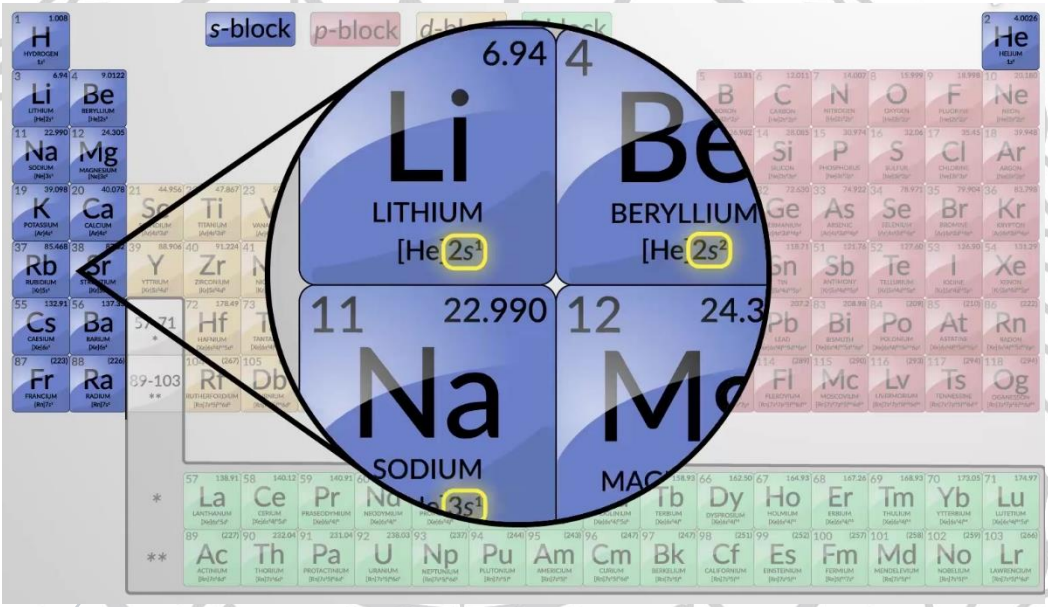
PERIODIC CLASSIFICATION OF ELEMENTS



EDUMAX CLASSES

EDUCATION AT MAXIMUM

Dobereiner's triads and Newland's law of octaves were early attempts at classifying elements into groups based on their properties. Since many new elements were discovered over the course of the 18th and 19th centuries, the broad classification of elements into metals and non-metals became inefficient. Several experiments were conducted in order to identify elements with similar properties and group them together.

It is important to note that the primitive methods of classifying elements, such as Newland's law of octaves and Dobereiner's triads, laid the foundation for the development of the modern periodic table.

What are Dobereiner's Triads?

Dobereiner's triads were groups of elements with similar properties that were identified by the German chemist Johann Wolfgang Dobereiner. He observed that groups of three elements (triads) could be formed in which all the elements shared similar physical and chemical properties.

Dobereiner stated in his **law of triads** that the arithmetic mean of the atomic masses of the first and third element in a triad would be approximately equal to the atomic mass of the second element in that triad. He also suggested that this law could be extended for other quantifiable properties of elements, such as density.

The first of Dobereiner's triads was identified in the year 1817 and was constituted by the alkaline earth metals calcium, strontium and barium. Three more triads were identified by the year 1829. These triads are tabulated below.

Triad 1

This triad was made up of the alkali metals lithium, sodium and potassium.

Triad	Atomic Masses
Lithium	6.94
Sodium	22.99
Potassium	39.1

The arithmetic mean of the masses of potassium and lithium corresponds to 23.02, which is almost equal to the atomic mass of sodium.

Triad 2

As mentioned earlier, calcium, barium and strontium formed another one of Dobereiner's triads.

Triad	Atomic Masses
Calcium	40.1
Strontium	87.6
Barium	137.3

The mean of the masses of barium and calcium corresponds to 88.7.

Triad 3

The halogens chlorine, bromine and iodine constituted one of the triads.

Triad	Atomic Masses
Chlorine	35.4
Bromine	79.9
Iodine	126.9

The mean value of the atomic masses of chlorine and iodine is 81.1.

Triad 4

The fourth triad was formed by the elements sulfur, selenium, and tellurium.

Triad	Atomic Masses
Sulfur	32.1
Selenium	78.9
Tellurium	127.6

The arithmetic mean of the masses of the first and third elements in this triad corresponds to 79.85.

Triad 5

Iron, cobalt and nickel constituted the last of Dobereiner's triads.

Triad	Atomic Masses
Iron	55.8
Cobalt	58.9
Nickel	58.7

However, the mean of the atomic masses of iron and nickel corresponds to 57.3.

Limitations of Dobereiner's Triads

The key shortcomings of Dobereiner's method of classifying elements are listed below.

