

NEET Revision Notes Chemistry

Classification of Elements and Periodicity in Properties

Periodic Classification

Periodic classification or periodic arrangement of elements refers to the arrangement of elements in such a way that elements with similar properties reappear at regular intervals.

Several people have tried to classify elements into groups, which resulted in the development of periodic classification and periodic table of elements.

Triads of Dobereiner

Dobereiner discovered in 1817 that elements could be arranged in a group of three called triads in such a way that the middle element had an atomic weight nearly equal to the average of the other two.

For example:

Elements Li, Na, K

Atomic mass 7, 23, 39

The average mean of atomic masses is shown below.

$$\frac{7 + 39}{2} = 23$$

Newland's Law of Octaves

John A.R. Newland (1866) discovered that arranging the elements in increasing order of atomic weight, the properties of the first eight elements were similar to those of the first one. It was known as the law of octaves by Newland.

Lothar Meyer's Arrangement of Elements

- Lothar Meyer (1869) stated that the properties of the elements were the periodic functions of their atomic volumes after plotting various physical properties such as atomic volume against atomic weight. This is also known as Lothar Meyer volume curves.
- He ordered the elements in a similar way to Mendeleev's.
- Alkaline earth metals (Be, Mg, Ca, Sr, Ba, etc.) are less electropositive than alkaline earth metals. Take up positions on the curve's descending portion.

- The ascending part of the curve is occupied by halogens and noble gases (excluding helium).
- Since transition elements have small volumes, they are found near the bottoms of the curve.

Mendeleev's Periodic Table

The periodic table of Mendeleev is based on atomic weight.

According to Mendeleev's periodic law, "the properties of elements are the periodic functions of their atomic masses."

- It has vertical columns known as groups and horizontal rows known as periods.
- Mendeleev arranged 63 elements in this table, which had 8 groups and 7 periods.

Defects of Mendeleev's Periodic Table

- The position of hydrogen in the periodic table was incorrectly defined.
- The increasing order of atomic masses was not followed in a certain pair of elements.
- Some similar elements were separated and dissimilar elements were grouped together in the periodic table.
- Mendeleev was unable to explain the cause of periodicity among the elements.

Modern Periodic Table

Moseley, a British chemist, established it in 1913 on the basis of his discovery that atomic number is the most fundamental property. It is a tabular representation of modern periodic law, which states that "the physical and chemical properties of the elements are the periodic function of their atomic numbers."

Characteristics of the Periodic Table in Long Form

- The horizontal rows are referred to as periods, and the vertical columns are referred to as groups.
- The groups are numbered 1 through 18, and there are a total of 7 periods.

- Elements in their atoms with similar outer electronic configurations are arranged in vertical columns, i.e., groups or families. As a result, elements within a group have similar chemical properties.
- The period number corresponds to the highest principal quantum number (n) of the elements in the period, and each period represents the filling of a new electronic shell.
- The first period has two elements, and the subsequent periods have 8, 8, 18, 18 and 32 elements respectively and seventh period is incomplete.
- In this version of the periodic table, 14 elements from the sixth and seventh periods (lanthanoids and actinoids, respectively) are grouped together in separate panels at the bottom.

Classification of Elements

- The periodic classification is theoretically supported by the electronic configuration of atoms.
- Elements in a group have similar chemical properties because their outermost shells have the same electronic configuration.
- The elements are classified into four blocks based on the type of orbital that receives the valence electron.

s-Block elements

- The general electronic configuration of *s*-block elements is ns^{1-2} .
- The elements whose last electron enters the *s*-orbital of their outermost energy level are known as *s*-block elements, and these elements are also known as representative elements.
- The *s*-block elements are found at the far left of the periodic table.
- It is made up of elements from groups IA and IIA. Elements in the IA group are known as alkali metals, while those in the IIA group are known as alkaline earth metals.

Properties of the *s*-block elements

- They are all metals that are malleable, ductile, and good conductors of both heat and electricity. They have a low ionisation potential.
- They are soft metals with a weak metallic bond and the melting and boiling points are both low.

