

UNIT 7

PERIODIC TABLE

The periodic table is a list of all known elements arranged in order of increasing atomic number, from 1 to 106. In addition to this, the elements are arranged in such a way that atoms with the same number of shells are placed together, and atoms with similar electronic configurations in the outer shell are also placed together. This is achieved as follows:

The elements are arranged in rows and columns. Elements with one shell are placed in the first row (ie H and He), Elements with two shells are placed in the second row (Li to Ne) and so on.

A row of elements thus arranged is called a **period**.

In addition, the elements are aligned vertically (in columns) with other elements in different rows, if they share a similar outer-shell electronic configuration. For example, elements with outer-shell configuration ns^1 are all placed in the same column (Li, Na, K, Rb, Cs, Fr).

A column of elements thus arranged is called a **group**.

According to these principles, the periodic table can be constructed as follows:

I	II											III	IV	V	VI	VII	0			
		H																He		
Li	Be											B	C	N	O	F	Ne			
Na	Mg											Al	Si	P	S	Cl	Ar			
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr			
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe			
Cs	Ba	La	Ce - Lu			Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Th - Lw																	

Since the electronic configurations of H and He are unusual, they do not fit comfortably into any group. They are thus allocated a group based on similarities in physical and chemical properties with other members of the group.

He is placed in group 0 on this basis, but hydrogen does not behave like any other element and so is placed in a group of its own.

The elements Ce - Lu and Th - Lw belong in the periodic table as shown above. However if they are placed there periods 6 and 7 do not fit onto a page of A4, so they are placed below the other elements in most tables.

The distance from the center of an atom's nucleus to its outermost energy level is its atomic radius. An atom's radius is determined by the size of the electron cloud that surrounds it.

Moving from top to bottom (moving down a group), with each new energy level, the size of the radius increases.

Generally, the radii of atoms decrease in size as you move from left to right on each period. Even though one electron is added to each successive atom (moving left to right on the table), the atomic radius still gets smaller, or "tighter." This happens because of atoms' strong nuclear charge. Each additional proton pulls even tighter to "pull electrons in" closer to the nucleus. Even though the charges are opposite and equal, a proton has almost 2x the attractive power of an electron.

The shell where an electron is added (*s, p, d* or *f*) also affects the atomic radius. When electrons are added to the *s* or *p* in the outermost shell the nuclear charge [positive charge from protons in the nucleus] pulls on the outer shell the most.

When electrons are added to a *d* or *f* in the inner shell, the outermost electrons have a 'shield' from the positive nuclear charge, so they are not as affected as inner electrons.

To summarize :

In general, atomic radius decreases from left to right along a period and it increases when moving from top to bottom down a group.

INCREASING ATOMIC RADIUS

← INCREASING ATOMIC RADIUS																							
1 H Hydrogen 1.00794																	2 He Helium 4.003						
3 Li Lithium 6.941	4 Be Beryllium 9.012182																	5 B Boron 10.811	6 C Carbon 12.0107	7 N Nitrogen 14.00674	8 O Oxygen 15.9994	9 F Fluorine 18.9984032	10 Ne Neon 20.1797
11 Na Sodium 22.989770	12 Mg Magnesium 24.3050																	13 Al Aluminum 26.981538	14 Si Silicon 28.0855	15 P Phosphorus 30.973761	16 S Sulfur 32.066	17 Cl Chlorine 35.4527	18 Ar Argon 39.948
19 K Potassium 39.0983	20 Ca Calcium 40.078	21 Sc Scandium 44.955910	22 Ti Titanium 47.867	23 V Vanadium 50.9415	24 Cr Chromium 51.9961	25 Mn Manganese 54.938049	26 Fe Iron 55.845	27 Co Cobalt 58.933200	28 Ni Nickel 58.6934	29 Cu Copper 63.546	30 Zn Zinc 65.39	31 Ga Gallium 69.723	32 Ge Germanium 72.61	33 As Arsenic 74.92160	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.80						
37 Rb Rubidium 85.4678	38 Sr Strontium 87.62	39 Y Yttrium 88.90585	40 Zr Zirconium 91.224	41 Nb Niobium 92.90638	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.90550	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.411	49 In Indium 114.818	50 Sn Tin 118.710	51 Sb Antimony 121.760	52 Te Tellurium 127.60	53 I Iodine 126.90447	54 Xe Xenon 131.29						
55 Cs Cesium 132.90545	56 Ba Barium 137.327	57 La Lanthanum 138.9055	72 Hf Hafnium 178.49	73 Ta Tantalum 180.9479	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.217	78 Pt Platinum 195.078	79 Au Gold 196.96655	80 Hg Mercury 200.59	81 Tl Thallium 204.3833	82 Pb Lead 207.2	83 Bi Bismuth 208.98038	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)						
87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (263)	107 Bh Bohrium (262)	108 Hs Hassium (268)	109 Mt Meitnerium (266)	110 (269)	111 (272)	112 (277)												

Forming Ions

An ion is an atom that has lost or gained electrons.

