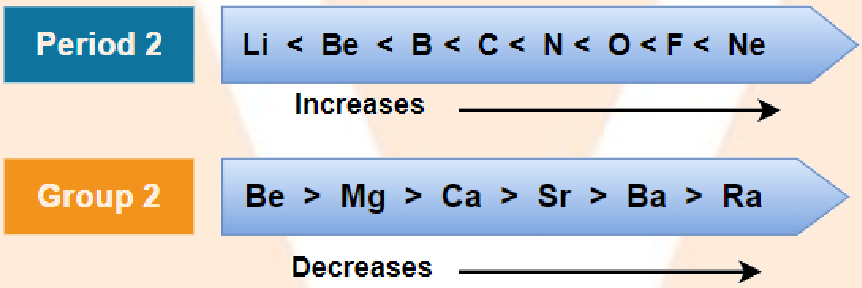
Decreases →

4) Electron affinity:

- It is termed as the amount of energy change as a result of an addition of an electron to the atom or the ability of an electron to accept electrons.

- In a period, from left to right as the nuclear charge increases and atomic size decreases, it is easy for the addition of electrons, which leads to generation of more energy. Hence **electron affinity increases** across a **period.**

- In the group, as atomic size increases, nuclear charge decreases, so lesser number of electrons can be added which leads to lesser energy generation. Hence **electron affinity decreases** from top to bottom in a **group**.



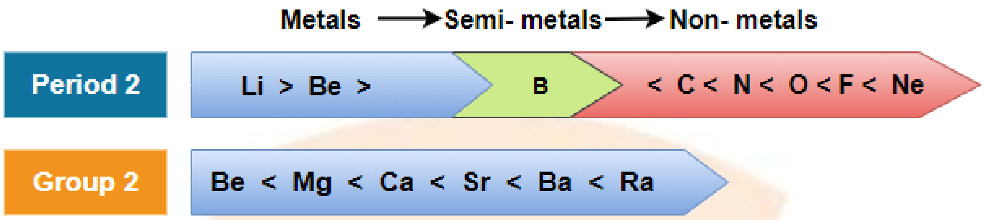
5) Metallic and non-metallic properties:

- The metals are those elements that have a tendency to lose electrons and attain a positive charge. So, the metallic nature of elements means that they are electropositive and have low ionization energy. This metallic character decreases along a period.

- Non-metallic nature of elements indicate that they have an ability to gain electrons and attain a negative charge. So, this non-metallic nature indicates that they are electronegative and have high ionization energies.

- Hence this **non-metallic nature increases** along a **period**. Hence, moving along a period from left to right, the metallic character decreases and non-metallic character increases and there is a semi-metallic nature in between them.

- The **metallic nature increases** down a **group** as the atomic size increases and they have an increased tendency of losing electrons. Example -



6) These trends can be **summarised** with the chart for easy reference: