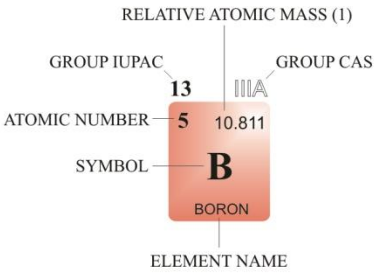
CLASSIFICATION
OF
THE ELEMENTS

PERIODIC TABLE OF THE ELEMENTS

http://www.periodni.com



- 🟦 Metal
- 🟥 Semimetal
- 🟩 Nonmetal

- 🟦 Alkali metal
- 🟦 Alkaline earth metal
- 🟦 Transition metals
- 🟪 Lanthanide
- 🟪 Actinide
- 🟩 Chalcogens element
- 🟩 Halogens element
- 🟩 Noble gas

STANDARD STATE (25 °C; 101 kPa)

Ne - gas   **Fe** - solid

Hg - liquid   **Tc** - synthetic

PERIOD	GROUP 1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9	10	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 1.0079 **H** HYDROGEN																	2 4.0026 **He** HELIUM
2	3 6.941 **Li** LITHIUM	4 9.0122 **Be** BERYLLIUM											5 10.811 **B** BORON	6 12.011 **C** CARBON	7 14.007 **N** NITROGEN	8 15.999 **O** OXYGEN	9 18.998 **F** FLUORINE	10 20.180 **Ne** NEON
3	11 22.990 **Na** SODIUM	12 24.305 **Mg** MAGNESIUM											13 26.982 **Al** ALUMINIUM	14 28.086 **Si** SILICON	15 30.974 **P** PHOSPHORUS	16 32.065 **S** SULPHUR	17 35.453 **Cl** CHLORINE	18 39.948 **Ar** ARGON
4	19 39.098 **K** POTASSIUM	20 40.078 **Ca** CALCIUM	21 44.956 **Sc** SCANDIUM	22 47.867 **Ti** TITANIUM	23 50.942 **V** VANADIUM	24 51.996 **Cr** CHROMIUM	25 54.938 **Mn** MANGANESE	26 55.845 **Fe** IRON	27 58.933 **Co** COBALT	28 58.693 **Ni** NICKEL	29 63.546 **Cu** COPPER	30 65.38 **Zn** ZINC	31 69.723 **Ga** GALLIUM	32 72.64 **Ge** GERMANIUM	33 74.922 **As** ARSENIC	34 78.96 **Se** SELENIUM	35 79.904 **Br** BROMINE	36 83.798 **Kr** KRYPTON
5	37 85.468 **Rb** RUBIDIUM	38 87.62 **Sr** STRONTIUM	39 88.906 **Y** YTTRIUM	40 91.224 **Zr** ZIRCONIUM	41 92.906 **Nb** NIOBIUM	42 95.96 **Mo** MOLYBDENUM	43 (98) **Tc** TECHNETIUM	44 101.07 **Ru** RUTHENIUM	45 102.91 **Rh** RHODIUM	46 106.42 **Pd** PALLADIUM	47 107.87 **Ag** SILVER	48 112.41 **Cd** CADMIUM	49 114.82 **In** INDIUM	50 118.71 **Sn** TIN	51 121.76 **Sb** ANTIMONY	52 127.60 **Te** TELLURIUM	53 126.90 **I** IODINE	54 131.29 **Xe** XENON
6	55 132.91 **Cs** CAESIUM	56 137.33 **Ba** BARIUM	57-71 **La-Lu** Lanthanide	72 178.49 **Hf** HAFNIUM	73 180.95 **Ta** TANTALUM	74 183.84 **W** TUNGSTEN	75 186.21 **Re** RHENIUM	76 190.23 **Os** OSMIUM	77 192.22 **Ir** IRIDIUM	78 195.08 **Pt** PLATINUM	79 196.97 **Au** GOLD	80 200.59 **Hg** MERCURY	81 204.38 **Tl** THALLIUM	82 207.2 **Pb** LEAD	83 208.98 **Bi** BISMUTH	84 (209) **Po** POLONIUM	85 (210) **At** ASTATINE	86 (222) **Rn** RADON
7	87 (223) **Fr** FRANCIUM	88 (226) **Ra** RADIUM	89-103 **Ac-Lr** Actinide	104 (267) **Rf** RUTHERFORDIUM	105 (268) **Db** DUBNIUM	106 (271) **Sg** SEABORGIUM	107 (272) **Bh** BOHRIUM	108 (277) **Hs** HASSIUM	109 (276) **Mt** MEITNERIUM	110 (281) **Ds** DARMSTADTIUM	111 (280) **Rg** ROENTGENIUM	112 (285) **Cn** COPERNICIUM	113 (. . .) **Uut** UNUNTRIUM	114 (287) **Fl** FLEROVIUM	115 (. . .) **Uup** UNUNPENTIUM	116 (291) **Lv** LIVERMORIUM	117 (. . .) **Uus** UNUNSEPTIUM	118 (. . .) **Uuo** UNUNOCTIUM