An atom that loses electrons is positively charged and is called a cation.

An atom that gains extra electrons is negatively charged and is called an anion.

The atomic radius of a cation is smaller than the parent atoms radius, because they have fewer electrons to pull against the positive nuclear charge, With fewer electrons, the positive nuclear charge pulls the remaining electrons tighter, making a smaller atomic radius.

The atomic radius of anions is increased compared to the parent atom. Excess electrons help shield other electrons from the positive nuclear charge. They also act to repel other electrons, pushing them away; this makes the electron cloud 'stretch' to a bigger size.

Some atoms easily lose electrons; these are atoms that are unstable until they get rid of some electrons. Other atoms quickly grab extra electrons to become stable.

The ease with which electrons donate or accept electrons (becoming "ionized") is important for predicting how different elements will react chemically.

Ionization Energy

"Ionization energy" is the minimum amount of energy required to remove one electron from the ground state (lowest energy state) of the gaseous-state atom.

The first ionization energy is the amount of energy required to remove the first electron from a neutral atom. For example, the first ionization energy of sodium:



Second ionization energy is the energy required to remove the second electron.



The higher the ionization energy, the more difficult it is to remove an electron; in other words, the less likely the atom will give up an electron.

For every atom, ionization energy increases as each electron is removed, because the atom becomes increasingly more positive causing a stronger pull on the remaining electrons.

Ionization energy depends also on atomic radius. When the radius is bigger (and electrons are further away from the nucleus) the ionization energy is lower.

Generally, ionization energy increases going left to right along each row and it decreases as you go top to bottom down a group. (This is the opposite of the trend for atomic radii)

Factors affecting: -

1. Size of the positive nuclear charge

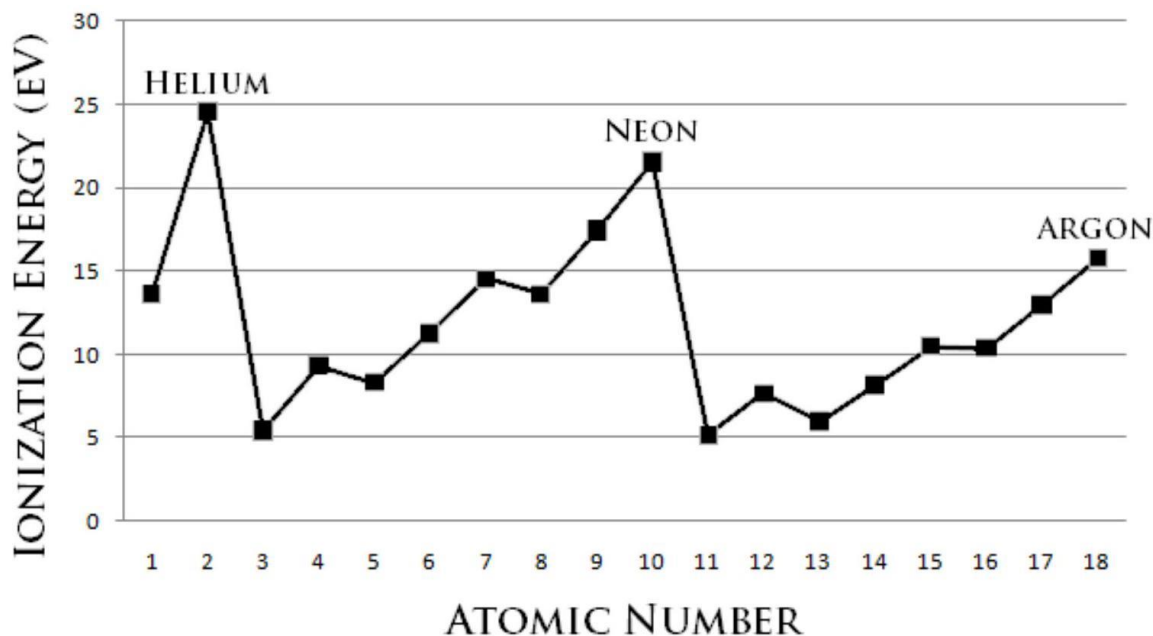
As the nuclear charge increases, its attraction for the outermost electron increases and more energy is required to remove an electron. This means that the ionization energy increases.

