

The Periodic Table and Periodicity

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1. The Periodic Table

Introduction

In this lesson, we will discuss the classification of elements in the Periodic Table. The Periodic Table helps us in the prediction of the properties of new compounds by comparison of the properties of known compounds. The modern Periodic Table is an *arrangement of all the chemical elements in the order of increasing atomic number with elements having similar properties (i.e. of the same chemical family) in the same vertical column*. In this session, we will learn about the group numbers, *s*, *p*, *d* and *f* blocks, and the electronic configuration. Some physical properties of elements such as atomic and ionic sizes/radii, ionization energy, electron affinity, electronegativity, melting and boiling points will be discussed in the next lesson. First of all we will briefly look at the history of the development of the Periodic Table.

1.1 Brief history of the development of the Periodic Table

In the early 19th century many new elements were being discovered and chemists were looking for similarities between these new elements and the existing elements.

Dobereiner (1829) suggested that elements could be grouped in three (triads), in which each member of the triad shows similar properties.

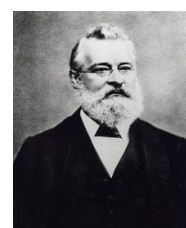
e.g.: Lithium, sodium, potassium,
Calcium, strontium, barium
Chlorine, bromine, iodine



Dobereiner⁷

Newlands (1863) arranged the elements in the order of increasing relative atomic mass. He noticed that there was some similarity between every eighth elements (octet rule).

Li	Be	B	C	N	O	F	
Na	Mg	Al	Si	P	S	Cl	<i>etc</i>



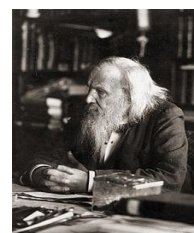
Newlands⁸

These were called **Newlands' Octaves**. Unfortunately the pattern broke down with the heavier elements because he left no gaps (free spaces) for undiscovered elements. His work did not receive much support at that time.



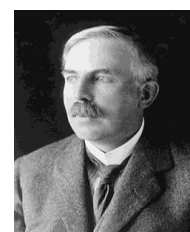
Meyer⁹

Meyer (1869) looked at the relationship between relative atomic mass and the density of an element. He then plotted a graph of atomic volume (mass of 1 mole of atoms divided by density) against the relative atomic mass for each element. The curve he obtained showed periodic variations.



Mendeleev¹⁰

Mendeleev/Mendeleef (1869) arranged the elements in order of increasing relative atomic mass but took into account the pattern of behaviour of the elements. He found that it was necessary to leave gaps in the table and said that these were for elements not known at that time. His table enabled him to predict the properties of the undiscovered elements. His work was proved correct by the accurate prediction of the properties of gallium and germanium. The Periodic Table we use today closely resembles that drawn up by Mendeleev.



Rutherford¹¹

A modification of the Periodic Table was made following the work of **Rutherford** and **Moseley**. It was realized that the elements should be arranged in the order of **atomic number**, *i.e.* the number of **protons** in the nucleus. In the Modern Periodic Table (Figure 1) the elements are arranged in the order of increasing atomic number with elements having similar properties in the same vertical column. Let us study about the Modern Periodic Table.



Moseley¹²

1.2 The Modern Periodic Table

Hydrogen, with the atomic number one, is the first element in the **Periodic Table**. Over one hundred elements are known and they are arranged in the order of increasing **atomic number** (**Z**). Elements are grouped (*e.g.* Group 1, 2, 3,, 17, 18) according to their chemical properties, which largely depend on the *number of electrons* in the *outer/valence shell*. The Periodic Table has vertical columns (**18 Groups**) and horizontal rows (**7 Periods**). The 1st, 2nd, 3rd, 4th, 5th and 6th periods have 2, 8, 8, 18, 18 and 32 elements, respectively. Majority of the elements are **metals** and there are about 17 **non-metals** placed at the right hand side of the Periodic Table. There are seven **metalloids** (B, Si, Ge, As, Sb, Te and At) which show

properties of both metals and non-metals. We will study about the groups and Group numbers now.

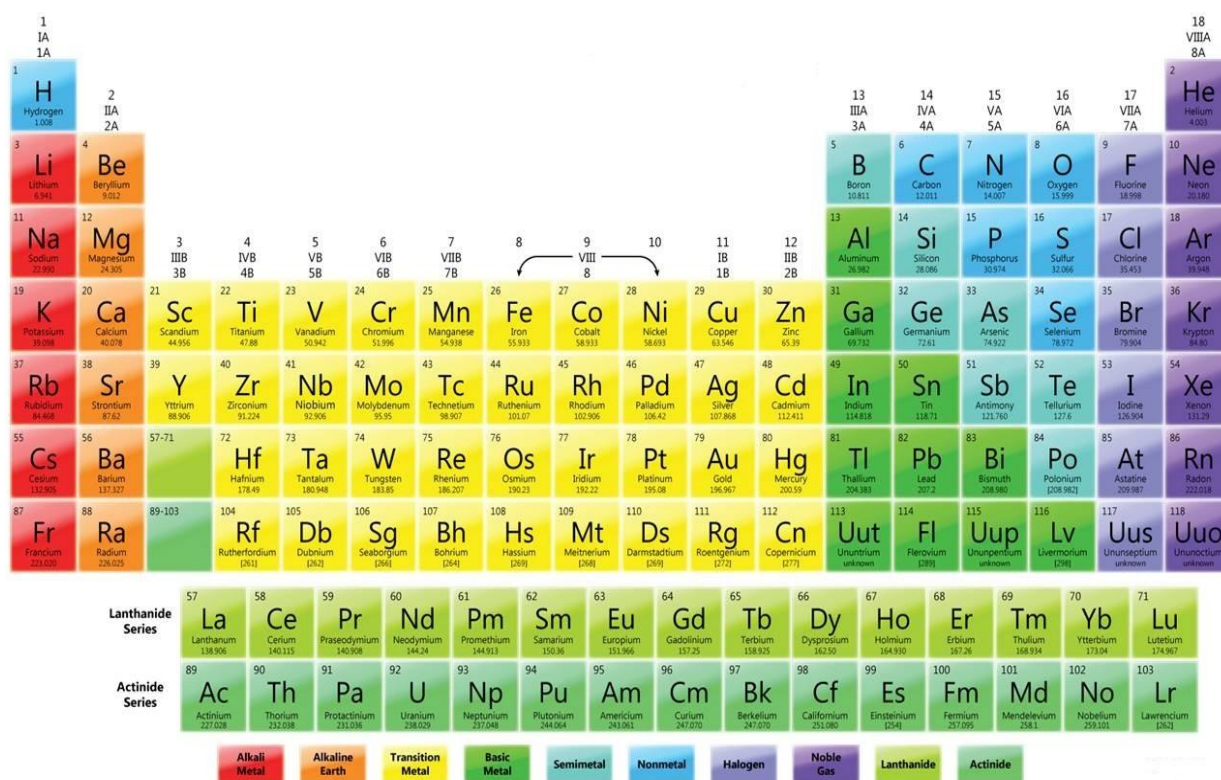


Figure 1: The Periodic Table with Group numbers and *s*, *p*, *d* and *f* blocks¹³

Group number

The most recent Periodic Table assigns **Group numbers** (1 to 18) to the vertical columns in the Periodic Table. Thus, the elements Li, Na, K, Rb, Cs and Fr belong to **Group 1** (formerly Group I or IA); Be, Mg, Ca, Sr, Ba and Ra belong to **Group 2** (formerly Group II or IIA); Sc, Y, La and Ac belong to **Group 3** (formerly Group IIIB); Fe, Ru and Os belong to **Group 8**; (formerly Group VIII); B, Al, Ga, In and Tl belong to **Group 13** (formerly Group III or IIIA); the halogens F, Cl, Br, I and At belongs to **Group 17** (formerly Group VII or VIIA) and the noble gases He, Ne, Ar, Kr, Xe and Rn belong to **Group 18** (formerly Group 0).

Blocks

The Periodic Table has **four blocks** (*s*, *p*, *d* and *f*) as *s*, *p*, *d* and *f* levels are being filled. The ***s*-block** consists of the elements in Groups 1 and 2. The elements in Groups 1 and 2 are also called *alkali metals* and *alkaline earth metals* respectively. The ***p*-block** consists of elements

in the Groups from 13 to 18. The **d-block** is located between s and p blocks and it consists of elements in the Groups from 3 to 12. The **f-block** has 28 elements and it is equally divided into **Lanthanides** ($Z = 58$ to 71) and **Actinides** ($Z = 90$ to 103). In most Periodic Tables, hydrogen is placed above lithium although it shows different properties to those of Li, Na, K and Cs. In this lesson, hydrogen will be considered separately.

Hydrogen

Hydrogen is the first element in the Periodic Table. Hydrogen atom has a very simple structure consisting of a nucleus with a single proton and one electron. Three isotopes of hydrogen are protium or proton ($H = {}^1_1H$), deuterium or deuteron ($D = {}^2_1H$) and tritium ($T = {}^3_1H$). At room temperature, hydrogen is a diatomic, colourless and odourless gas.

In some Periodic Tables, hydrogen is placed above the Group 1 elements but sometimes it is placed above the Group 17 elements. *It is best to have a separate place for hydrogen.*

Q : Give reasons for placing hydrogen with the Group 17 elements.

A : Hydrogen needs just one electron to attain the inert gas configuration ($1s^2$). Halogens also need only one electron to achieve the inert gas configuration (ns^2np^6). Hydrogen forms the hydride ion, H^- with electropositive metals. Halogens also form mono-negative anions, X^- . Hydrogen (H_2) is a diatomic gas similar to most halogens (X_2).



Activity

1. Give reasons for placing hydrogen with the Group 1 elements.

The most important use of hydrogen is for the synthesis of ammonia (NH_3) by the Haber process. It is also used to synthesize methyl alcohol from carbon monoxide (CO) in the Fischer-Tropsch synthesis. Hydrogen is also used for the conversion of vegetable fats and oils to give edible fats (*e.g.* margarine). It is sometimes useful as a reducing agent in the

extraction of metals, as a fuel (hydrogen powered cars), in oxy-hydrogen welding, and the production of organic compounds.

s-Block Elements

The elements in the Groups 1 and 2 are called *s-block* elements.