LANTHANIDE

57 138.91 **La** LANTHANUM	58 140.12 **Ce** CERIUM	59 140.91 **Pr** PRASEODYMIUM	60 144.24 **Nd** NEODYMIUM	61 (145) **Pm** PROMETHIUM	62 150.36 **Sm** SAMARIUM	63 151.96 **Eu** EUROPIUM	64 157.25 **Gd** GADOLINIUM	65 158.93 **Tb** TERBIUM	66 162.50 **Dy** DYSPROSIUM	67 164.93 **Ho** HOLMIUM	68 167.26 **Er** ERBIUM	69 168.93 **Tm** THULIUM	70 173.05 **Yb** YTTERBIUM	71 174.97 **Lu** LUTETIUM

ACTINIDE

89 (227) **Ac** ACTINIUM	90 232.04 **Th** THORIUM	91 231.04 **Pa** PROTACTINIUM	92 238.03 **U** URANIUM	93 (237) **Np** NEPTUNIUM	94 (244) **Pu** PLUTONIUM	95 (243) **Am** AMERICIUM	96 (247) **Cm** CURIUM	97 (247) **Bk** BERKELIUM	98 (251) **Cf** CALIFORNIUM	99 (252) **Es** EINSTEINIUM	100 (257) **Fm** FERMIUM	101 (258) **Md** MENDELEVIUM	102 (259) **No** NOBELIUM	103 (262) **Lr** LAWRENCIUM

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Relative atomic masses are expressed with five significant figures. For elements that have no stable nuclides, the value enclosed in brackets indicates the mass number of the longest-lived isotope of the element. However three such elements (Th, Pa and U) do have a characteristic terrestrial isotopic composition, and for these an atomic weight is tabulated.

Electronic structure and Periodic Law

Types of elements: According to the electronic configurations, the elements may be divided into four types. The four types of elements are:

(1) The Inert Gases (Elements of **0** group).
(2) The Representative elements (**s** and **p** block elements).
(3) The Transition Elements (**d** block elements).
(4) The Inner Transition Elements (**f** block elements).

The Inert Gases: The zero group elements have been placed at the end of each period in the periodic table. It appears that these elements having s^2p^6 electronic arrangements in the outermost level are very stable. Helium has $1s^2$ stable arrangement and all other inert gases have s^2p^6 outer configurations. These elements are colourless gases.

The Representative Elements (s and p block elements): These elements generally belongs to a **A** sub-group of the Periodic Table. These elements have the outermost energy level incomplete just after the complete or stable groupings of s^2p^6.

The chemical behavior of these elements depends upon the valence electrons and these are both metals and non metals. The alkali metals (Group IA) and alkaline earth metals (Group IIA) are **s** block elements. The valence electrons of all the elements from boron to halogens (Group IIIA to VIIA vertically) occupy **p** orbitals. Hence these elements are called **p** block elements. They generally form colourless compounds.

The Transition Elements (d block elements): These elements are generally heavy metals of sub-group **B** and contains two incomplete energy levels because of the building up of the inner **d** electrons. The chemical properties of these elements depend upon the electrons from the two outermost levels (**s** and **d** electrons). These elements generally form coloured compounds.

These elements have normally the same number of electrons in the outermost level but have a progressively greater number of electrons in an inner level (such as **d** level) and hence they are called as "Transition Elements".

The Inner Transition Elements (f block elements): These elements have three incomplete outer levels. The orbital in which the electron is added on increasing the atomic number is an **f** orbital. The inner transition elements (lanthanides and actinides) are all metals and show variable oxidation states. Their compound are highly coloured.

