

Periodic Table

■ Development of Periodic Table:

→ In the nineteenth century, when chemists had only a vague idea of atoms and molecules and did not know of the existence of electrons and protons, they devised the periodic table using their knowledge of atomic masses.

→ In 1864 the English chemist John Newlands noticed that when the elements were arranged in order of atomic mass, every eighth element had similar properties.

Newlands referred to this peculiar relationship as the law of octaves.

However, this “law” turned out to be inadequate for elements beyond calcium, and Newlands’s work was not accepted by the scientific community.

Periodic Table

■ Development of Periodic Table:

→ In 1869 the Russian chemist Dmitri Mendeleev and the German chemist Lothar Meyer independently proposed a much more extensive tabulation of the elements based on the regular, periodic recurrence of properties.

→ First, it grouped the elements together more accurately, according to their properties.

Equally important, it made possible the prediction of the properties of several elements that had not yet been discovered.

For example, Mendeleev proposed the existence of an unknown element that he called eka-aluminum and predicted a number of its properties.

Periodic Table

■ Development of Periodic Table:

→ When gallium was discovered four years later, its properties matched the predicted properties of eka-aluminum remarkably well:

→ Mendeleev's periodic table included 66 known elements. By 1900, some 30 more had been added to the list, filling in some of the empty spaces.

→ The early versions had some glaring inconsistencies. For example, the atomic mass of argon (39.95 amu) is greater than that of potassium (39.10 amu). If elements were arranged solely according to increasing atomic mass, argon would appear in the position occupied by potassium in our modern periodic table. But no chemist would place argon, an inert gas, in the same group as lithium and sodium, two very reactive metals.

Periodic Table

■ Development of Periodic Table:

→ This and other discrepancies suggested that some fundamental property other than atomic mass must be the basis of periodicity. This property turned out to be associated with atomic number, a concept unknown to Mendeleev and his contemporaries.

→ With a few exceptions, Moseley found that atomic number increases in the same order as atomic mass. For example, calcium is the twentieth element in order of increasing atomic mass, and it has an atomic number of 20. The discrepancies that had puzzled earlier scientists now made sense. The atomic number of argon is 18 and that of potassium is 19, so potassium should follow argon in the periodic table. **A modern periodic table usually shows the atomic number along with the element symbol.**

Periodic Table

■ Development of Periodic Table:

→ Organizing the Elements Chemists used the properties of elements to sort them into groups. JW. Dobreiner grouped elements into triads. A triad is a set of three elements with similar properties.

→ Mendeleev's Periodic Table, In 1869, a Russian chemist and teacher published a table of the elements. Mendeleev arranged the elements in the periodic table in order of increasing atomic mass.

→ Henry Mosele In 1913, he determined the actual nuclear charge (atomic number) of the elements. He rearranged the elements in order of increasing atomic number.

Periodic Table

Periodic Table of the Elements

Periodic Table of the Elements																	
1 H Hydrogen 1.008																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.066	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.631	33 As Arsenic 74.922	34 Se Selenium 78.971	35 Br Bromine 79.904	36 Kr Krypton 84.798
37 Rb Rubidium 84.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium 98.907	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.414	49 In Indium 114.818	50 Sn Tin 118.711	51 Sb Antimony 121.760	52 Te Tellurium 127.6	53 I Iodine 126.904	54 Xe Xenon 131.294
55 Cs Cesium 132.905	56 Ba Barium 137.328	57-71 Lanthanides	72 Hf Hafnium 178.49	73 Ta Tantalum 180.948	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.217	78 Pt Platinum 195.085	79 Au Gold 196.967	80 Hg Mercury 200.592	81 Tl Thallium 204.383	82 Pb Lead 207.2	83 Bi Bismuth 208.980	84 Po Polonium [208.982]	85 At Astatine 209.987	86 Rn Radon 222.018
87 Fr Francium 223.020	88 Ra Radium 226.025	89-103 Actinides	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [269]	109 Mt Meitnerium [268]	110 Ds Darmstadtium [269]	111 Rg Roentgenium [272]	112 Cn Copernicium [277]	113 Uut Ununtrium unknown	114 Fl Flerovium [289]	115 Uup Ununpentium unknown	116 Lv Livermorium [298]	117 Uus Ununseptium unknown	118 Uuo Ununoctium unknown