2. Size of atom (distance of outermost electron from the nucleus)

As atomic size increases, the attraction of the positive nucleus for the negative electron decreases and less energy is required to remove an electron. This means that the ionization energy decreases.

3. Screening (shielding) effect of inner shell electrons

The outermost electron is screened (shielded) from the attraction of the nucleus by the repelling effect of the inner electrons. As shielding increases, the attraction of the positive nucleus for the negative electron decreases and less energy is required to remove an electron. This means that the ionization energy decreases.



Electron Affinity

Affinity: "a natural liking for, or attraction to a thing."

The energy involved in the addition of an electron to an atom is its electron affinity. This energy is a measure of the attraction (or "affinity") of the atom for the added electron.

Usually, when an electron is added (creating an anion), energy is released, so the value is negative.

Example of electron affinity for Chlorine: $\text{Cl}(\text{g}) + \text{e}^- \rightarrow \text{Cl}^- (\text{g}) \Delta E = -349 \text{ J}$

In general, electron affinity increases from left to right along a period and decreases from top to bottom in a group. (Similar to ionization energy and opposite of atomic radii)

Exception: This is NOT true for noble gases. They do not want extra electrons, because that takes more energy.

Remember: Ionization energy is a measure of the energy required to remove an electron while electron affinity is a measure of the energy involved in adding an electron.

Electro negativity: The tendency for an atom to attract electrons to itself when forming a bond with another element.

The vertical and horizontal trends for electro negativity (and their explanations) are the SAME as those for ionization energy. There is an exception: fluorine, not helium has the highest electro negativity on the periodic table and similarly the halogens are the group with the highest electro negativity, not the noble gases. This is because the noble gases have a full outer

shell of electrons and do not tend to attract additional electron density The halogens, on the other hand strongly attract electrons in order to fill their outer shell (remember that they have seven valence electrons, but would be most stable with eight)

- **From left to right across a period of elements, electro negativity increases.** If the valence shell of an atom is less than half full, it requires less energy to lose an electron than to gain one. Conversely, if the valence shell is more than half full, it is easier to pull an electron into the valence shell than to donate one.
- **From top to bottom down a group, electro negativity decreases.** This is because atomic number increases down a group, and thus there is an increased distance between the valence electrons and nucleus, or a greater atomic radius.
- **Important exceptions of the above rules include the noble gases, lanthanides and actinides.** The noble gases possess a complete valence shell and do not usually attract electrons. The lanthanides and actinides possess more complicated chemistry that does not generally follow any trends. Therefore, noble gases, lanthanides, and actinides do not have electro negativity values.
- **As for the transition metals, although they have electro negativity values, there is little variance among them across the period and up and down a group.** This is because their metallic properties affect their ability to attract electrons as easily as the other elements.

According to these two general trends, the most electronegative element is fluorine, with 3.98 Pauling units.

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Figure . Periodic Table showing Electro negativity Trend

Effective nuclear charge - The attractive positive charge of nuclear protons acting on valence electrons.

1. The effective nuclear charge is always less than the total number of protons present in a nucleus due to shielding effect.
2. Effective nuclear charge is behind all other periodic table tendencies.

Shielding effect - the lessening of attractive electrostatic charge difference between nuclear protons and valence electrons by partially or fully filled inner shells.

1. Shielding effect increases with the number of inner shells of electrons.
2. Electrons sharing the same shell do not shield one another from the attractive pull of the nucleus.

Calculating the effective nuclear charge:

An estimate of effective nuclear charge can be obtained from $Z_{\text{eff}} = Z - S$, where Z_{eff} = effective nuclear charge, Z = atomic number, and, S = the screening constant. Consider aluminum: $[\text{Ne}]3s^23p^1$ $Z = 13$ $S = 10$ $Z_{\text{eff}} = Z - S = 13 - 10 = 3 +$

Don't forget that Z_{eff} is only an estimate. Actual shielding effect is always greater than the screening constant S because core electrons are much closer to the nucleus than are valence electrons.