1	2
3 Li	4 Be
11 Na	12 Mg
19 K	20 Ca
37 Rb	38 Sr
55 Cs	56 Ba
87 Fr	88 Ra

Group 1 elements

The Group 1 elements (formerly Group I or IA) include **lithium** (Li), **sodium** (Na), **potassium** (K), **rubidium** (Rb), **caesium** (Cs) and **francium** (Fr). They are also called “**alkali metals**” and they conduct electricity and heat. Alkali metals are *reactive* and they are not found in nature as *free* metals. Sodium and potassium are relatively more abundant than the other metals in this group. The general valence electron configuration of these elements is ns^1 .



Lithium¹⁴



Sodium¹⁵



Potassium¹⁶



Rubidium¹⁷



Caesium¹⁸



Francium¹⁹

Group 2 elements

Group 2 (formerly Group II or IIA) elements include **beryllium** (Be), **magnesium** (Mg), **calcium** (Ca), **strontium** (Sr), **barium** (Ba) and **radium** (Ra). They are also called “*alkaline earth metals*”. The Group 2 elements are too *reactive* to occur as *free elements* in nature and all are *powerful reducing agents*. The general valence electron configuration of Group 2 elements is ns^2 .



Beryllium²⁰



Magnesium²¹



Calcium²²



Strontium²³



Barium²⁴



Radium²⁵

p-Block Elements

The elements in the Groups 13 to 18 belong to *p-block*.

					18
13	14	15	16	17	2 He
5 B	6 C	7 N	8 O	9 F	10 Ne
13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn

Group 13 elements

This family (formerly Group III or IIIA) consists of **boron** (B), **aluminium** (Al), **gallium** (Ga), **indium** (In) and **thallium** (Tl). Boron is regarded as a *non-metallic* element with some metallic characteristics (*i. e.* B is a semi-metal). The other elements are all metals. The general valence electron configuration of these elements can be written as ns^2np^1 . All the elements show the oxidation state +3.



boron²⁶



Aluminium²⁷



Gallium²⁸



Indium²⁹



Thallium³⁰

Group 14 elements

The Group 14 elements (formerly Group IV or IVA) include **carbon** (C), **silicon** (Si), and **germanium** (Ge), **tin** (Sn) and **lead** (Pb). Carbon is a *non-metal*; silicon and germanium are *metalloids*; tin and lead are *metals*. The general valence electron configuration of Group 14 elements is ns^2np^2 and they exhibit a wide range of oxidation states from -4 to +4. The oxidation numbers of carbon in CH₄, diamond, CO and CO₂ are -4, 0, +2 and +4,

respectively. Elemental carbon exists in three allotropic forms; allotropes are diamond, graphite and fullerene.

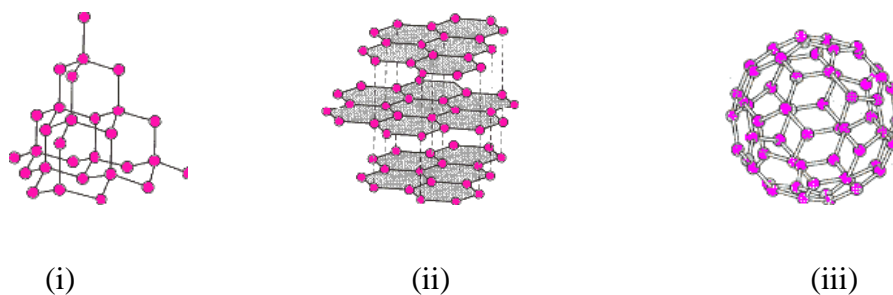


Figure 2: Allotropes of carbon; diamond, graphite and fullerene (C_{60})³¹

Coal is still used as a fuel source. Carbon is also used as an electrode material. Carbon has three isotopes: ^{12}C , ^{13}C and ^{14}C . Carbon-14 isotope is used in archaeological dating, for example, to determine the age of fossils. Graphite is a good lubricant and diamond is a precious material. Fullerenes have found various applications in the field of nanotechnology. Ultra-pure silicon is used in electronic industry. Polymeric silicones are used as oils and waxes.



Group 15 elements

The Group 15 elements include **nitrogen** (N), **phosphorus** (P), **arsenic** (As), **antimony** (Sb) and **bismuth** (Bi). N and P are *non-metals*, As and Sb are *metalloids* whereas Bi has characteristic properties of a main group metal. The general valence electron configuration of Group 15 elements is ns^2np^3 and they exhibit a range of oxidation states from -3 to $+5$. The oxidation numbers of nitrogen in NH_3 , N_2 , NF_3 and NO_3^- are -3 , 0 , $+3$ and $+5$, respectively. Air contains 78% of N_2 (by volume). Nitrogen is obtained by the *fractional distillation* of liquefied air. Conversion of inert nitrogen gas into useful nitrogen compounds is called “Nitrogen fixation”. NH_3 is *commercially* produced in the Haber synthesis by reacting N_2 and H_2 in the presence of an *iron catalyst* at 500°C and at a pressure of 300 atm. Nitric acid is manufactured by oxidising ammonia using air.



Nitrogen³⁷

Phosphorus³⁸

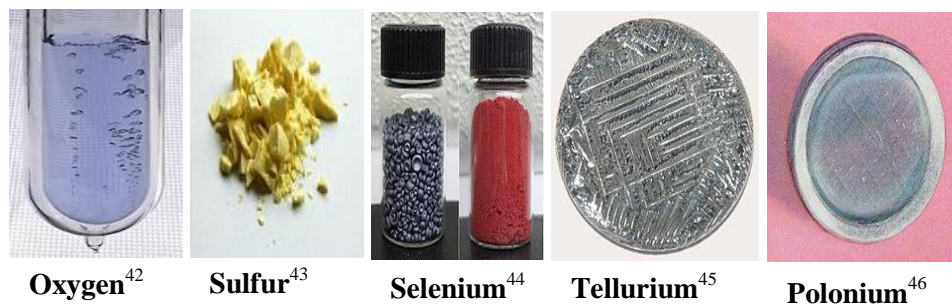
Arsenic³⁹

Antimony⁴⁰

Bismuth⁴¹

Group 16 elements

This family consists of the elements: **oxygen** (O), **sulfur** (S), **selenium** (Se), **tellurium** (Te) and **polonium** (Po). The Group 16 elements are also known as “**chalcogens**”. Polonium is the only true metal; Te is a semi-metal; oxygen, sulfur and selenium are *non*-metals. The general valence electron configuration of Group 16 elements is ns^2np^4 and they exhibit a wide range of oxidation states from -2 to $+6$. The oxidation numbers of sulphur in H_2S , S_8 , SO_2 and SO_3 are -2 , 0 , $+4$ and $+6$, respectively. In general, the reactivity of these elements decreases in the following order $O > S > Se > Te$. Oxygen (O_2) is the second most abundant gas in the earth’s atmosphere and it is obtained from liquid air by fractional distillation. Photolysis of the oxygen generates two oxygen radicals which combine with another oxygen molecule to form ozone (O_3).



Oxygen⁴²

Sulfur⁴³

Selenium⁴⁴

Tellurium⁴⁵

Polonium⁴⁶

Group 17 elements

This family consists of **fluorine** (F), **chlorine** (Cl), **bromine** (Br), **iodine** (I) and **astatine** (At). Halogens are non-metals and they exist as **diatomic** molecules. At room temperature, fluorine and chlorine are yellow green gasses; bromine is a red liquid and iodine a black solid. Solutions of iodine are purple in colour. Fluorine is the most electronegative element and a powerful oxidizing agent. Fluorine is very reactive and forms compounds with most elements.

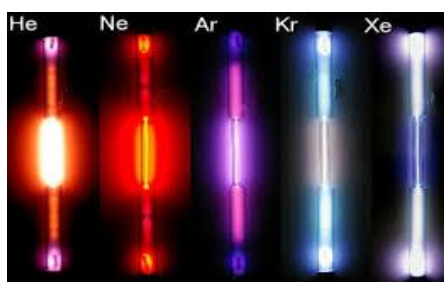
The general valence electron configuration of an element is ns^2np^5 . They achieve the next noble gas electron configuration by gaining one electron from a metal or sharing an electron with a non-metal. Thus, all elements show the oxidation number -1 . Positive oxidation

numbers such as +1, +3, +5 and +7 occur for chlorine, bromine and iodine, mainly in oxyanions and inter-halogen compounds. The bond dissociation energies, $D(X-X)$, of the halogen molecules decrease in the following order. $D(Cl-Cl) > D(Br-Br) > D(F-F) \approx D(I-I)$



Group 18 elements

The Group 18 elements are normally called the **noble gases** and they include **helium** (He), **neon** (Ne), **argon** (Ar), **krypton** (Kr), **xenon** (Xe) and **radon** (Rn). They are called noble gases because they are chemically not very reactive. They are also called **rare gases** as they are found only in very small quantities in the atmosphere and in the earth's crust. All are **monatomic**, colourless and odourless gases. Radon is radioactive. The general valence electron configuration of Ne, Ar, Kr, Xe and Rn is ns^2np^6 (the total number of electrons in the outermost shell is 8). The electron configuration of He is $1s^2$. In all noble gases, the valence shell has the maximum number of electrons, thus atoms do not combine with one another (to form diatomic gas molecules). They also do not combine readily with other elements. This is the reason why they are referred to as 'inert' gases. But since 1960s fluorides and oxides of Xe are known.



Group 18 elements⁵¹

d-Block Elements (Groups 3 to 12)

All *d*-elements are metals and they are generally called transition metals. These elements are obtained by the filling up of *d*-levels. Each metal consists of a positively charged lattice and a pool of free electrons. All *d*-elements are good heat and electrical conductors.