Lanthanide Series	57 La Lanthanum 138.905	58 Ce Cerium 140.116	59 Pr Praseodymium 140.908	60 Nd Neodymium 144.243	61 Pm Promethium 144.913	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.925	66 Dy Dysprosium 162.500	67 Ho Holmium 164.930	68 Er Erbium 167.259	69 Tm Thulium 168.934	70 Yb Ytterbium 173.055	71 Lu Lutetium 174.967
Actinide Series	89 Ac Actinium 227.028	90 Th Thorium 232.038	91 Pa Protactinium 231.036	92 U Uranium 238.029	93 Np Neptunium 237.048	94 Pu Plutonium 244.064	95 Am Americium 243.061	96 Cm Curium 247.070	97 Bk Berkelium 247.070	98 Cf Californium 251.080	99 Es Einsteinium [254]	100 Fm Fermium 257.095	101 Md Mendelevium 258.1	102 No Nobelium 259.101	103 Lr Lawrencium [262]

Alkali Metal	Alkaline Earth	Transition Metal	Basic Metal	Semimetal	Nonmetal	Halogen	Noble Gas	Lanthanide	Actinide
--------------	----------------	------------------	-------------	-----------	----------	---------	-----------	------------	----------

Periodic Table

■ Components Periodic Table?

- The periodic table is made up of rows of elements and columns.
- An element is identified by its chemical symbol.
- The number above the symbol is the atomic number.
- The number below the symbol is the rounded atomic weight of the element.
- A row is called a period and a column is called a group.
- There are seven period and eighteen groups in the periodic table.
- The elements can be grouped into three broad classes based on their general properties.
- Three classes of elements are Metals, Nonmetals, and Metalloids.
- Across a period, the properties of elements become less metallic and more nonmetallic.

Periodic Table

Group

- Columns of elements are called groups or families.
- Elements in each group have similar but not identical properties.
- For example, lithium (Li), sodium (Na), potassium (K), and other members of group I are all soft, white, shiny metals.
- All elements in a group have the same number of valence electrons.

Period

- Each horizontal row of elements is called a period.
- The elements in a period are not alike in properties.
- In fact, the properties change greatly across even given row.
- The first element in a period is always an extremely active solid. The last element in a period, is always an inactive gas.

Periodic Table

- Determination of the position of elements in periodic table from the electronic configuration.

→ **Determining the period number:**

The number of the outermost main energy level ---> Period number of that given element.

For example, Li (3) ---> $1s^2 2s^1$ Here, period number of Li is 2.

Mg (12) ---> $1s^2 2s^2 2p^6 3s^2$. Here, period number of Mg is 3.

Ca (20) ---> $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ Here, period number of Ca is 4.

Periodic Table

- Determination of the position of elements in periodic table from the electronic configuration.

→ **Determining the group number:**

If Last Electron enters in s orbital – it is a s block element

If p orbital – it is a p block element

If d orbital – it is a d block element

If it is an element of s block ---> The number of electrons in the valence shell. For example; Be (4) ---> $1s^2 2s^2$

Francium (87) ---> $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^1$

Periodic Table

■ Determination of the position of elements in periodic table from the electronic configuration.

→ **Determining the group number:**

If Last Electron enters in s orbital – it is a s block element

If p orbital – it is a p block element

If d orbital – it is a d block element

If it is an element of s block ---> The number of electrons in the valence shell. For example; Be (4) ---> $1s^2 2s^2$

Francium (87) ---> $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^1$

If it is an element of p block ---> The number of electrons in valence shell (summation of outermost electrons present in s and p orbitals)

10: N (7) ---- > $1s^2 2s^2 2p^3$

Periodic Table

■ Properties of Metals:

- Metals are good conductors of heat and electricity.
- Metals are shiny.
- Metals are ductile (can be stretched into thin wires).
- Metals are malleable (can be pounded into thin sheets).
- A chemical property of metal is its reaction with water which results in corrosion.
- Solid at room temperature except Hg.