Trends

The periodic table tendency for effective nuclear charge:

1. Increase across a period (due to increasing nuclear charge with no accompanying increase in shielding effect).
2. Decrease down a group (although nuclear charge increases down a group, shielding effect more than counters its effect).

ELECTRONIC CONFIGURATION

The **electron configuration** is the distribution of electrons of an atom or molecule (or other physical structure) in atomic or molecular orbital. Electronic configurations describe each electron as moving independently in an orbital, in an average field created by all other orbital.

Atomic number	Symbol	Electron configuration	Atomic number	Symbol	Electron configuration	Atomic number	Symbol	Electron configuration
1	H	1s ¹	37	Rb	[Kr]5s ¹	73	Ta	[Xe]6s ² 4f ¹⁴ 5d ³
2	He	1s ²	38	Sr	[Kr]5s ²	74	W	[Xe]6s ² 4f ¹⁴ 5d ⁴
3	Li	[He]2s ¹	39	Y	[Kr]5s ² 4d ¹	75	Re	[Xe]6s ² 4f ¹⁴ 5d ⁵
4	Be	[He]2s ²	40	Zr	[Kr]5s ² 4d ²	76	Os	[Xe]6s ² 4f ¹⁴ 5d ⁶
5	B	[He]2s ² 2p ¹	41	Nb	[Kr]5s ¹ 4d ⁴	77	Ir	[Xe]6s ² 4f ¹⁴ 5d ⁷
6	C	[He]2s ² 2p ²	42	Mo	[Kr]5s ¹ 4d ⁵	78	Pt	[Xe]6s ¹ 4f ¹⁴ 5d ⁸
7	N	[He]2s ² 2p ³	43	Tc	[Kr]5s ² 4d ⁵	79	Au	[Xe]6s ¹ 4f ¹⁴ 5d ¹⁰
8	O	[He]2s ² 2p ⁴	44	Ru	[Kr]5s ¹ 4d ⁶	80	Hg	[Xe]6s ² 4f ¹⁴ 5d ¹⁰
9	F	[He]2s ² 2p ⁵	45	Rh	[Kr]5s ¹ 4d ⁶	81	Tl	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ¹
10	Ne	[He]2s ² 2p ⁶	46	Pd	[Kr]4d ¹⁰	82	Pb	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ²
11	Na	[Ne]3s ¹	47	Ag	[Kr]5s ¹ 4d ¹⁰	83	Bi	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ³
12	Mg	[Ne]3s ²	48	Cd	[Kr]5s ² 4d ¹⁰	84	Po	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁴
13	Al	[Ne]3s ² 3p ¹	49	In	[Kr]5s ² 4d ¹⁰ 5p ¹	85	At	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁵
14	Si	[Ne]3s ² 3p ²	50	Sn	[Kr]5s ² 4d ¹⁰ 5p ²	86	Rn	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁶
15	P	[Ne]3s ² 3p ³	51	Sb	[Kr]5s ² 4d ¹⁰ 5p ³	87	Fr	[Rn]7s ¹
16	S	[Ne]3s ² 3p ⁴	52	Te	[Kr]5s ² 4d ¹⁰ 5p ⁴	88	Ra	[Rn]7s ²
17	Cl	[Ne]3s ² 3p ⁵	53	I	[Kr]5s ² 4d ¹⁰ 5p ⁵	89	Ac	[Rn]7s ² 6d ¹
18	Ar	[Ne]3s ² 3p ⁶	54	Xe	[Kr]5s ² 4d ¹⁰ 5p ⁶	90	Th	[Rn]7s ² 6d ²
19	K	[Ar]4s ¹	55	Cs	[Xe]6s ¹	91	Pa	[Rn]7s ² 5f ⁶ 6d ¹
20	Ca	[Ar]4s ²	56	Ba	[Xe]6s ²	92	U	[Rn]7s ² 5f ⁶ 6d ¹
21	Sc	[Ar]4s ² 3d ¹	57	La	[Xe]6s ² 5d ¹	93	Np	[Rn]7s ² 5f ⁶ 6d ¹
22	Ti	[Ar]4s ² 3d ²	58	Ce	[Xe]6s ² 4f ¹ 5d ¹	94	Pu	[Rn]7s ² 5f ⁶
23	V	[Ar]4s ² 3d ³	59	Pr	[Xe]6s ² 4f ³	95	Am	[Rn]7s ² 5f ⁷
24	Cr	[Ar]4s ¹ 3d ⁵	60	Nd	[Xe]6s ² 4f ⁴	96	Cm	[Rn]7s ² 5f ⁷ 6d ¹
25	Mn	[Ar]4s ² 3d ⁵	61	Pm	[Xe]6s ² 4f ⁶	97	Bk	[Rn]7s ² 5f ⁷
26	Fe	[Ar]4s ² 3d ⁶	62	Sm	[Xe]6s ² 4f ⁶	98	Cf	[Rn]7s ² 5f ¹⁰
27	Co	[Ar]4s ² 3d ⁷	63	Eu	[Xe]6s ² 4f ⁷	99	Es	[Rn]7s ² 5f ¹¹
28	Ni	[Ar]4s ² 3d ⁸	64	Gd	[Xe]6s ² 4f ⁷ 5d ¹	100	Fm	[Rn]7s ² 5f ¹²
29	Cu	[Ar]4s ¹ 3d ¹⁰	65	Tb	[Xe]6s ² 4f ⁹	101	Md	[Rn]7s ² 5f ¹³
30	Zn	[Ar]4s ² 3d ¹⁰	66	Dy	[Xe]6s ² 4f ¹⁰	102	No	[Rn]7s ² 5f ¹⁴
31	Ga	[Ar]4s ² 3d ¹⁰ 4p ¹	67	Ho	[Xe]6s ² 4f ¹¹	103	Lr	[Rn]7s ² 5f ¹⁴ 6d ¹
32	Ge	[Ar]4s ² 3d ¹⁰ 4p ²	68	Er	[Xe]6s ² 4f ¹²	104	Rf	[Rn]7s ² 5f ¹⁴ 6d ²
33	As	[Ar]4s ² 3d ¹⁰ 4p ³	69	Tm	[Xe]6s ² 4f ¹³	105	Db	[Rn]7s ² 5f ¹⁴ 6d ³
34	Se	[Ar]4s ² 3d ¹⁰ 4p ⁴	70	Yb	[Xe]6s ² 4f ¹⁴	106	Sg	[Rn]7s ² 5f ¹⁴ 6d ⁴
35	Br	[Ar]4s ² 3d ¹⁰ 4p ⁵	71	Lu	[Xe]6s ² 4f ¹⁴ 5d ¹	107	Bh	[Rn]7s ² 5f ¹⁴ 6d ⁵
36	Kr	[Ar]4s ² 3d ¹⁰ 4p ⁶	72	Hf	[Xe]6s ² 4f ¹⁴ 5d ²	108	Hs	[Rn]7s ² 5f ¹⁴ 6d ⁶
						109	Mt	[Rn]7s ² 5f ¹⁴ 6d ⁷
						110	Ds	[Rn]7s ² 5f ¹⁴ 6d ⁸
						111	Rg	[Rn]7s ² 5f ¹⁴ 6d ¹⁰