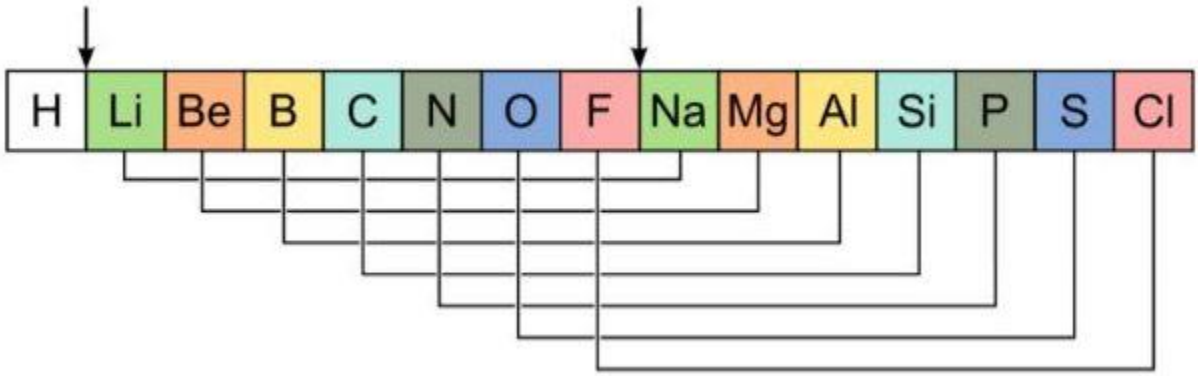
- The identification of new elements made this model obsolete.
- Newly discovered elements did not fit into the triads.
- Only a total of 5 Dobereiner's triads were identified.
- Even several known elements did not fit into any of the triads.

Owing to these shortcomings, other methods of classifying elements were developed.

Newland's Law of Octaves

In the year 1864, the British chemist John Newlands attempted the 62 elements known at that time. He arranged them in an ascending order based on their atomic masses and observed that every 8th element had similar properties. On the basis of this observation, Newland's law of octaves was formulated.

The **law of octaves** states that every eighth element has similar properties when the elements are arranged in the increasing order of their atomic masses. An illustration detailing the elements holding similar properties as per Newland's law of octaves is provided below.



Newlands compared the similarity between the elements to the octaves of music, where every eighth note is comparable to the first. This was the first attempt at assigning an atomic number to each element. However, this method of classifying elements was met with a lot of resistance in the scientific community.

Limitations of Newland's Law of Octaves

The key shortcomings of Newland's law of octaves are listed below.

- Several elements were fit into the same slots in Newland's periodic classification. For example, cobalt and nickel were placed in the same slot.
- Elements with dissimilar properties were grouped together. For example, the halogens were grouped with some metals such as cobalt, nickel and platinum.
- Newland's law of octaves held true only for elements up to calcium. Elements with greater atomic masses could not be accommodated into octaves.
- The elements that were discovered later could not be fit into the octave pattern. Therefore, this method of classifying elements did not leave any room for the discovery of new elements.

Mendeleev's Periodic Table and Law

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

Features of Mendeleev's Periodic Table

● Twelve horizontal rows, which were condensed to 7, known as periods.
● Eight vertical columns known as groups.
● Groups I to VII subdivided into A and B subgroups.
● Group VIII doesn't have any subgroups and contains three elements in each row.
● Elements in the same group exhibit similar properties.

Group →	I	II	III	IV	V	VI	VII	VIII
Oxide Hydride:	R_2O RH	RO RH_2	R_2O_3 RH_3	RO_2 RH_4	R_2O_5 RH_3	RO_3 RH_2	R_2O_7 RH	RO_4
Periods ↓	A B	A B	A B	A B	A B	A B	A B	Transition series
1.	H 1.008							
2.	Li 6.939	Be 9.012	B 10.81	C 12.011	N 14.007	O 15.999	F 18.998	
3.	Na 22.99	Mg 24.31	Al 29.98	Si 28.09	P 30.974	S 32.06	Cl 35.453	
4. First series : Second series	K 39.102 Cu 63.54	Ca 40.08 Zn 65.37	Sc 44.96 Ga 69.72	Ti 47.90 Ge 72.59	V 50.94 As 74.92	Cr 50.20 Se 78.96	Mn 54.94 Br 79.909	Fe Co Ni 55.85 58.93 58.71
5. First series : Second series	Rb 85.47 Ag 107.87	Sr 87.62 Cd 112.40	Y 88.91 In 114.82	Zr 91.22 Sn 118.69	Nb 92.91 Sb 121.75	Mo 95.94 Te 127.60	Tc 99 I 126.90	Ru Rh Pd 101.07 102.91 106.4
6. First series : Second series :	Cs 132.90 Au 196.97	Ba 137.34 Hg 200.59	La 138.91 Tl 204.37	Ht 178.49 Pb 207.19	Ta 180.95 Bi 208.98	W 183.85		Os Ir Pt 190.2 192.2 195.0