	3	4	5	6	7	8	9	10	11	12
3d	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn
4d	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd
5d	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg
6d	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn

Most of the *d*-elements show variable oxidation states. *d*-Block elements have higher melting and boiling points than *s*- and *p*-block elements. In the *3d*-series, the melting point gradually increases from Sc to V and then decreases from Fe to Zn. *d*-Block elements show higher densities compared to *s*- and *p*-block elements. In the *3d*-series, the density increases from Sc to Cu. The *d*-block elements are less electropositive than the Group 1 metals. In the *3d*-series, the electronegativity tends to increase slightly from Sc to Cu. Majority of the transition metal complexes are coloured. *d*-Block elements show variable oxidation states and their metal complexes play an important role as catalysts. Zn, Cd and Hg are not considered as transition metals. Since scandium does not form stable Sc^+ (d^2) or Sc^{2+} (d^1) ions, in many text books, it is not considered a transition metal.



Activity

2. Why is Zn not considered as a transition metal?

f-Block Elements

f-Block contains 28 elements and it is placed at the bottom of the Periodic Table. These elements are obtained by the filling up of *f*-levels, see section on electron configuration for more details.

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	71 Lu
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	103 Lr

f-Block has two horizontal rows: the elements in the first row are called **Lanthanides** (or Lanthanoids) whilst the elements in the second row are called **Actinides** (or actinoids). However, the lanthanide elements (*i.e.* elements after lanthanum (La) having atomic numbers 57 to 71) should be placed at the period 6 between the elements La and Hf. Similarly, the actinide elements (*i.e.* elements after actinium having atomic numbers 90 to 103) should be placed at the period 7. Electron configurations of elements can be used to group elements into respective blocks. Reactivity of elements, metallic and non-metallic properties of elements, and ionization energies of elements can be easily predicted by considering the electron configuration of an element.

Electron configuration

The electrons are found in different **energy levels** or **shells** (K, L, M, N, *etc.*). Each shell consists of a number of **sub-shells** (*s, p, d, f, etc.*). The number of **orbitals** found in *s, p, d* and *f* sub-shells are 1, 3, 5 and 7, respectively. The maximum number of electrons found in an orbital is **two**. Distribution of orbitals among the first four shells is given in Table 1.

Table 1: Distribution of orbitals among the first four shells

n	Shell	Orbitals	Total number of orbitals
1	K	1s	1
2	L	2s 2p _x , 2p _y , 2p _z	4
3	M	3s 3p _x , 3p _y , 3p _z 3d _{xy} , 3d _{yz} , 3d _{xz} , 3d _{x²-y²} , 3d _{z²}	9
4	N	4s 4p _x , 4p _y , 4p _z 4d _{xy} , 4d _{yz} , 4d _{xz} , 4d _{x²-y²} , 4d _{z²} Seven 4f orbitals	16

Q: What is the order of filling of **energy levels**?

A: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, *etc.*

In terms of energy: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s \dots\dots$

The elements of the Periodic Table can be further classified into four main categories depending on the *electronic configurations* of the atoms. They are:

1. Noble gases (Group 18)
2. Representative elements (*s*- and *p*-blocks)
3. Transition elements (*d*-block)
4. Inner-transition elements (*f*-block)

Noble gases

These elements are classified under Group 18 of the Periodic Table. Except helium (which has two electrons in the 1s orbital; thus the valence electron configuration is $1s^2$), all the other elements have the valence electron configuration ns^2np^6 . These elements have a **filled valence shell** thus they are inert. All are monatomic gases.

Q : What are the electron configurations of Ar and K?

A : The electron configuration of Ar is $1s^2 2s^2 2p^6 3s^2 3p^6$, which can be represented as [Ar].

The electron configuration of K is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$.

Thus, the electron configuration of K can be written as [Ar] $4s^1$.

Note that the electron configuration of K^+ is $1s^2 2s^2 2p^6 3s^2 3p^6 = [Ar]$; *i.e.* K^+ and Ar have the same number of electrons (**isoelectronic**). Similarly, Ca^{2+} , Ga^{3+} , P^{3-} , S^{2-} and Cl^- are isoelectronic with Ar.

Representative elements

These are the elements with the valence electron configurations from ns^1 , ns^2 , ns^2np^1 , ns^2np^2 , ns^2np^3 , ns^2np^4 to ns^2np^5 . These elements belong to s - and p -blocks.

Q : What is the electron configuration of P?

A : The electron configuration of P is $1s^22s^22p^63s^23p^3 = [\text{Ne}]3s^23p^3$.

Note that the electron configuration of P^{3-} is $1s^22s^22p^63s^23p^6 = [\text{Ar}]$.

Transition elements

The atoms or the stable ions of these elements have *incompletely filled d-shells* thus they are called “*transition metals*”. In general they can have either one or two electrons in the last s -shell and one to ten electrons in the penultimate (*i.e.* one before the last) d -shell. The valence electron configuration can be written as $ns^{1 \text{ or } 2} (n-1)d^{1 \text{ to } 10}$. The valence electron configuration of a d -block element can also be represented as $(n+1)s^2nd^m$ or $(n+1)s^1nd^m$ where $n = 3, 4, 5$ or 6 and $m = 1, 2, 3, \dots$ or 10 .

The first transition series ($3d$ -series) starts with the Group 3 element scandium Sc (with the atomic number, Z of 21) which has the electron configuration $1s^22s^22p^63s^23p^64s^23d^1$ and ends up with the Group 11 element Cu ($Z = 29$) with the electron configuration $1s^22s^22p^63s^23p^64s^13d^{10}$. Although copper in the element form has a filled d shell, it forms the Cu^{2+} ion with the electron configuration $1s^22s^22p^63s^23p^63d^9$, and it is a transition metal. Note that for free atom, the energies of the $3d$ and $4s$ orbitals are very close ($3d > 4s$). In metal cations $3d$ orbitals are much more stabilised than the $4s$ orbitals (*i.e.* $3d < 4s$), thus $4s$ electrons are lost first. Therefore the electron configuration of Ti^{2+} and V^{3+} ions can be written as $[\text{Ar}]3d^2$.

Q : What is the electron configuration of the Cu^+ ion?

A : The electron configuration of the Cu^+ ion is $1s^22s^22p^63s^23p^63d^{10} = [\text{Ar}]3d^{10}$

Inner-transition elements

These elements are obtained by the filling up of f -levels. For example, the fourteen elements from Ce ($Z = 58$) to Yb ($Z = 70$) are obtained by filling of the $4f$ -level. In this series, called **Lanthanides**, the filling of electrons is rather **irregular** in the sense that the $5d$ energy level is occasionally filled in preference to the $4f$ -level. Their electron configurations can be written as $[\text{Xe}]4f^m5d^16s^2$ ($m = 1-14$). The elements obtained by similar filling up of the $5f$ level are called **Actinides**.

Q : What is the electron configuration of Ce?

A : The electron configuration of the Ce is $[\text{Xe}]4f^15d^16s^2$.

Note that in the free atom, the order of filling of energy levels is $6s < 5d < 4f$.



Activity

3. (a) Write the electron configurations of the following atoms/ions.

(i) K (ii) Kr (iii) Cr (iv) Fe^{2+} (v) Cl^- (vi) Si

(b) Write the valence electron configurations of the following atoms/ions.

(i) Cs^+ (ii) Ir^{3+} (iii) Sr (iv) Mn^{7+} (v) S^{2-}

2. Periodicity

Introduction

We have studied the classification of the elements in the Periodic Table. Let us now consider some physical properties of the elements: atomic size, atomic radius, ionic radius, ionization energy, electron affinity, electronegativity, melting and boiling points *etc.*

2.1 Atomic and ionic sizes

A few elements exist as atoms at ambient temperatures. Most elements form neutral molecules (*e.g.* N₂, H₂, *etc.*) or ions. *Addition or removal of an electron(s) to or from an atom generates ions.* The size of an atom depends on the effective volume of the outer electrons. Atomic sizes of elements are important because properties such as boiling point, melting point, ionization energy and electronegativity depend on their sizes. *The size of an atom/ion can be estimated from the measured distance between the nuclei of two neighbouring atom/ion by making the assumption that this distance is equal to the sum of the radii of the two atom/ion.* Because this distance depends on the type of bond between the atoms, there are few types of radii, namely, atomic, covalent, ionic, metallic radii *etc.* Each radius is calculated for a particular type of bond.

2.2 Atomic radii

The radius of an atom can be considered as the distance of closest approach to another identical atom. This definition does not mention whether the atoms are chemically bonded or not, thus radius depends on the type of bond between the atoms/ions. Therefore, we use *covalent radius* for covalent compounds, *i.e.* the equilibrium distance between two covalently bonded atoms. In free metals we use *metallic radius*, *i.e.* the half distance between the neighbouring atoms in the metal.

Several experimental methods are available for measuring inter nuclear distances and hence for calculating the radii of atoms. For example, in the Cl₂ molecule, which is formed by sharing a pair of electrons between two chlorine atoms, the Cl–Cl distance is found to be

0.198 nm. Half of this distance (0.099 nm) will be the *covalent radius* of chlorine. In diamond, the C–C bond distance is found to be 0.154 nm and the same value is found for the C–C bond distance in many other aliphatic compounds. Hence the *covalent radius* of carbon is 0.077 nm. In free metals, each atom has many neighbours and the bonds between the atoms are known as “metallic bond”. *Metallic radii are usually larger than covalent bond radii.*

The atomic radius increases as the number of energy levels occupied by electrons increases. Hence the elements within a group in the Periodic Table would show *a gradual increase in atomic radius as the atomic number increases.* The atomic radii of the alkali metals (Li to Cs) are given below for comparison.

Element	Li	Na	K	Rb	Cs
Atomic radius (nm)	0.123	0.157	0.203	0.216	0.235

Activity



4. Compare the atomic radii of the alkaline earth metals.

Within a period there is *a general decrease in the radii of the atoms from left to right* of the Periodic Table. This decrease is mainly due to the increase of *positive charge* in the nucleus (*i.e.* the number of protons). The decrease in atomic radius as the atomic number increases from Li to F is given below for comparison.