CLASSIFICATION OF ELEMENTS AS s, p, d, f block elements

s block elements :-

The **s-block** is one of four blocks of elements in the periodic table. The element of s- group have a common property. The electron in their most outward electron shell are in the s- orbital. Elements in the s- are in the first two periodic table groups. The elements in group

one are called the alkali metals. The elements in group two are called the alkaline earth metals.

Properties:-

- All of the s- elements are metals (except Hydrogen).
- In general, they are shiny, silvery, good conductors of heat and electricity.
- They lose their valence electrons easily. In fact, they lose their trademark s orbital valence electrons so easily that the s- elements are some of the most reactive elements on the periodic table.
- The elements in group 1, known collectively as the alkali metals (except hydrogen), always lose their one valence electron to make a +1 ion. These metals are characterized by being silvery, very soft, not very dense and having low melting points.
- The elements in group 2, known as the alkaline earth metals (except helium), always lose their two valence electrons to make a +2 ion. Like the alkali metals, the alkaline earth metals are silvery, shiny and relatively soft.
- S- elements are famous for being ingredients in fireworks. The ionic forms of potassium, strontium and barium make appearances in firework displays as the brilliant purples, reds and greens.
- The first element in group one, Lithium, and the first in group two, Beryllium, behave differently to other members of their groups. Their behavior is like the second element of the next group. So lithium is similar to magnesium, and beryllium is similar to aluminum.
- In the periodic table this is known as a 'diagonal relationship'. The diagonal relationship is because of similarities in ionic sizes and charge/radius ratio of the element. The similarity between lithium and magnesium is because of their similar sizes Radii, $\text{Li} = 152\text{pm}$ $\text{Mg} = 160\text{pm}$

character tends to increment down every group while it diminishes as we go from left to right over a period.