Steel

Periodic Table

■ Properties of Non-Metals:

- Non-metals are poor conductors of heat and electricity.
- Non-metals are not ductile or malleable.
- Solid non-metals are brittle and break easily.
- They are dull.
- Many non-metals are gases.

Sulfur



Periodic Table

■ Properties of Metalloids:

- Metalloids (metal-like) have properties of both metals and non-metals.
- They are solids that can be shiny or dull.
- They conduct heat and electricity better than non-metals but not as well as metals.
- They are ductile and malleable.

Silicon



Periodic Relationships Among the Elements

■ Effective Nuclear Charge:

→ The effective nuclear charge (Z_{eff}) is the nuclear charge felt by an electron when both the actual nuclear charge (Z) and the repulsive effects (shielding) of the other electrons are taken into account.

$$Z_{\text{eff}} = Z - \sigma$$

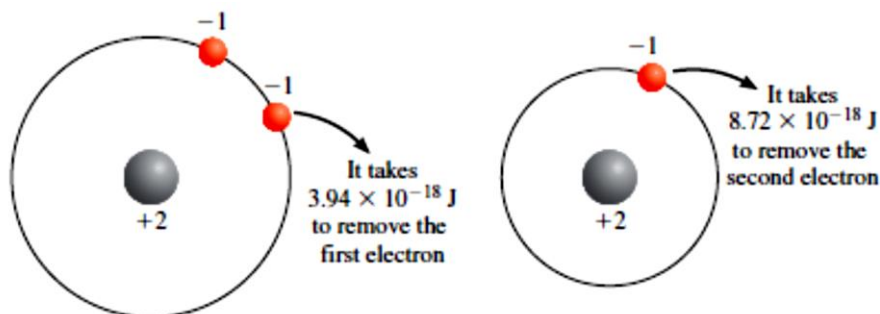
where σ (sigma) is called the shielding constant (also called the screening constant). The shielding constant is greater than zero but smaller than Z .

The presence of other electrons in an atom reduces the electrostatic attraction between a given electron and the positively charged protons in the nucleus.

Periodic Relationships Among the Elements

■ Effective Nuclear Charge:

→ Moving across the period is a greater effective nuclear charge felt by the valence electrons because the added electron is a valence electron and valence electrons do not shield each other well.



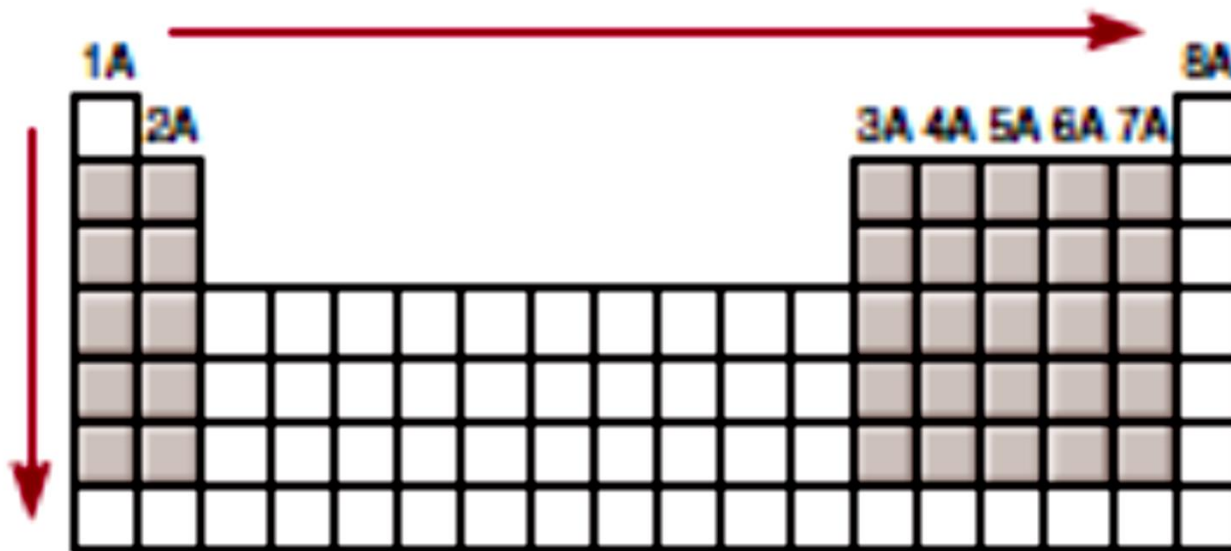
→ Moving from left to right, number of core electrons ($1s^2$) remains constant while the nuclear charge increases.