Element	Li	Be	B	C	N	O	F
Atomic radius/nm	0.123	0.106	0.088	0.077	0.074	0.073	0.064

Across a period, from one element to the next, each additional electron added occupies the same principal energy level. Thus, the *effective positive charge* increases from left to right across a period but the increase in positive charge is not neutralized by increasing *shielding by shielding electrons* (see section 2.5). Hence, the *outermost electrons of an atom are drawn closer to the nucleus* resulting in a smaller atomic radius. Note that atomic radii of noble gases may be greater than those of the other elements in the same period.

The variation in covalent radii of the elements from hydrogen to potassium is given in Fig. 3.

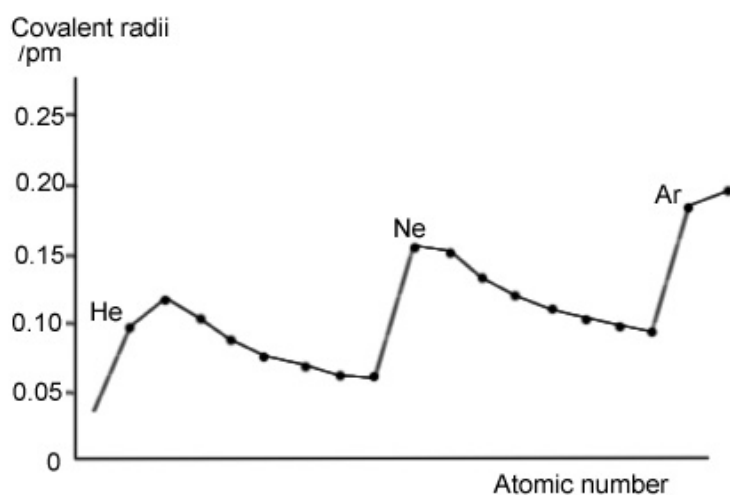


Figure 3: The variation in covalent radii of the elements from H to K

Activity

5. Compare the atomic radii of the elements in the 3rd period (Na to Cl).



2.3 Ionic radii

When atoms lose or gain electrons, ions are formed. The radii of such ions (*i.e.* ionic radii) can be determined from the distances between the ions in a crystal. For example, the ionic radii of the sodium ion (Na^+) and chloride ion (Cl^-) can be obtained from the inter-nuclear distances in sodium chloride crystals. X-ray diffractometer is used to determine the molecular crystal structure of a compound and atomic/ionic distances can be easily determined.

Positive ions (cations) are formed by the loss of electron(s) and are always smaller than the parent atom. For example, the atomic radius of vanadium is 0.131 nm whereas the ionic radii of V^{2+} , V^{3+} and V^{4+} are 0.088, 0.074 and 0.060 nm, respectively.

Q : Explain the above observation.

A : The size of an atom or ion decreases as the net positive charge on the particle increases.

The net positive charge present in a positive ion attracts the electrons more strongly towards the nucleus; this would lead to a contraction of the electron cloud. Thus, ionic radii decrease as the positive charge increases.



Activity

6. The atomic radius of Li is 0.123 nm whereas the radius of the Li^+ ion is 0.068 nm. Explain this observation.

Addition of an electron to an atom generates a negatively charged ion (or anion). The incoming electron enters the highest energy level without an increase in nuclear charge (*i.e.* there is no change in the number of proton in the nucleus). There is electron-electron repulsion with an expansion of the electron cloud. Therefore, the size of the negatively charged ion increases over that of the atom. For example, the atomic radius of chlorine atom is 0.099 nm whereas the radius of the Cl^- ion is 0.181 nm.



Activity

7. The atomic radius of iodine is 0.128 nm whereas the radius of the I^- ion is 0.219 nm. Explain.

2.4 Metallic and non-metallic character

The tendency of an element to lose electrons to generate a positively charged ion (cation) is called its metallic character (which is related to electro-positivity). The tendency of an element to gain electrons to form negatively charged ions (anions) is called its non-metallic character. The metallic character of the elements increases as we go down a group; for example metallic character increases from Li to Cs as the size of the atoms increases when we go down a group. As a result, the electrons of the outermost shell can be easily removed. We can also say that the non-metallic character decreases from Li to Cs.



Activity

8. (a) Arrange the following elements in the order of increasing metallic character.

As, N, P and Sb

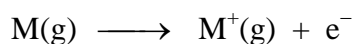
(b) Arrange the following elements in the order of increasing non-metallic character.

Be, Ca, Mg and Sr

Across a period, from left to right, the metallic character of the elements decreases and the non-metallic character increases. The reasons are (i) the atomic radius decreases from left to right and (ii) the effective nuclear charge increases from left to right. For example, among the elements in the second period, Li is the most metallic element and F is the most non-metallic element. The noble gases do not exhibit any metallic or non-metallic character as they have no tendency to lose or gain electrons.

2.5 Ionization Energy

The **first ionization energy** of an atom is the energy required to remove an electron from an atom in the gaseous state to an infinitely large distance. The process can be written as



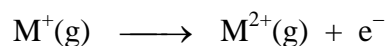
Since ionization is a process common to many chemical reactions, it is an important factor in interpreting reactivity of many elements.

The factors that affect the ionization energy of an atom are given below.

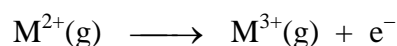
- (a) **Atomic radius:** As the atomic radius increases, the ionization energy becomes smaller.
- (b) **The effective nuclear charge:** When the effective charge of the nucleus increases, the ionization energy becomes larger.
- (c) **Shielding by inner electrons:** When an atom has more inner shells between the nucleus and the outer most electrons, the outer most electrons can be easily removed. The reasons are (i) inner-shell electrons (**shielding electrons**) repel outer most electrons and (ii) inner-shell electrons shield the outer most electrons from the nucleus; thus the effect of effective nuclear charge decreases as shielding increases (**shielding effect**). Therefore, the ionization energy becomes smaller with the increase of shielding. Shielding effect decreases in the following order $s > p > d > f$.

The *first ionization energy* refers to the removal of the most loosely bound electron from an isolated atom in the gaseous state. The second, third, fourth ...ionization energies refer to the removal of the second, third, fourthelectrons and so on. Since the energy required to remove an electron depends on how strongly it is bound to the nucleus of an atom, the ionization energy will be lowest when the effective nuclear charge is smallest and when the distance separating the electron from the nucleus is largest. The second ionization energy

refers to the energy required to remove one electron from the gaseous mono-cation, $M^+(g)$, as shown below.



The third ionization energy refers to the energy required to remove one electron from the gaseous dication, $M^{2+}(g)$, as shown below.



Ionization energy generally increases from left to right as (i) the atomic radius decreases from left to right and (ii) the effective nuclear charge increases from left to right of the periodic table. Shielding effect is very small as all outermost electrons are in the same shell.

However, there are several irregularities; *i.e.* some elements have ionization energy higher than that of the next element in a period. This slight deviation can be explained by considering the **electron configurations of the atoms and the ions formed**. It is known that completely filled and half-filled shells and sub shells show extra stability. Thus elements with s^2, s^2p^6 (completely filled) and s^2p^3 (half filled) configurations show higher ionization energies. It is also found that when the resulting cation has an electron configuration with a completely filled or half-filled shell then the ionization energy is found to be low. Figure 4 shows the first ionization energies of the first fifty four elements in the Periodic Table.

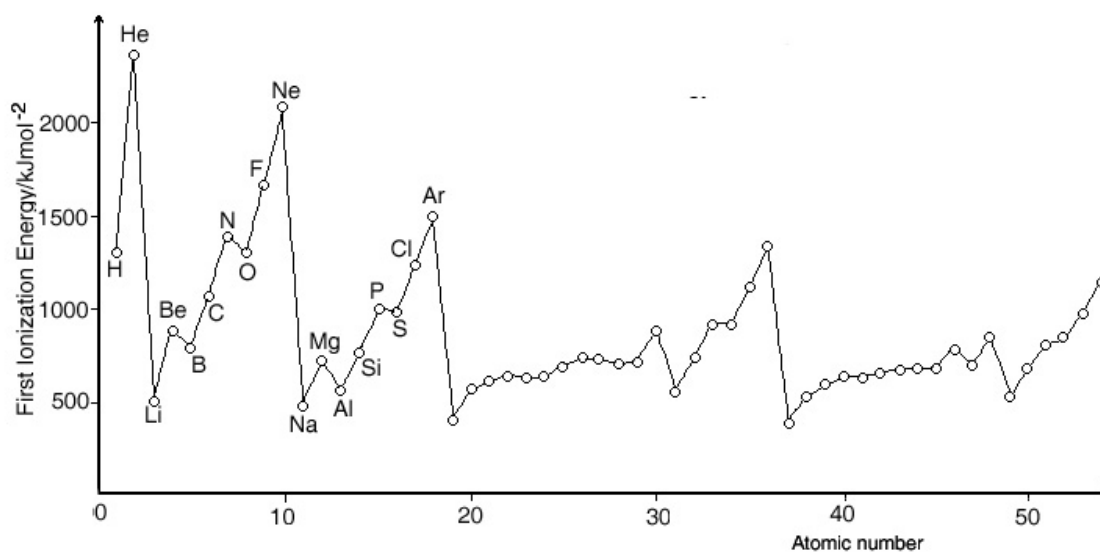


Figure 4: The first ionization energies of the first fifty four elements

It is clear from Figure 4 that the first ionization energy increases across a period with some irregularities. For example, *the first ionization energy of beryllium is greater than that of boron.*

Q : How would you explain the above observation?

A : The valence electron configuration of boron is $2s^2 2p^1$. The valence electron configuration of beryllium is $2s^2$, which has a *filled s sub shell*. Thus removal of an electron from beryllium is much more difficult than boron. Removal of one electron from boron gives a *filled s sub shell ($2s^2$) with extra stability*; therefore, ionization energy of boron is smaller than that of beryllium

The first ionization energy of noble gases is much higher when compared with those of other elements in the same period. Ionization energy is expressed in kJ mol^{-1} or eV where $1 \text{ eV} = 96.85 \text{ kJ mol}^{-1}$.