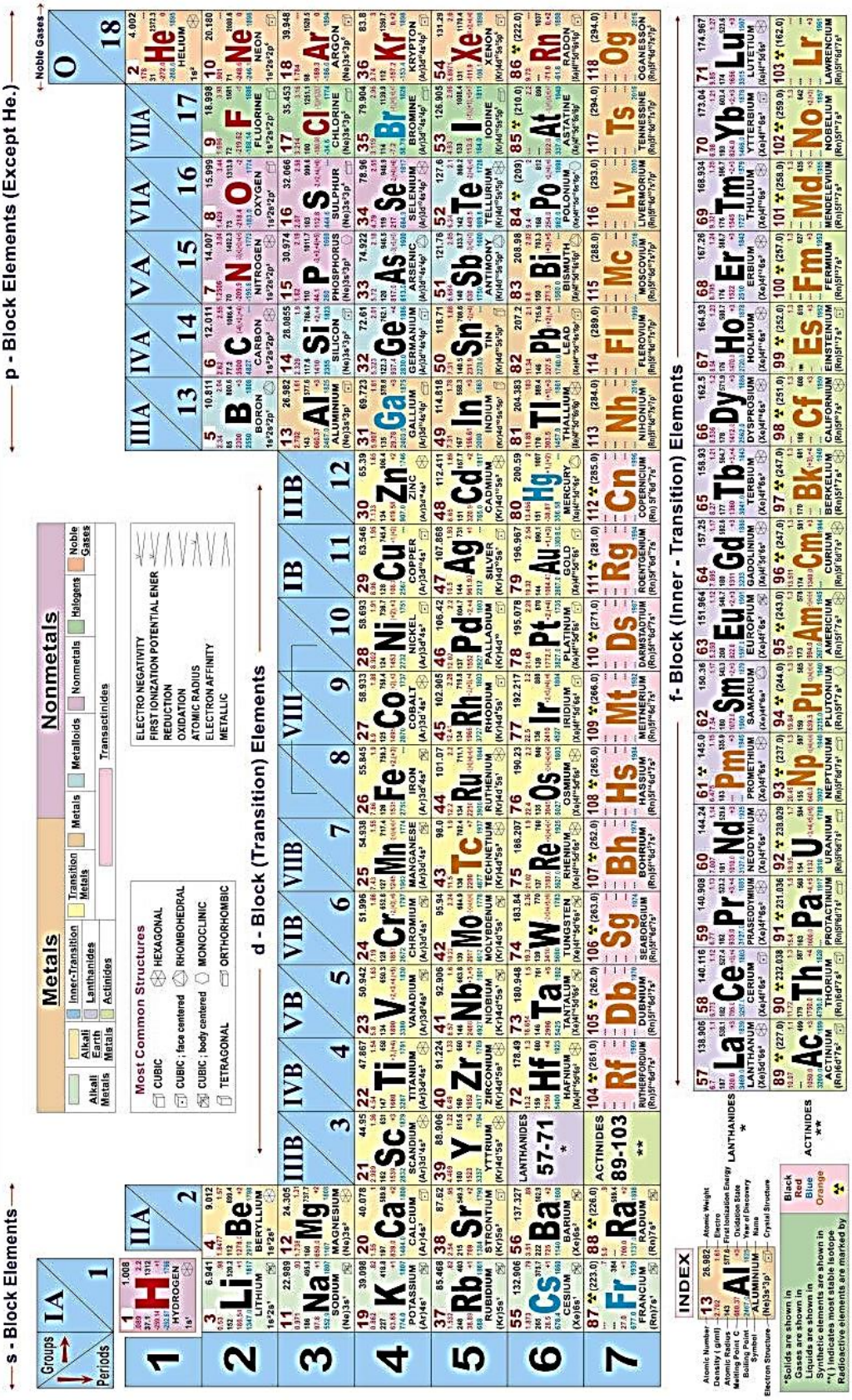
Achievements of Mendeleev's Periodic Table

1. A systematic study of elements: Elements with similar properties were grouped together, that made the study of their chemical and physical properties easier.
2. Correction of atomic masses: Placement of elements in Mendeleev's periodic table helped in correcting the atomic masses of certain elements. For example, the atomic mass of beryllium was corrected from 13.5 to 9. Similarly, atomic masses of indium, gold, platinum etc., were also corrected.
3. Prediction of properties of yet to be discovered elements: Eka-boron, eka-aluminium and eka-silicon were the names given to yet to be discovered elements. The properties of these elements could be predicted accurately from the elements that belonged to the same group. These elements, when discovered were named scandium, gallium, and germanium, respectively.
4. Placement of noble gases: When discovered, they were placed easily in a new group called zero group of Mendeleev's table, without disturbing the existing order.