- With the exception of Li and Be, they form ionic compounds (due to Li and Be, due to their small sizes, form covalent compounds in accordance with the Fajan's rule).
- Compounds' ionic nature increases from top to bottom.
- IA elements form monovalent cations, whereas IIA elements form divalent cations because of the loss of electrons from the outermost shell.
- They have a constant oxidation state and form diamagnetic (no unpaired electrons) and colorless salts, with the exception of dichromates and permanganates, which are colored.
- Because of their high reactivity, alkali metals are stored in liquid form.

***p*-block elements**

The general electronic configuration of *p*-block elements is ns^2np^{1-6} .

p-block elements are those in which the final electron enters the *p*-orbital of the outermost energy levels. These elements can be found on the right side of the periodic table. The elements of groups 13 to 18 are contained within the *p*-block.

Properties of *p*-block elements:

- Metals, nonmetals, and metalloids are all examples of *p*-block elements.
- The atomic radii of these elements are smaller than those of *s*-block elements. Because of the increase in nuclear charge, the atomic radius decreases from left to right over time.
- They have a lot of electronegativity. Because of their smaller atomic size, *p*-block elements have greater electron affinities than *s*-block elements.
- Reactive *p*-block elements include halogens, oxygen, sulfur, and phosphorus.
- They usually combine to form covalent compounds.
- Some of these elements have variable valency and can be found in multiple oxidation states in their compounds.
- Only a few elements, such as oxygen, sulfur, and phosphorus, exhibit allotropy.
- They have a high ionisation energy value.
- Their oxides are naturally acidic.

***d*-block elements**

The general electronic configuration of d -block elements is $(n-1)d^{1-10}ns^{1-2}$.

The element in which the last electron enters the d -orbitals, i.e. $(n-1)$ energy level, is referred to as the d -block element. In the periodic table, they are classified as groups 3 to 12.

The d -block is divided into three series:

- (i) the first transition series scandium ($z = 21$) to zinc ($z = 30$);
- (ii) the second transition series yttrium ($z = 39$) to cadmium ($z = 48$);
- (iii) the third transition series lanthanum ($z = 57$); hafnium ($z = 72$) to mercury ($z = 80$).

Properties of d -block elements:

- They are metals with extremely high melting and boiling points (strong metallic bond).
- Because of the free and mobile electrons, they are good conductors of heat and electricity. Silver is the best heat conductor, while lead is the worst.
- In their compounds, almost all of them have variable valence and exist in multiple oxidation states.
- They combine to form complex compounds.
- Transition elements and their derivatives serve as catalysts.
- They can combine to form both ionic and covalent compounds.
- d -block elements have a much higher density than s -block elements.
- Because the last two orbits are incomplete, the properties of transition elements do not change gradually as those of s and p -block elements when moving across a period from left to right.
- Because the majority of these ions contain unpaired electrons, they are paramagnetic and colored (due to the $d-d$ transition).
- Colored salts are not formed by zinc, cadmium, or mercury with d^{10} configuration.

f -block elements

- The general electronic configuration of f -block elements is $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$. The elements in which the final electron enters the atom's f -orbitals are referred to as f -block elements (penultimate orbit).
- The first series (lanthanide series) follows lanthanum from $z = 58$ to $z = 71$. This is also referred to as the $4f$ inner-transition series.

- The second series (actinide series) follows actinium from $z = 90$ to $z = 103$, also known as the $5f$ inner-transition series.

Properties of f -block elements

- All f -block elements are metals and are highly reactive, with high melting and boiling points.
- They have varying valency. Their ions are colored and have a paramagnetic nature.
- They combine to form complex compounds.
- They contain radioactive material.
- Because of their low ionisation energies, these metals are highly electropositive and usually form ionic compounds.

Periodic properties

With the change in atomic number, there is a regular gradation (gradual increase or decrease in a particular property) in the physical and chemical properties of elements within a period and within a group. Periodicity refers to the regular gradation of properties. Periodicity in properties is caused by the repetition of similar configurations at regular intervals.

Effective nuclear charge:

(a) In a polyelectronic atom, the inner electrons repel the outermost orbit electrons. The nuclear attraction on the electrons in the outermost orbit is reduced as a result of this.