Ionization energy decreases as you go down a group. In this case, all elements have the same number of valence electrons. As we go down a group the valence electrons are held less tightly as they are shielded (and also repelled) by the electrons in the inner shells. Thus, removal of an electron from outermost orbital is much easier. As a result, the first ionization energy decreases as you go down a group. The first ionization energies of the alkali metals (Li to Cs) are given below for comparison.

Element	Li	Na	K	Rb	Cs
1 st Ionization Energy/ kJ mol^{-1}	520	513	419	400	380

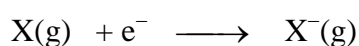


Activity

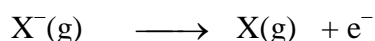
9. The first ionization energy of magnesium is greater than that of aluminium. Explain

2.6 Electron affinity

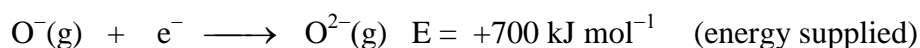
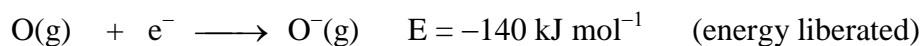
Electron affinity is defined as the amount of energy released when an electron is added to a *neutral gaseous atom*. It is the *energy released* during this *reduction process*,



This can be considered as the *reverse* of the ionization of $X^-(g)$ ion.



Electron affinity increases (increase in the negative value or the energy liberated) from left to right of the Periodic Table as (i) the effective nuclear charge increases and (ii) the size of the atom decreases, from left to right. For most elements, energy is liberated when electron is added to a neutral gaseous atom, and the electron affinity therefore takes a negative value. However, there are some elements with *positive electron affinity* values *i.e.* energy should be supplied from outside to force an electron to combine with the gaseous atom, to form an ion. Affinity for the second electron needs a large energy because the second electron must be forced on against the *net negative charge of the ion*. This is illustrated by the following example:



Activity

10. Electron affinity of chlorine is negative but that of argon is positive, explain.

2.7 Electronegativity

The electronegativity of an atom is a *measure of the tendency of the atom to attract electrons* when it is covalently bonded to other atom(s). Generally small atoms attract electrons more strongly than larger ones and are therefore more negative. Like ionization energy and electron affinity, electronegativity describes the tendency of an atom to attract electrons but here we consider a particular bonding situation. Unlike ionization energy and electron affinity, electronegativity cannot be measured directly.

There are several ways of expressing the electronegativity of elements. For example, in the Pauling's method the most electronegative element fluorine (F) is given an arbitrary value of 4.0. The electronegativity of other atoms is defined with reference to the value of fluorine. Electronegativity values for some selected elements are given below.

Element	F	O	Cl	N	C	H	Li
Electronegativity	4.0	3.5	3.0	3.0	2.5	2.1	1.0

Electronegativity is a useful quantity for predicting bond type, dipole moments and bond energies. For example, covalent bonds (*i.e.* formed by sharing of electrons between atoms) are formed by elements which have similar electronegativity values and ionic bonds (*i.e.* transfer of electrons from one to the other) are formed between elements with a large electronegativity difference.

The electronegativity of elements increase sharply across a period of the s and p block elements (for example from Li to F). This can be explained on the basis of the increased effective nuclear charge across a period. (Refer section 2.6 to see how the electron affinity varies across a period). The electronegativity generally decreases when going down a group. Figure 5 shows the variation of electronegativity across the third period.

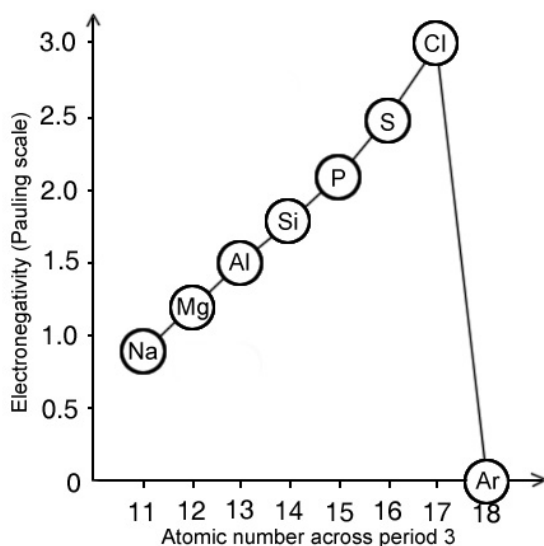


Figure 5: The variation of electronegativity across the third period



Activity

11. Arrange the following elements in the order of increasing electronegativity.

- (a) H, Li, C, N, O, F (b) Be, Mg, Ca, Ba (c) F, Cl, Br, I

2.8 Melting and boiling points

The melting point is defined as *the temperature at which pure solid is in equilibrium with pure liquid*. This temperature depends on the strength of the forces holding the

particles together in the solid and the extent to which these must be broken to form a liquid. The melting points you normally find in text books are usually measured at atmospheric pressure.

Within a group, there is a gradual change of melting points. As given in Table 2, the melting points of the alkali metals decrease when you go down the group and in the case of the halogens the melting points *increase* when you go down the group. In the case of alkali metals, the atoms are held together by weak metallic bonding. The smaller Li atoms pack more compactly in the solid and the metallic bonds in Li are stronger than that in Na, which is a larger atom. This is the reason for the gradual decrease in the melting points of alkali metals as you go down the group. In contrast halogens are held together by weak van der Waals forces. Thus halogens have relatively low melting points.

However, since structural changes are abrupt in going along a period, the melting point could change drastically when going from one element to the other across a period. As an example let us consider the adjacent element carbon and nitrogen; carbon is a solid at room temperature with a very high melting point whereas nitrogen is a gas with a very low melting point.

Table 2: Melting and boiling points of elements in Groups 1 and 17

Element	m.pt./°C	b.pt./°C
Li	181	1331
Na	98	890
K	64	766
Rb	39	701
Cs	29	685
F	-220	-188
Cl	-101	-34
Br	-7	58
I	114	183

Activity

12. Discuss the variation of melting points in Group 2 and Group 16 elements.



The variation in melting points of the elements from hydrogen to rubidium is given in Fig. 6.

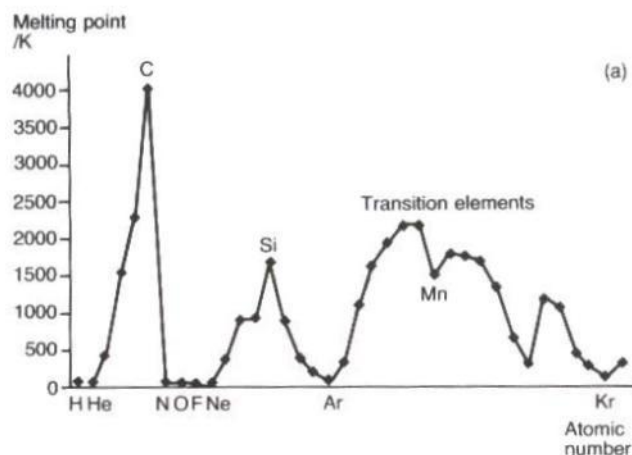


Figure 6: The variation in melting points of the elements from H to Rb

Boiling Points

The periodic trends of boiling points are similar to those of the melting point (see Table 2). In the case of metals the process of boiling requires complete breaking of bonds since there is little or no bonding of atoms in the vapour phase. Appreciable metallic bonding exists in the liquid phase and these bonds must be broken during vaporization. Thus a large amount of thermal energy must be supplied to vaporize hence the boiling points are high. There is also a large difference between melting and boiling points. Non-metals such as N_2 and Cl_2 require a small energy to break up the weak van der Waals forces thus the boiling points are also relatively low. In contrast, non-metals such as C and Si which have giant covalent structures show high boiling points.

2.9 Oxidation states

There is a simple periodic relationship between the common oxidation states (numbers) of elements and their *electron configuration*. The oxidation number of any *free element* is *zero*. Fluorine shows only one oxidation state, -1 .

Let us consider the *s*- and *p*-block elements. The maximum *positive oxidation number* shown by any representative element is equal to the total number of *s*- and *p*-electrons in the valence shell. For example, the Group 13 elements with the s^2p^1 configuration would have the maximum oxidation number of $+3$. The elements in the Groups 1 and 2 show only one *positive oxidation state*, e.g. Na^+ ($+1$) and Ca^{2+} ($+2$).

Q : What are the maximum oxidation numbers of Chromium and Chlorine?

A : +6 and +7 respectively.

The common oxidation states for each Group of elements from Li to Ne generally correspond to the number of electrons gained or lost in order to achieve the noble gas configuration. An element with the ns^2np^4 valence electron configuration would gain two electrons from another element and achieve the octet configuration; thus, oxygen would show an oxidation state of -2 . Thus, the oxide of aluminium is Al_2O_3 and the formula of barium chloride is $BaCl_2$.

Transition elements generally have varying numbers of d -electrons and they show several oxidation states.

Q : What are the formulae of oxides of Mn?

A : MnO , Mn_2O_3 , MnO_2 , MnO_3 , Mn_2O_7 , and MnO_4^- .

The oxidation numbers of Mn in MnO , Mn_2O_3 , MnO_2 , MnO_3 , Mn_2O_7 , and MnO_4^- are +2, +3, +4, +6, +7 and +7, respectively.



Activity

13. Determine the common oxidation number(s) of Groups 1, 2 and 16 using their electron configuration of representative element.