- The Atomic Density of elements in p-block increments down the group, this is because of increment in the atom's size down the group. Though it diminishes as we move from left to right over the period, this is because of the lessening in nuclear size of all elements in the p-block over the period.
- The Melting and Boiling points slowly increment down the group in light of the fact that the atomic mass increments down the group and thus the intermolecular forces also increase.

The p-block elements demonstrate a variable oxidation state. The oxidation states increments as we move from left towards right in the periodic table. Boron family (Group 13): - + 3 Carbon family (Group 14): - + 4 Nitrogen family (Group 15): - + 5, Oxygen family (Group 16): - + 6, Halogen family (Group 17): - + 7, Noble gases (Group 18): - + 8

d Block Elements : -

Transition metals are defined as those elements that have (or readily form) partially filled *d* orbitals. The *d*-block elements in groups 3 – 11 are transition elements. The *f* block elements, also called inner transition metals (the lanthanides and actinides), also meet this criterion because the *d* orbital is partially occupied before the *f* orbitals.

The *d*-block elements are divided into the **first transition series** (the elements Sc through Cu), the **second transition series** (the elements Y through Ag), and the **third transition series** (the element La and the elements Hf through Au). Actinium, Ac , is the first member of the fourth transition series, which also includes Rf through Rg.

CHARACTERSTICS

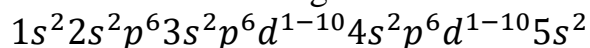
- 1. Metallic nature:** As the number of electrons in the outermost shell is very less i.e. All the transition elements are metals. They show the characteristics of metals such as malleability and ductile in nature and form alloys with several other metals. They also serve as good conductors of heat and electricity. The hard and brittle nature of these elements indicates the presence of covalent bond which is due to the presence of unfilled d-orbital. They are said to form covalent bonding as well as the metallic bonding.
- 2. Melting and boiling points:** They show very high melting and boiling points. This can be attributed to the presence of strong metallic bonding due to the overlapping of ($n - 1$) *d* orbital and covalent bonding of the unpaired d orbital electrons. Since Zn, Cd and Hg have completely filled ($n - 1$)d orbital they are not expected to form covalent bonds. Hence, they show comparatively lower melting point than other d-block elements.

- 3. Atomic radii:** A great degree of variation is seen in the atomic radii across each transition series. The atomic radii of the d -block elements within a given series decreases with increase in the atomic number. This is due to the increase in the nuclear charge that attracts the electron cloud inwards resulting in decrease in size. However, the decrease a uniform decrease in atomic radius is not observed across a period. The decrease in atomic radii is small compared to the S and P block elements. This is due to the screening effect caused by the electrons of the $(n - 1)d$ sub shell on the outermost shell. The atomic radius increases on descending the group. In a given series, the atomic radius decreases to a minimum for the group VIII elements and then it increases towards the end of the series. This increase in radius towards the end of the series is due to the force of repulsion among the added electrons. A close similarity is observed in the radii of the elements of the second and third transition series due to the filling of $4f$ subshells.
- 4. Ionic radii:** The ionic radius is similar to the pattern of atomic radii. Thus, for ions of a given charge the radius decreases slowly with increase in atomic number.
- 5. Atomic volume and Densities:** The atomic volume of transition elements is much lower than those of s and p block elements. This is because of the filling of the $(n - 1)d$ orbital that cause an increase in the nuclear charge and pulls the electrons inward. This results in decrease in atomic volume. With the decrease in the atomic volume, the atomic density these elements increases. Osmium is having a maximum density.
- 6. Ionization potentials:** Transition elements have high ionization energy due to their small size. Their ionization potentials lie between those of s and p block elements. Thus, they are less electropositive than the s -block elements. Hence, they do not form ionic compounds readily like the alkali and alkaline earth metals. They also have the ability o form covalent compounds.
The ionization potentials of d -block elements increase as we move across each series from left to right. However, the increase is not as much as in case of S and P blocks elements. This is due to the screening effect caused by the new electrons that are added into the $(n - 1)d$ subshell.
The second ionization energies of the first transition series also increases with the increase in atomic number. However, Cr and Cu are sufficiently higher than those of their neighbors. This is due to their stable electronic configuration.
- 7. Electronic configuration:** The outer electronic configuration remains constant. But, a electron is added to the penultimate shell till the d -sub shell reaches its full capacity.