Limitations of Mendeleev's Periodic Table

1. Position of hydrogen: Hydrogen resembles both, the alkali metals (IA) and the halogens (VIIA) in properties, so, Mendeleev could not justify its position.
2. Position of isotopes: Atomic weight of isotopes differ, but, they were not placed in different positions in Mendeleev's periodic table.
3. Anomalous pairs of elements: Cobalt (Co) has higher atomic weights but was placed before Nickel (Ni) in the periodic table.
4. Placement of like elements in different groups: Platinum (Pt) and Gold (Au) have similar properties but were placed in different groups.
5. Cause of periodicity: He could not explain the cause of periodicity among the elements.

The Modern Periodic Table: In 1913, Henry Moseley showed that the atomic number of an element is a more fundamental property than its atomic mass.



Modern Period Law: The physical and chemical properties of elements are the periodic function of their atomic number.
Modern periodic table is based on atomic number of elements.
Atomic number (Z) is equal to the number of protons present in the nucleus of an atom of an element.
Modern periodic table contains 18 vertical column known as group and seven horizontal rows known as periods.
On moving from left to right in a period, the number of valence electrons increases from 1 to 8 in the elements present.
On moving from left to right in a period, number of shell remains same.
All the elements of a group of the periodic table have the same number of valence electrons.

Trends in Modern Periodic Table: Valency, Atomic size, metallic and non-metallic characters, and Electronegativity.
(i) Valency: The valency of an element is determined by the number of valence electrons present in the outermost shell of its atom (i.e. the combining capacity of an element is known as its valency).
In Period: On moving from left to right in a period, the valency first increases from 1 to 4 and then decreases to zero (0).

Example ; Valency of 2nd period elements are :

	Li	Be	B	C	N	O	F	Ne
Valency	1	2	3	4	3	2	1	0

In Groups: On moving from top to bottom in a group, the valency remains same because the number of valence electrons remains the same.
Example: Valency of first group elements = 1 Valency of second group elements = 2.

(ii) Atomic size: Atomic size refers to radius of an atom. It is a distance between the centre of the nucleus and the outermost shell of an isolated atom.
In Period : On moving from left to right in a period, atomic size decreases because nuclear charge increases.
Example: Size of second period elements: Li > Be > B > C > N > O > F
Point to know: The atomic size of noble gases in corresponding period is largest due to presence of fully filled electronic configuration (i.e. complete octet).
In Group: Atomic size increases down the group because new shells are being added in spite of the increase in nuclear charge.
Example ; Atomic size of first group element : Li < Na < K < Rb < Cs < Fr
Atomic size of 17th group elements : F < Cl < Br < I

(iii) **Metallic character**: It is the tendency of an atom to lose electrons. In Period: Along the period from left to right, metallic characters decreases because a tendency to lose electron decreases due to the increase in nuclear charge. Example: Metallic character of second period elements: Li > Be > B > C >> N > O > F
In Group: Metallic character, when moving from top to bottom increases because the atomic size and tendency to lose electrons increases.
Example: First group element : Li < Na < K < Rb < Cs

(iv) Non-metallic character: It is tendency of an atom to gain electrons.
In Period: Along the period from left to right, non-metallic character increases because tendency to gain electrons increases due to increase in nucleus charge. Example ; Non-metallic character of 2nd period elements : Li < Be < B < C < N < O < F In Group: On moving from top to bottom in a group, non-metallic character

decreases because atomic size increases and tendency to gain electrons decreases. Ex. Non-metallic character of 17th period element: F > Cl > Br > I

(v) Chemical Reactivity
In metals: Chemical reactivity of metals increases down the group because tendency to lose electrons increases. Example ; Li < Na < K < Rb < Cs (1st group) In non-metals: Chemical reactivity of non-metals decreases down the group because tendency to gain electrons decreases. Example: F > Cl > Br > I (17th group)

(vi) Electronegativity: It is tendency of an element to attract the shared pair of electrons towards it in a covalently bonded molecule. It increases with increase of nuclear charge or decrease in atomic size.
Along the period electronegativity increases. Example ;Li < Be < B < C < N < O < F. Down the group electronegativity decreases. Example ; Li > Na > K > Rb > Cs
F > Cl > Br > I

(vii) Nature of Oxides: Metal oxides are basic in nature. Ex. Na2O, MgO etc.
Non-metal oxides are acidic in nature. Ex. Cl_2O_7, SO_3, P_2O_5,

In the case of metal reactivity, it increases down the group because of the tendency to lose electrons increases.

In the case of non-metal reactivity, decreases down the group because of the tendency to gain electrons decreases.