2.10 Diagonal relationships

There are certain irregularities related to the positions of many lighter elements. For example, the Group 2 element **beryllium** is slightly different from the rest of the members of the group; it resembles the Group 13 element **aluminium** in many ways. Another example is the similarity in the properties of **lithium** and **magnesium**. The oxidation states of these elements correspond to their group numbers, but the nature of

their compounds, their acidic and basic characteristics are related in a diagonal fashion. This is called *diagonal relationship* (relationship along the diagonal). This relationship can be explained in terms of **ionic potential** (equals to ratio of ionic charge/ionic radius). This term is really a measure of the charge density on the ion.

$$\text{Ionic potential for Be}^{2+} = \text{ionic charge (q)} / \text{ionic radius (r)} = 2/0.038 = 52.6$$

$$\text{Ionic potential for Al}^{3+} = 3/0.052 = 57.6$$

Since the ionic potentials for the two ions given above are similar, the properties such as *hydration energy*, *acid-base character* and *polarizing power* of the elements are expected to be similar.

Summary

- The Periodic Table is an arrangement of all the chemical elements in the order of increasing atomic number with elements having similar properties in the same vertical column.
- The Periodic Table has eighteen groups and seven periods. The 1st, 2nd, 3rd, 4th, 5th and 6th periods have 2, 8, 8, 18, 18 and 32 elements, respectively.
- The Periodic Table has four blocks (*s*, *p*, *d* and *f*) due to the filling of *s*, *p*, *d* and *f* levels.
- Atomic sizes of elements are important because properties such as boiling point, melting point, ionization energy and electronegativity depend on their sizes.
- The atomic radius increases as the number of energy levels occupied by electrons increases.
- Within a period, from left to right, there is a general decrease in the radii of the elements.
- The factors that affect the ionization energy of an atom are atomic radius, effective nuclear charge, and shielding and repulsion by inner electrons.
- Generally, ionization energy increases from left to right of the Periodic Table and ionization energy decreases as you go down a group.
- Electron affinity is defined as the amount of energy released when an electron is added to a neutral gaseous atom. Electron affinity increases (increase in the negative value or the energy liberated) from left to right of the Periodic Table.

- The electronegativity of an atom is a measure of the tendency of the atom to attract electrons when it is covalently bonded to other atom(s). The electronegativity generally decreases when going down a group.

Learning Outcomes

At the end of the session you should be able to

- explain briefly the history of the development of the Periodic Table
- discuss the classification of elements into *s*, *p*, *d* and *f*-blocks
- write the electron configuration of the elements in the *s*, *p*, *d* and *f*-blocks
- discuss the periodic trends in atomic size, atomic radius, and ionic radius of elements.
- discuss the metallic and non-metallic character of elements
- define ionization energy of elements
- explain trends within groups and periods
- discuss the electron affinity, electronegativity, melting and boiling points
- comment on elements which show diagonal relationships

Activity

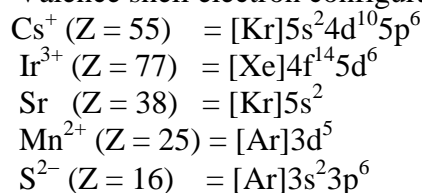
14. Define the term “electronegativity”.
15. Account for the difference in the ionization energies between the following successive pairs of elements in the Periodic Table. The values of ionization energy (in kJ mol^{-1}) are given in parentheses.
 - a) He (2372) and Li (520)
 - b) Li (520) and Be (900)
 - c) Be (900) and B (800)
 - d) N (1400) and O (1300)
16. There is a marked difference between the values for the fifth and the sixth of ionization energies of nitrogen. Explain.
17. Define the first electron affinity of an atom.
18. Discuss the variation of electron affinity in the Periodic Table.
19. What are the formulae of common oxides of Cr?
20. What are the formulae of common chlorides of vanadium?

Answer Guide to Activities

- The reasons for placing hydrogen with the Group 1 elements are (i) hydrogen has one electron in the outermost shell, (ii) hydrogen forms a mono-positive ion, H^+ , (iii) like Group 1 metals, it gets discharged at the cathode.
- Zn is not considered as a transition metal because (i) its d -shell is completely filled (d^{10}), (ii) it does not form coloured compounds, and (iii) it does not show variable oxidation states (Zn^{2+}).

- 3 (a) $K (Z = 19) = 1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{4s^1}$
 $Kr (Z = 36) = 1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{4s^2 3d^{10} 4p^6}$
 $Cr (Z = 24) = 1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{4s^1 3d^5}$
 $Fe^{2+} (Z = 26) = 1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{3d^6}$
 $Cl^- (Z = 17) = 1s^2 2s^2 2p^6 \mathbf{3s^2 3p^6}$
 $Si (Z = 14) = 1s^2 2s^2 2p^6 \mathbf{3s^2 3p^2}$ Note: Valence shell electron configuration is in bold.

(b) Valence shell electron configurations are given below.



4

Element	Be	Mg	Ca	Sr	Ba
Atomic radii (nm)	0.106	0.14	0.17	0.19	0.198

The atomic number and the number of shells increase from Be to Ba. Therefore the size of the atoms increases from Be to Ba.

- The atomic radii generally decrease as we move from Na to Cl. As atomic number increases, the effective nuclear charge (positive) increases. Therefore, the attraction of electrons to the nucleus is stronger when moving from left to right.
- $Li = 1s^2 2s^1$, $Li^+ = 1s^2$
Once lithium forms a cation by giving out the outermost electron, the number of shells is reduced to 1. The effective nuclear charge has increased. As a result, the two $1s$ electrons are more strongly attracted to the nucleus causing a significant reduction in its radius.
- After receiving an electron, iodine atom generates the negatively charged iodide ion (I^-). This incoming electron enters the outermost sub-shell ($5p$) which is now completely filled. The electron cloud is expanded due to electron-electron repulsion. Therefore, the radius of the I^- ion is much larger than that of the iodine atom.
- Metallic Character: $N < P < As < Sb$;
 - Non-metallic Character: $Sr < Ca < Mg < Be$

9 $\text{Mg} = 1s^2 2s^2 2p^6 3s^2$ $\text{Al} = 1s^2 2s^2 2p^6 3s^2 3p^1$
 The first ionization energy means the energy required to remove an electron from the gaseous atom. Mg has two outermost electrons in the completely filled *s* sub-shell. Since the electron configurations with filled shells/sub-shells show an extra stability than those with unfilled shells/sub-shells, the removal of an electron from Mg is not easy. In the case of Al, the electron is removed from the unfilled *p* sub-shell and the removal of an electron leads to a filled electron configuration, therefore it is much easy to remove an electron from Al. Therefore, the first ionization energy of Mg is greater than that of Al.

10 $\text{Cl} = 1s^2 2s^2 2p^6 3s^2 3p^5$ $\text{Ar} = 1s^2 2s^2 2p^6 3s^2 3p^6$
 By getting one electron, Cl atom can easily attain the stable noble gas configuration, the electron affinity of Cl atom is a negative value (energy is liberated when an electron is added). But it is difficult to give an electron to Ar, since it destroys the stable noble gas configuration. Therefore energy should be supplied from outside to force an electron to combine with Ar(g). Hence electron affinity of Ar is positive.

11 Electronegativity: a) $\text{Li} < \text{H} < \text{C} < \text{N} < \text{O} < \text{F}$
 b) $\text{Ba} < \text{Ca} < \text{Mg} < \text{Be}$
 c) $\text{I} < \text{Br} < \text{Cl} < \text{F}$

12

Group 2	m. pt./°C	Group 16	m. pt./°C
Be	1283	O	-219
Mg	650	S	114
Ca	850	Se	217
Sr	770	Te	450
Ba	710		

In general, the melting points of the Group 2 elements decrease with the increase of atomic number. The atoms are held together by weak metallic bonds and the strength of bonding decreases when moving down the group. Be atoms are packed more compactly (since its metallic bonds are stronger than others) and the melting point of Be is significantly high. The melting point of Mg is unexpectedly lower than Ca.

In Group 16, the oxygen has the lowest melting point. In general, melting points increase as we go down the group because the covalent/ionic bond strengths increase with increasing atomic number. Since the strength of the forces which hold the atoms in the solid state is high, it is difficult to form liquid by breaking these bonds.

13 (a) **Group 1 elements**

For example, the electron configuration of $\text{Na} = 1s^2 2s^2 2p^6 3s^1$

Na attains its stable electron configuration by giving away the outermost electron, hence the oxidation number of the Group 1 elements is always +1

(b) **Group 2 elements**

For example, the electron configuration of $\text{Mg} = 1s^2 2s^2 2p^6 3s^2$

By giving away 2 electrons it can attain the stable noble gas configuration. Therefore, the oxidation number of the Group 2 elements is always +2.

(c) Group 16 elements

For example, the electron configuration of O = $1s^2 2s^2 2p^4$

Oxidation number of O in its compound is often -2 , because it tries to attain noble gas configuration by gaining 2 electrons.

S = $1s^2 2s^2 2p^6 3s^2 3p^4$

Oxidation number of S varies from -2 to $+6$; most common oxidation numbers are -2 , $+4$ and $+6$. This is due to the availability of *d* orbitals.

- 14 The electronegativity of an atom is a measure of the tendency of the atom to attract electrons when it is covalently bonded to other atom(s).
- 15 a) He (2372) and Li (520)
The first ionization energy of He is larger than that of Li. The reason is, He has two outermost electrons in a stable filled *s* sub shell. But Li has the outermost electron in a half filled *s* orbital. The removal of one electron from the filled *s* sub shell is much harder than removal of a single electron from the half-filled sub shell.
- b) Li (520) and Be (900)
Li has one outermost electron in the *s* sub shell while Be has two outermost electrons in *s* sub shell. Since the removal of one electron from the filled *s* sub shell is much harder than the removal of a single electron from a half filled sub shell, Be has larger first ionization energy than that of Li.
- c) Be (900) and B (800); Be = $1s^2 2s^2$, B = $1s^2 2s^2 2p^1$
The first ionization energy of Be is larger than that of B. The reason is that Be has two outermost electrons in stable filled *s* sub shell. But the outermost electron of B is in the *p* sub shell which is not filled. Therefore the removal of one electron from stable *s* sub shell is difficult and hence the first ionization energy of Be is larger than that of B.
- 16 N ($Z = 7$) = $1s^2 2s^2 2p^3$
The 5th ionization energy of N is the energy required to remove an electron from the tetravalent gaseous cation. These electrons are removed from the *s* sub shell which has one electron (half-filled sub-shell). The 6th ionization energy of N means the energy required for removing an electron from the gaseous pentavalent cation and these electrons have to be removed from the most stable filled *s* sub shell. Removal of an electron from the filled sub shell is much harder because this 1*s* sub shell is situated very much closer to the nucleus. Therefore, the 6th ionization energy of N is significantly high.
- 17 The amount of energy released when an electron is added to a neutral gaseous atom.
- 18 The electron affinity increases from left to right of the Periodic Table as the effective nuclear charge increases and the size of the atom decreases, from left to right.
- 19 Cr₂O₃, CrO₂ and CrO₃
- 20 VCl₂, VCl₃ and VCl₄.