(b) As a result, only a portion of the nuclear charge has an influence on electrons in the outermost orbit. As a result, the inner electrons shield or protect the nucleus, reducing the effect of nuclear charge on the electrons in the outermost orbit.

(c) An effective nuclear charge is the portion of the nuclear charge that operates against outside electrons.

$$Z^* = Z - \sigma$$

Here, Z is the nuclear charge, σ is the shielding constant and Z^* is the effective nuclear charge.

Slater, a scientist, calculated the value of the shielding constant and proposed the following rules.

(1) Each electron in the $1s$ orbital has a shielding or screening effect of 0.30.

(2) Each ns and np electron, i.e. the electron in the outermost orbit, has a shielding effect of 0.35.

(3) Each electron in the s , p , or d orbitals of the penultimate orbit ($n-1$) has a shielding effect of 0.85.

(4) Each electron in the s , p , d , or f orbital of the prepenultimate orbit ($n-2$) and below has a shielding effect of 1.0.

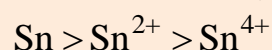
Atomic Radius:

It is defined as the distance from the nucleus's center to the outermost shell of electrons in the case of a covalent bond; in the case of an ionic bond, the radius is ionic radius; and in the absence of a bond, the radius is van der Waal's radius. In general, van der Waal's radius is greater than covalent radius.

Atomic radii of elements increase when moving down the group because the number of shells increases by a factor of one, reducing effective nuclear charge.

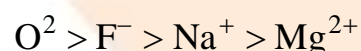
Atomic radii of elements decrease when moving left to right in a period because electrons enter the same shell throughout the period, increasing effective nuclear charge.

Because of the increase in effective nuclear charge per electron, the radius of a cation is always smaller than that of a neutral atom. For example:



Because of the decrease in effective nuclear charge per electron, the radius of an anion is always larger than that of its neutral atom. For example: $\text{O}^{2-} > \text{O}^- > \text{O}$

Species that are isoelectronic. Both of these species have the same number of electrons. In the case of isoelectronic species, the ionic radii decrease as the atomic number increases. For example:



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Ionization Potential (IP):

- It is the smallest amount of energy required to remove an electron from an isolated gaseous atom's outermost orbit.
- IP assesses the proclivity of cation formation.
- IP increases from left to right as the effective nuclear charge increases, while it decreases from top to bottom.

$$IP \propto \frac{1}{\text{size of atom}}$$

IP_1 of elements of group 2 (alkaline earth metal) is greater than that of group 13. For example: IP_1 of Magnesium is greater than IP of Aluminum.

It is due to the group 2 element's stable configuration (ns^2). Similarly, IP_1 of group 15 elements is greater than IP_1 of group 16 elements. For example: IP_1 of nitrogen is greater than that of oxygen.

Because of the stable configuration, zero group elements have an exceptionally high IP value (highest in its period).

IP decreases in order within the same orbit

$$s > p > d > f$$

Electron affinity (EA):

- It is the energy released when an extra electron is added to a gaseous atom in isolation. It assesses the proclivity of anion formation.
- EA rises from left to right in a period and falls from top to bottom in a group.
- The EA of zero group elements is very low.
- Because of its small size, fluorine has less EA than chlorine. F has more electron density and thus greater electron electron repulsion.
- The order of EA of halogens is $Cl > F > Br > I$.

Electronegativity:

Electronegativity is an atom's proclivity to attract the shared pair of electrons to itself. The decreasing order of electronegativity is

$$F > O > Cl \approx N > Br > C \approx I > H$$

$$43.5332 \ .82 \ .52 \ .52 \ .1$$

Note: Trends in periodic properties

Periodic property	In a period from left to right	In a group from top to bottom
Ionisation energy	Increase	Decrease

Electron affinity	Increase	Decrease
Electronegativity	Increase	Decrease
Non-metallic character	Increase	Decrease
Oxidising character	Increase	Decrease
Acidic nature of oxides	Increase	Decrease
Atomic size	Decrease	Increase
Electropositivity	Decrease	Increase
Metallic character	Decrease	Increase
Basic nature of oxides	Decrease	Increase