→ Because the core electrons are, on average, closer to the nucleus than valence electrons, core electrons shield valence electrons much more than valence electrons shield one another.

Periodic Relationships Among the Elements

■ Effective Nuclear Charge:

→ The effective nuclear charge also increases as we go down a particular periodic group. However, because the valence electrons are now added to increasingly large shells as n increases, the electrostatic attraction between the nucleus and the valence electrons actually decreases.

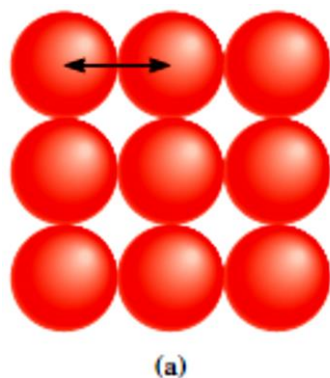


Periodic Relationships Among the Elements

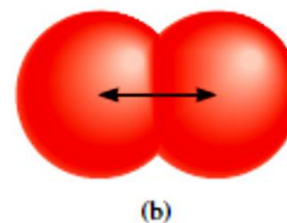
■ Atomic Radius:

→ A number of physical properties, including density, melting point, and boiling point, are related to the sizes of atoms, but atomic size is difficult to define.

→ **we define the size of an atom in terms of its atomic radius, which is one-half the distance between the two nuclei in two adjacent metal atoms or in a diatomic molecule.**