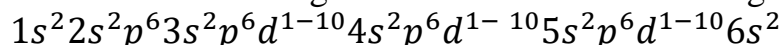
There are three series of elements depending on the $n - 1d$ orbital that is being filled. The orbitals are filled in order of their increasing energy i.e. an orbital of lower energy is filled first. Thus 4s orbital with lesser energy is filled first to its full extent then the 3d orbital with higher energy is filled. The exactly half-filled and completely filled d-orbitals are extra stable.

The electronic configuration of the first series is given as $1s^2 2s^2 p^6 3s^2 p^6 d^{1-10} 4s^2$

The electronic configuration of the second series is given as



The electronic configuration of the third series is given as



Transition elements also show variable oxidation states, tendency to form complexes, magnetic nature and other properties.

- 8. f Block elements :-** These elements have an incomplete f-orbital of the anti-penultimate shell i.e. inner to the penultimate shell in addition to the incomplete d-orbital of the penultimate shell. f-block elements are also termed as inner transition elements as they are transition elements within the transition elements i.e. d-block elements. An f-orbital can accommodate 14 electrons. This means f-block series can be said to include 14 elements. The general electronic configuration of the f-block elements is given as $(n - 2)f^{1-14}, (n - 1)s^2 (n - 1)p^6 (n - 1)d^{10}, ns^2$. The f-block elements are grouped into two series basing on the nature of the f-orbital of the anti-penultimate shell ($4f$ or $5f$) into which the differentiating electron enters.

General Characteristics:

- 1. Electronic Configuration of Lanthanides:** as the 4 f and 5 d electrons are so close in energy it is not possible to decide whether the electron has entered the 5 d or 4 f orbital. However, it is considered that the 5d orbital remains vacant and the electrons enter into the 4f orbital except for gadolinium, $Gd(Z = 64)$ where the electron enters into the 5d orbital due to the presence of half filled d-orbital. At Ytterbium ($z = 70$) all the 4f orbitals are completely filled and hence, the differentiating electron of the next element that is lutetium ($z=71$) enters into the 5d orbital.
- 2. The complete electronic configuration of Lanthanides can be given as**

$$1s^2 2s^2 p^6 3s^2 p^6 d^{10} 4s^2 p^6 d^{10} f^{0-14} 5s^2 p^6 d^{0-1} 6s^2$$
- 3. Oxidation States:** Lanthanides show variable oxidation states but the degree of variability is less compared to the transition elements. The most stable oxidation state of Lanthanides is +3. In addition to the most stable +2 state, Lanthanides also show +2 and +4 oxidation states. These additional oxidation states also show stability due to presence of either half filled or completely filled or empty 4 f subshell. It should be noted that the +2 and +4 oxidation states are unstable in aqueous solutions except for

Ce^{+4} salts such as the ceric sulphate which acts as an oxidising agent in volumetric analysis.

- 4. Ionic Radii and Lanthanide contraction:** the atomic size decreases with the increase in atomic number as we move across from La to Lu. Thus, among lanthanides, lanthanum has the largest atomic radius and lutetium has the smallest atomic radius. This gradual decrease in the size of an atom is said as lanthanide contraction.
- 5. Color:** All lanthanide metals are silver white. The trivalent lanthanide ions are colored both in solid state and in aqueous solution. It should be noted that the color change is seen only in case of cations. The color of a cation depends on the number of unpaired f electrons. Lanthanides with either half-filled or completely filled orbitals are colorless.
- 6. Oxidation States:** Unlike lanthanides which show the +3 oxidation states, actinides show a variety of oxidation states from +3 to +6. However, +3 and +4 are the principal oxidation states. The +3 oxidation state is the most stable in Ac and all the other elements of the series. The +4 oxidation state is the most stable in Th and Pu. +5 in Pa and Np and +6 is seen in U .
- 7. Ionic radius:** In actinides, the ionic radius decreases as we move down the series. This decrease in ionic radius is termed as actinide contraction. This effect is due to poor screening offered by f electrons.
- 8. Complex formation:** Unlike lanthanides, they have the tendency to form complexes. Complex formation with ligands such as thio-ethers is also seen.

POLARIZABILITY:- Polarizability is a measure of how easily an electron cloud is distorted by an electric field. Typically the electron cloud will belong to an **atom** or molecule or ion. The electric field could be caused, for example, by an electrode or a nearby cation or anion.