Study Questions

- Which element has a single electron in its outermost sub-shell?
1) Ca 2) C 3) P 4) Ge 5) Cu
- The element with the electron configuration of $s^2p^6d^{10}s^1$ is
1) Mo 2) K 3) Fe 4) Ag 5) Zn
- Number of d electrons in the Mn^{2+} ion is
1) 3 2) 4 3) 5 4) 6 5) 7
- The ion/element which is isoelectronic with the Fe^{3+} ion is
1) Cr^{3+} 2) Cr 3) Mn^{2+} 4) Mn 5) Co^{2+}
- The highest oxidation state of Mn is
1) +2 2) +4 3) +5 4) +6 5) +7
- The highest oxidation state of fluorine is
1) -1 2) 0 3) +1 4) +5 5) +7
- The highest oxidation state of chlorine is
1) -1 2) 0 3) +1 4) +5 5) +7
- The molecular formula of the oxide formed by the highest oxidation state of the element X with the atomic number 17 is;
1) X_2O 2) XO_2 3) XO_3 4) X_2O_5 5) X_2O_7
- Which one of the following is the smallest in size?
1) N^{3-} 2) O^{2-} 3) F^- 4) Na 5) K
- Which one of the following electron configurations belongs to an alkaline earth element?
1) $1s^2$ 2) ns^2np^1 3) ns^2np^6 4) $nd^{10}(n+1)s^2$ 5) None of the above
- The element/ion isoelectronic with Cr^{3+} is
1) V^{2+} 2) Ti 3) Mn^{2+} 4) Fe^{3+} 5) B^{2+}
- Which one of the following atoms has the d^7s^2 electron configuration?
1) Fe 2) Co 3) Ni 4) Cr 5) Zn
- The ion which does not form a coloured compound is
1) Zn^{2+} 2) Cr^{2+} 3) Mn^{2+} 4) Fe^{2+} 5) Ni^{2+}
- Which one of the following elements is a metalloid?
1) Sb 2) Sr 3) Se 4) Ga 5) Bi
- Which one of the following elements is a non-metal?
1) Se 2) As 3) Pt 4) Ga 5) B

16. Which element has the lowest first ionization energy?
 1) N 2) O 3) F 4) Ni 5) He
17. The 1st, 2nd, 3rd...ionization energies of an element are 685, 1320, 2240, 8950, 11600, 14500 kJ mol⁻¹ respectively. The Group number of this element is
 1) 3 2). 13 3) 14 4) 15 5) 16
18. The molecule which exhibits hydrogen bonding is
 1) CH₃I 2) N₂ 3) H₂ 4). HF 5) CH₄
19. The removal of one electron is easiest in
 1). Ni 2) Na⁺ 3) Mg²⁺ 4) Al³⁺ 5) O²⁻
20. Which one of the following ions has the smallest ionic radius?
 1) Al³⁺ 2) Ga³⁺ 3) Se²⁻ 4) N³⁻ 5). O²⁻
21. The lowest oxidation state of the element X is -3. The valence electron configuration of X is
 1) s² p¹ 2). s² p³ 3) d² s¹ 4) d¹ s² 5) s² p⁵
22. The highest oxidation state of the element M is +3. The valence electron configuration of M is,
 1) s² p³ 2) d³ s² 3) d² s⁰ 4) s¹ p² 5). d¹ s²
23. Select the element with the highest second ionization energy.
 1). Na 2) Mg 3) Ca 4) Be 5) Ba
24. Which element has a single electron in its d sub-shell?
 1). Sc 2) Ti 3) V 4) Cr 5) Mn
25. Which one is not involved in the development of the Periodic Table?
 1) Newlands 2) Mendeleev 3) Rutherford 4). Gay-Lussac 5) Mayer
26. What is the Group number of the element with the valence electron configuration ns² np⁵?
 1) Group 1 2) Group 5 3) Group 7 4) Group 13 5). Group 17
27. The Group number of the element with the atomic number 33 is
 1) 5 2) 14 3). 15 4) 16 5) 17
28. The concept of triads was introduced by
 1). Dobereiner 2) Newlands 3) Rutherford 4) Mayer 5) Bohr
29. Which element (M) forms the oxide with the molecular formula MO₃?
 1) Be 2) Al 3). Se 4) Si 5) C
30. "Octet Rule" was introduced by
 1). Newlands 2) Mayer 3) Bohr 4) Dalton 5) Rutherford
31. Which one of the following set is a Dobereiner's triad?
 1) Na, K, Cs 2). Cl, Br, I 3) Mg, Ca, Sr 4) S, Se, Te 5) Be, B, C

32. The outermost electron configuration of halogens is
 1) $(n-1)d^{10}ns^2np^5$ 2) ns^2np^5 3) $(n-1)d^{10}ns^2np^4$
 4) ns^2np^4 5) None of the above
33. The atomic number of the element which has a single electron in its outermost shell is
 1) 5 2) 14 3) 20 4) 25 5) 29
34. The correct order of increasing the ionic radii is
 1) $Na^+ < Mg^{2+} < Ca^{2+}$ 2) $Mg^{2+} < Na^+ < Ca^{2+}$
 3) $Ca^{2+} < Mg^{2+} < Na^+$ 4) $Mg^{2+} < Ca^{2+} < Na^+$
 5) $Ca^{2+} < Na^+ < Mg^{2+}$
35. The element which is more likely to lose two electrons easily is
 1) Cs 2) Al 3) Ba 4) Be 5) K
36. Which element has the lowest first ionization energy?
 1) K 2) Mg 3) Ca 4) Rb 5) Sr
37. The element with the highest second ionization energy is
 1) Na 2) Mg 3) Ca 4) Be 5) Li
38. X, Y and Z are consecutive elements with increasing atomic numbers. Element X is a noble gas. The symbol of the ion of element Z in its compounds is
 1) Z^- 2) Z^{2-} 3) Z^{3-} 4) Z^+ 5) Z^{2+}
39. Which one of the following has the electron configuration of N^{3-} ?
 1) Cl^- 2) Li^+ 3) Ar 4) Al^{3+} 5) Be^{2+}
40. Which pair of elements shows diagonal relationship?
 1) B, Al 2) Be, Mg 3) Li, Mg 4) C, Si 5) Be, B
41. Consider the following ions.
 a) Cl^- b) N^{3-} c) Al^{3+} d) Be^{2+}
 The ions with the electronic configuration of Ne are
 1) a & b only 2) b & c only 3) c & d only 4) a & d only 5) a, b & c
42. Consider the following pairs of elements.
 a) Be, Mg b) Li, Be c) Be, Al d) Li, Mg
 The pairs of elements that show the diagonal relationship are
 1) a & b 2) b & c 3) c & d 4) a & d 5) a, b & c
43. Which one of the following elements has the highest first ionization energy?
 1) Si 2) P 3) As 4) Ge 5) C
44. Among the following elements, select the largest atom.
 1) Sr 2) Ba 3) Cs 4) Rb 5) Ge
45. The valencies of an element with an outer electron configuration ns^2np^4 are
 1) 1 and 4 2) 2 and 1 3) 2 and 5 4) 2 and 6 5) 1 and 5

Glossary

- Alkali metal** : The Group 1 elements, lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr).
- ක්ෂාරලෝහ** : පළමු කාණ්ඩයේ ලෝහ ලිතියම් (Li), සෝඩියම් (Na), පොටෑසියම් (K), රුබිඩියම් (Rb), සීසියම් (Cs), ෆ්‍රැන්සියම් (Fr).
- කාරුඤ්ඤ** : ලිතියම් (Li), සෝඩියම් (Na), පොටෑසියම් (K), රුබිඩියම් (Rb), සීසියම් (Cs) மற்றும் பிரான்சியம் (Fr) ஆகியன கூட்டம் 1 மூலகங்கள் ஆகும்.
- Alkaline earth metal** : The Group 2 elements, beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra).
- ක්ෂාරපාංශු ලෝහ** : දෙවන කාණ්ඩයේ ලෝහ බෙරිලියම් (Be), මැග්නීසියම් (Mg), කැල්සියම් (Ca), ස්ට්‍රොන්ටියම් (Sr), බේරියම් (Ba), සහ රේඩියම් (Ra).
- කාරමණ්ඤ්ඤ** : பெரிலியம் (Be), மக்னீசியம் (Mg), கல்சியம் (Ca), எஸ்ரொஞ்சியம் (Sr), பேரியம் (Ba), ரேடியம் (Ra) ஆகியன கூட்டம் 2 மூலகங்களாகும்.
- Atom** : The smallest object that retains properties of an element. It is composed of electrons and a nucleus containing protons and neutrons.
- පරමාණුව** : මූලද්‍රව්‍යයක පවතින, තව දුරටත් බෙදා වෙන් කළ නොහැකි කුඩාම අංශුව
- அணு** : ஒரு மூலகத்தின் இயல்புகளைக் கொண்டிருக்கின்ற மிகச்சிறிய பொருள். இது புரோத்திரன், நியூத்திரன் ஆகியவற்றைக் கொண்டிருக்கின்ற கருவையும் இலத்திரன்களையும் கொண்டிருக்கின்றது.
- Block** : A region of the periodic table that corresponds to the type of subshell (*s*, *p*, *d*, or *f*) being filled during the Aufbau construction of electron configurations.
- ගොනුව** : අවුල්ලා ගොඩනැංවීමේ මූලධර්මයට අනුව උපකාක්ෂිකවල ඉලෙක්ට්‍රෝන පිරීම අනුව ආවර්තික වගුවේ ඇති ප්‍රදේශය
- தொகுதி** : இலத்திரன் நிலையமைப்பு Aufbau கொள்கைப்படி பூர்த்தி செய்யப்பட்டிருக்கின்ற போது உப ஒட்டுவகைகளைக் (s,p,d,f) கொண்டிருக்கின்ற ஆவர்த்தன அட்டவணையின் ஒரு பகுதியாகும்.
- Electron affinity** : The amount of energy released when an electron is added to a neutral gaseous atom.
- ඉලෙක්ට්‍රෝන බන්ධුතාවය** : වායුමය උදාසීන පරමාණුවකට ඉලෙක්ට්‍රෝනයක් ලබා දීමේ දී නිකුත් වන ශක්ති ප්‍රමාණය
- இலத்திரன் நாட்டம்** : நடுநிலையான வாயுநிலை அணுவிற்கு இலத்திரனை சேர்க்கும் போது வெளியேற்றப்பட்ட சக்தியின் அளவு.
- Electronegativity** : The tendency of an atom to attract electrons to itself when combined in a molecule.
- විද්‍යුත් ඝෘණතාවය** : අණුවක ඇති එක් පරමාණුවක් සතු, එයට බැඳී ඇති අනෙක් පරමාණුවට සාපේක්ෂව, ඉලෙක්ට්‍රෝන ඇද ගැනීමේ හැකියාව
- மின்னெதிரியல்பு** : மூலக்கூறுகள் இணையும் போது அணு இலத்திரன்களை தன்னை நோக்கி கவரும் தன்மை.
- Element** : A substance consisting of only one type of atoms.
- මූලද්‍රව්‍ය** : එක ම ආකාරයේ පරමාණු වලින් සමන්විත ද්‍රව්‍ය
- மூலகம்** : ஒரு வகையான அணுக்களை மட்டும் கொண்டிருக்கின்ற பதார்த்தம்.