The properties of Transition Elements (d block elements):

The properties of transition elements are summarized in the following points:

1. All the elements are of high melting points, electropositive and heavy metals.
2. These metals have almost the same atomic and ionic sizes. There is only slight increase in the ionization energy of the formation of M^{+2} ions.
3. All these elements show positive oxidation states of +2 and +3 generally and form mostly ionic compounds. Higher oxidations states are also exhibited in some cases.
4. As a general rule, the transition elements form coloured compounds.
5. These elements are also effective catalytic agents.
6. All these form quite a large number of complex compounds.

These properties are due to the influence of the incomplete inner d orbitals in the transition elements.

Variation of Properties within Periods and Groups:

1. **Variation of Metallic Character of the elements:** Generally, it may be noticed that in the periodic table the metallic character of the elements decreases from left to right progressing in the series but increase in moving vertically from top to bottom in the groups. The term “metallic character” is a rough and combination of a number of specific properties, such as electrical and thermal conductivities, metallic lustre, reducing properties etc. Except the transition elements, the trend in the variation of metallic character of elements follows the above generalization. For instance, the most non-metallic elements, fluorine, chlorine, oxygen, sulphur, nitrogen are found at the upper right of the periodic table whereas the most basic metals, the alkali and alkaline earth metals are at the lower left of the table.

2. **Variation in atomic size:** The atomic size in each succeeding element in a period decreases. When the succeeding electrons go into the same energy levels they are subject to greater attraction by the increased nuclear charge and hence the elements in a series show gradual decrease in the atomic size.

   Vertically in the groups the succeeding elements increasing atomic radii. This is due to the fact that the additional electron occupies a new sub-level with a quantum number higher than those of already filled energy levels.

3. **Variation in Ionic Radii:** It is obvious that the size of a positive ion will be less than that of atom from which it is formed. There is considerable decrease in size due to the loss of the outermost electron particularly in the case of alkali metals. Thus in the series, the decrease in the ionic sizes of Na^+, Mg^{+2}, Al^{+3} and Si^{+4} appear to be considerable as compared to the atomic sizes of the parents atoms. It will also be seen that the greater the nuclear charge, the smaller is the ionic radius in a series. In a given group of the periodic table positive ions of succeeding elements have larger ionic radii.

   A simple negative ion formed by the addition of one or more electrons to the outermost energy levels of an atom, is expected to be mach larger than the parent atom. In a given group the negative ions will have larger radii due to larger number of electron levels. The ionic sizes in a series also follow the same trend as in case of positive ions, i.e., the sizes decrease gradually from element to element.

4. **Variation in the ionization potential:** Ionization potential is defined as the energy required to remove the outermost electron from an atom. In general, the greater the nuclear charge of atoms having the same number of electron orbit, the greater the ionization potential. Thus the ionization potential increases in a series and shows decreasing tendency with a group in the periodic classification.

5. **Variation in Electronegativities:** The power of attraction that an atom shows for electron in a covalent bond (electronegativity) also shows periodic variations. The most electronegative elements are found towards the end of the periods. Metal having low electronegativities are found at the beginning of the periods. Thus the alkali metals show gradually decreasing electronegativity values within the group. The halogens are most electronegative elements and the values decrease from fluorine to iodine.

6. **Variation in the Oxidizing and Reducing Power:** Oxidizing substances have tendency to accept electrons and are converted into lower oxidation states. The non-metals at the extreme right of the periodic table tend to act as oxidizing agents in chemical reactions with other substances. Similarly, reducing substances give up electrons and are converted into higher oxidation state during chemical reaction. The reducing power is the highest with the metals at the beginning of the periods. Thus alkali metals have the greatest reducing power. In general, the reducing power of the elements is progressively lower as we pass across the periods and higher, as we go down the groups.

Usefulness of the Periodic Table:

1. Classification of the elements.
2. Prediction of undiscovered elements.
3. Correction of atomic weight.
4. Periodic table in industrial research.

Limitation of the periodic Table:

1. **Position of hydrogen:** The position of hydrogen in the periodic table is left undecided. It has similarities in properties with both the alkali metals and halogens. According to the atomic number or atomic weight, hydrogen should occupy a position just before helium.