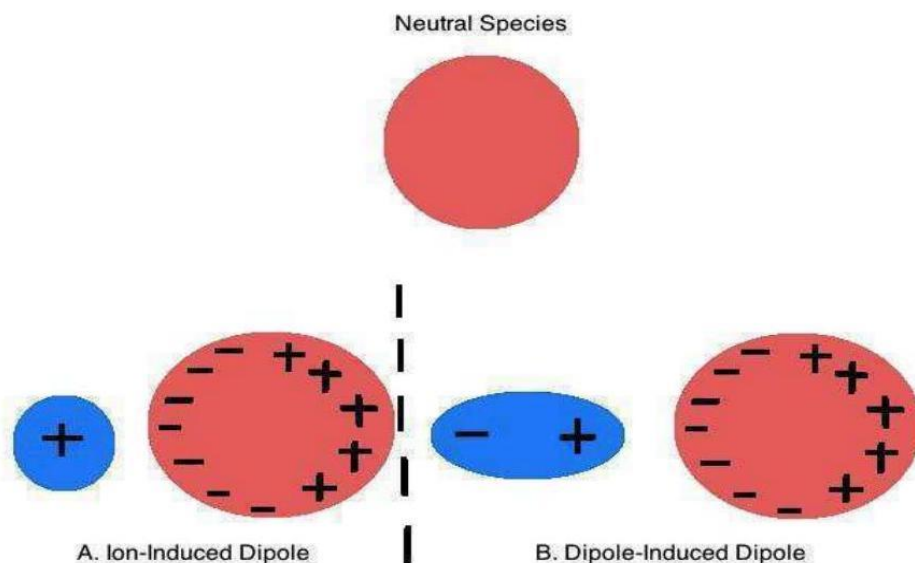
Neutral non polar species have spherically symmetric arrangements of electrons in their electron clouds. When in the presence of an electric field, their electron clouds can be distorted. The ease of this distortion is defined as the **polarizability** of the atom or molecule. The created distortion of the electron cloud causes the originally non polar molecule or atom to acquire a dipole moment. This induced dipole moment is related to the polarizability of the molecule or atom and the strength of the electric field by the following equation:

$$\mu_{ind} = \alpha E \quad (1.1) \quad \mu_{ind} = \alpha E$$

Where

- E denotes the strength of the electric field and

- α is the polarizability constant with units of $\text{Cm}^2 \text{V}^{-1}$.



Factors that Influence Polarizability

The relationship between polarizability and the factors of electron density, atomic radii, and molecular orientation is as follows:

1. The greater the number of electrons, the less control the nuclear charge has on charge distribution, and thus the increased polarizability of the atom.
2. The greater the distance of electrons from nuclear charge, the less control the nuclear charge has on the charge distribution, and thus the increased polarizability of the atom.
3. Molecular orientation with respect to an electric field can affect polarizability (labeled Orientation-dependent), except for molecules that are: tetrahedral, octahedral or icosahedral (labeled Orientation-independent). This factor is more important for unsaturated molecules that contain areas of electron dense regions, such as 2,4hexadiene. Greatest polarizability in these molecules is achieved when the electric field is applied parallel to the molecule rather than perpendicular to the molecule.

Polarizability Influences Dispersion Forces

The dispersion force is the weakest intermolecular force. It is an attractive force that arises from surrounding temporary dipole moments in non polar molecules or species. These temporary dipole moments arise when there are instantaneous deviations in the electron clouds of the non polar species. Surrounding molecules are influenced by these temporary dipole moments and a sort of chain reaction results in which subsequent weak, dipole-induced dipole interactions are created. These cumulative dipole-induced dipole interactions

create the attractive dispersion forces. Dispersion forces are the forces that make non polar substances condense to liquids and freeze into solids when the temperature is low enough.

Polarizability affects dispersion forces in the following ways:

- As polarizability increases, the dispersion forces also become stronger. Thus, molecules attract one another more strongly and melting and boiling points of covalent substances increase with larger molecular mass.
- Polarizability also affects dispersion forces through the molecular shape of the affected molecules. Elongated molecules have electrons that are easily moved increasing their polarizability and thus strengthening the dispersion forces. In contrast, small, compact, symmetrical molecules are less polarizable resulting in weaker dispersion forces.



If you have any queries please visit- <https://studywithakash.in/>

Gmail – studywithakash311@gmail.com

+918871317984

THANK YOU