First ionization energy : The energy needed to remove an electron from an isolated, neutral gaseous atom.

පළමු අයනීකරණ ශක්තිය : නිදහස්, උදාසීන, වායුමය පරමාණුවකට බැඳී ඇති ඉලෙක්ට්‍රෝනය ඉවත් කිරීමට අවශ්‍ය ශක්තිය

முதலாவது அயனாக்கற்சக்தி : தனிப்படுத்தப்பட்ட நடுநிலையான வாயுநிலை அணுவிலிருந்து இலத்திரனை நீக்குவதற்கு தேவையான சக்தி.

Group : A vertical column on the periodic table.

කාණ්ඩය : ආවර්තිතා වගුවේ සිරස් තීරු

கூட்டம் : ஆவர்த்தன அட்டவணையின் நிலைக்குத்து நிரல்.

Halogen : Elements of Group 18. Fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At) are known at this time.

හැලජන : 18 වන කාණ්ඩයේ මූලද්‍රව්‍ය ෆ්ලෝරීන් (F), ක්ලෝරීන් (Cl), බ්‍රෝමීන් (Br), අයඩීන් (I), සහ ඇස්ටටීන් (At)

அலசன்கள் : புளோரின் (F), புரோமின் (Br), அயடீன் (I), அஸ்டடைன் (At) ஆகியன இந்த நேரத்தில் அறியப்பட்ட கூட்டம் 18 இன் மூலகங்களாகும்.

Ionic radius : The radii of anions and cations in crystalline ionic compounds, as determined by consistently partitioning the center-to-center distance of ions in those compounds.

අයනික අරය : ස්ඵටිකීය අයනික සංයෝගයක ඇත්‍යයන සහ කැටායනවල කේන්ද්‍ර අතර දුර

අයනාරා : பளிங்குருவான அயன் சேர்வைகளின் மறை அயன், நோ அயன்களிற்கான ஆரை அந்த சேர்வைகளின் தொடர்ச்சியாகப் பங்கிடப்படுகின்ற அயன்களின் மையத்துக்குமான தூரத்தினால் தீர்மானிக்கப்படுகின்றது.

Ionisation energy : The energy needed to remove an electron from a gaseous atom or ion.

අයනීකරණ ශක්තිය : වායුමය පරමාණුවකින් හෝ අයනයකින් ඉලෙක්ට්‍රෝනයක් ඉවත් කිරීමට අවශ්‍ය ශක්තිය යි.

අයනාකරණ ශක්තිය : වායුநிலை அணு அல்லது அயனிலிருந்து இலத்திரனை வெளியேற்றுவதற்கு தேவையான சக்தி.

Liquefaction : The act or process of turning a gas into a liquid.

ද්‍රව කිරීම : වායුවක් ද්‍රවයක් බවට පත් කිරීමේ ක්‍රියාවලිය

திரவமாக்கல் : வாயுவானது திரவமாக மாற்றப்படும் ஒரு பொறிமுறை.

Mass number : The total number of protons and neutrons in the nucleus of an atom.

ස්කන්ධ ක්‍රමාංකය : පරමාණුවක න්‍යෂ්ටියේ ඇති ප්‍රෝටෝන හා නියුට්‍රෝනවල එකතුව

திணிவெண் : அணுவொன்றின் கருவிலுள்ள புரோதன், நியூத்திரன்களின் மொத்த எண்ணிக்கை.

Metal : It is a substance that conducts heat and electricity, it is shiny and can be hammered into sheets or drawn into wire. Metals lose electrons easily to form cations.

ලෝහය : විදුලිය සහ තාපය සන්නයනය කළ හැකි, දිලිසෙන සුළු, ඇදීමේ හෝ තැලීමේ ගුණ ඇති ද්‍රව්‍යයකි. ඉතා පහසුවෙන් ඉලෙක්ට්‍රෝන නිදහස් කර කැටායන සාදයි.

உலோகம் : வெப்பம், மின் ஆகியவற்றைக் கடத்துகின்ற ஒரு பதார்த்தமாகும், பளபளப்பானது அடித்து தாளாகவும் அல்லது இழுத்து வயராகவும் மாற்ற

முடியும். உலோகங்கள் இலகுவாக இலத்திரன்களை இழந்து நேர் அயன்களை உருவாக்கும்.

Metalloid : It is a substance that exhibits properties in between to metals and non-metals.

லோகாலோக : லோகமீய சன டலோகமீய டுன் டநர ஡ுடி டுதி டுவ

உலோகப்போலி : உலோகத்திற்கும் அலோகத்திற்கும் இடையிலான இயல்புகளை காட்டுகின்ற பதார்த்தமாகும்.

Nitrogen Fixation : The process by which free nitrogen from the air is combined with other elements to form inorganic compounds.

நடிகுபன் திர திரீம : வாகடே பவநன திடனச தடிகுபன் வெனன் இலடுவா சமற ஸகநு கர் டகாவநிக சமடேய டுடிமே தியாவலேய டி.

N₂ பதித்தல் : வளிமண்டலத்தில் இருக்கும் தனித்த N₂ வேறு முலகங்களுடன் இணைவதன் முலம் அசேதன சேர்வைகளை உருவாக்கும் பொறிமுறை.

Nonmetal : It is a substance that conducts heat and electricity poorly, is brittle or waxy or gaseous, and cannot be hammered into sheets or drawn into wire. Nonmetals gain electrons easily to form anions.

டலோக : விடுலேய சன நாகப சன்நடயன டுர்வல, நிடேன டுப டோ வாகுமீய வன, டுடிமே டோ துலேமே டுன் துதி டுவடகி. பனடுவென் டுலேக்டுரேன் டலா டென டுநாகன சாடி.

அல்லுலோகம் : ஓரு பதார்த்தமானது வெப்பம், மின் ஆகியவற்றை குறைந்தளவு கடத்துகின்ற , நொருங்கும் தன்மையானது, மெழுகு அல்லது வாயு நிலையிலுள்ளது மற்றும் அடித்து தாளாகவும் அல்லது இழுத்து வயராகவும் மாற்ற முடியாதது. அல்லுலோகங்கள் இலகுவாக இலத்திரன்களை ஏற்று மறை அயன்களை உருவாக்கும்.

Neutron : A neutral particle in the nucleus of an atom.

நிடுடுரேய : பரமாளுவக நாகசீவெ டுதி டுடிசீன டுடிவ

நியுத்திரன் : அணுக்கருவினுள் உள்ள நடுநிலையான துணிக்கைகள்

Orbital : A region around the nucleus where there is a high probability of finding an electron.

காக்மீக : டுலேக்டுரேன் புவநீமே வுவி சமீகாவிகாவ டுதி நாகசீவெ வலா வு புடேடய

ஓழுக்கு : அதிகளவு நிகழ்தகவு இலத்திரன் காணப்படும் கருவை சூழவுள்ள பகுதி

Oxidation number: A number assigned to each atom to help keep track of the electrons during a redox-reaction.

மக்சிகரன் டுகய : ரெவொக்சீ ப்ரிகியாவக டி டுடில வன டுலேக்டுரேன் ப்ரமாளய பன்வன டுலச ஸக ஸக பரமாளவெ சநு டுகய

ஓட்சியேற்ற எண் : ஓட்சியேற்ற தாழ்த்தல் தாக்கத்தின் போது சுவட்டில் இலத்திரன்களை வைத்திருப்பதற்கு உதவுகின்ற ஒவ்வொரு அணுவிற்கும் ஒதுக்கப்பட்ட ஒரு இலக்கம்.

Paramagnetic : A substance that shows magnetic properties when placed in a magnetic field.

டுடிபுமீக : ப்ரமீக க்சேநுயக் நுல துடி விடு ப்ரமீக டுக்சன் பன்வன டுவடி.

பரகாந்தத்திற்குரிய : காந்த புலத்திலுள்ள போது காந்த இயல்பைக் காட்டும் ஒரு பதார்த்தம்.

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