   Hydrogen is a gas like fluorine and chlorine and forms compounds like CH_4, SiH_4 which are like CCl_4 and $SiCl_4$ respectively. Even Solid hydrogen is a non-metal resembling iodine. The hydrogen molecule is diatomic like halogen.

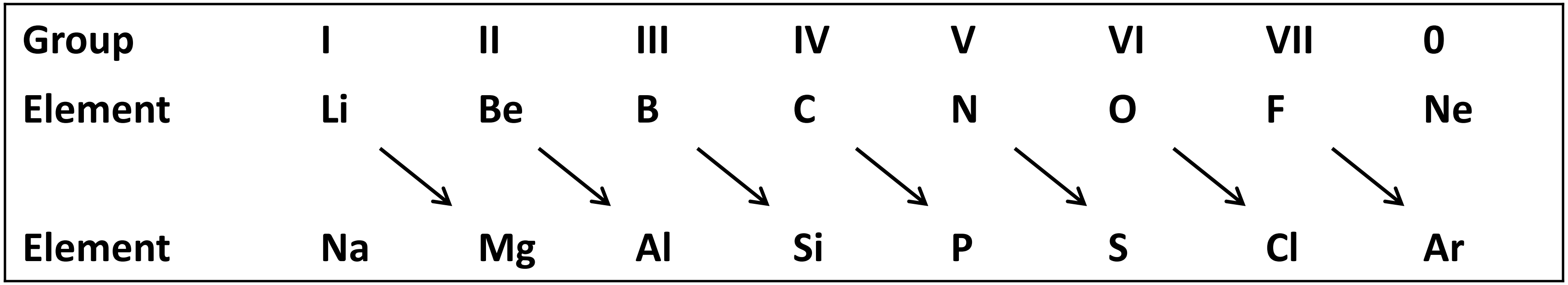
   On the other hand, hydrogen resembles lithium and other alkali metals in having one electron $1s^1$ which can be lost in forming the hydrogen ion. Again, sodium hydride and sodium fluoride are both crystalline ionic solids, a point which shows similarity of hydrogen with fluorine. But in most cases it assumes a +1 oxidation state. For this reason, hydrogen is usually included in group IA of the table

2. **Position of lanthanides (Rare Earths) and Actinides:** The rare earths are also known as lanthanides. Lanthanide series are metals. On the other hand, the elements of actinide series are trans-uranium elements.

   The lanthanides elements have two outermost energy levels identically occupied by electrons which give them great similarities in properties. Similarly, the electronic configurations of the elements of actinides series have been found to very similar to that of lanthanides. Both of these groups of elements contain **f** energy levels which are being systematically filled. Thus, in lanthanide series, cerium to lutccium contains $4f^2$ to $4f^{14}$ and the actinide series of the $5f^2$ to $5f^{14}$ is being completed at lawrencium (103). Their compounds are very closely related to one another which involves tremendous difficulties in their separation. It may be assumed that these elements form a sort of bridge between the preceding and following elements.

   Thus, for the same reason, lanthanide and actinide series are generally omitted from the main table and placed at the bottom of the periodic table.

Diagonal relationships: Lithium of group IA resembles magnesium of group IIA in many respects contrary to its group properties. Similar relationship exists between the some other elements. Beryllium of group IIA shows similarity with aluminum of group IIIA. Boron of group IIIA shows likeness with silicon of group IVA. Thus the light elements of one group shows similarity in properties with the second elements of the following groups. This similarity is generally referred to as diagonal relationship in the periodic table as shown below:



Group	**I**	**II**	**III**	**IV**	**V**	**VI**	**VII**	**0**
Element	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**	**Ne**
Element	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**

The diagonal relationship between the elements may be explained in terms of the electropositive character of the elements. Although an element present in a given group is more electropositive than the corresponding element of the next higher group, the elements more electropositive in passing down the group. Thus, Li in group IA is more electropositive than Be in group IIA, but Mg is also more electropositive than Be. Thus, both Li and Mg are more electropositive than Be and less electropositive than Na.

The other explanation is based on the sizes of the ions formed by the removal of valence electrons. Thus, Li^{+} ion is almost of the same size as Mg^{+2} ion. Similarly, Be^{+2} and Al^{+3} ions have approximately the same ionic size. B^{+3} and Si^{+4} also present the same situation. Compounds having similar properties of the elements showing diagonal relationship are formed due to the effect of ionic sizes.

Thank You