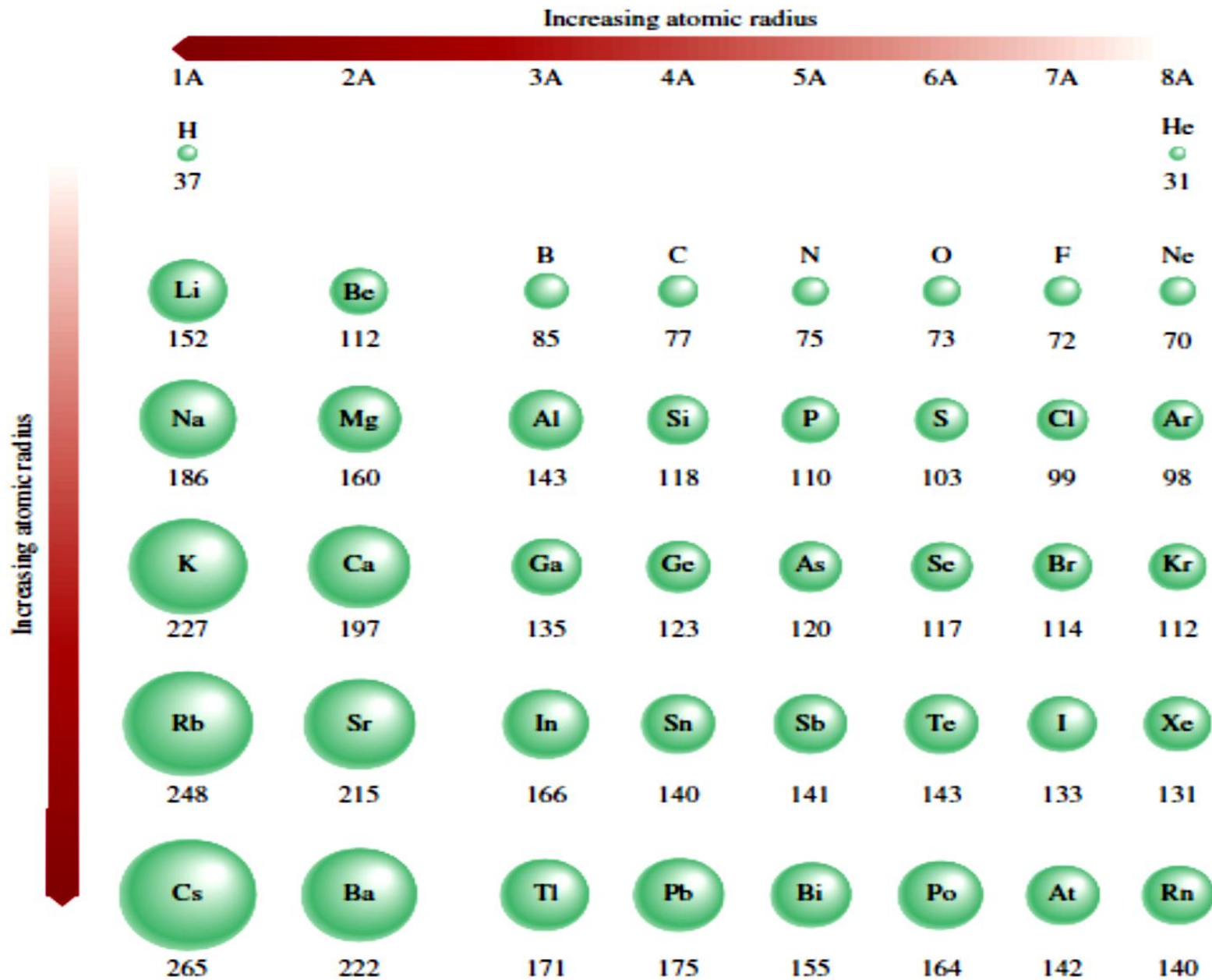


(a) In metals such as beryllium, the atomic radius is defined as one-half the distance between the centers of two adjacent atoms.



(b) For elements that exist as diatomic molecules, such as iodine, the radius of the atom is defined as one-half the distance between the centers of the atoms in the molecule.

Periodic Relationships Among the Elements



Periodic Relationships Among the Elements

■ Atomic Radius:

→ Periodic trends are clearly evident. Consider the second-period elements. Because the effective nuclear charge increases from left to right, the added valence electron at each step is more strongly attracted by the nucleus than the one before. Therefore, we expect and indeed find the atomic radius decreases from Li to Ne.

→ Within a group, we find that atomic radius increases with atomic number. For the alkali metals in Group 1, the valence electron resides in the ns orbital. Because orbital size increases with the increasing principal quantum number n , the size of the atomic radius increases even though the effective nuclear charge also increases from Li to Cs.

Periodic Relationships Among the Elements

■ Ionic Radius:

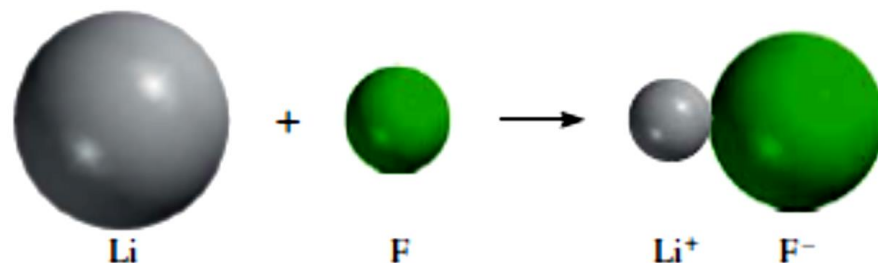
→ Ionic radius is the radius of a cation or an anion . It can be measured by X-ray diffraction. Ionic radius affects the physical and chemical properties of an ionic compound. For example, the three-dimensional structure of an ionic compound depends on the relative sizes of its cations and anions.

→ When a neutral atom is converted to an ion, we expect a change in size. If the atom forms an anion, its size (or radius) increases, because the nuclear charge remains the same but the repulsion resulting from the additional electron(s) enlarges the domain of the electron cloud.

Periodic Relationships Among the Elements

■ Ionic Radius:

→ On the other hand, removing one or more electrons from an atom reduces electron-electron repulsion but the nuclear charge remains the same, so the electron cloud shrinks, and the cation is smaller than the atom.



→ We can see parallel trends between atomic radii and ionic radii. For example, from top to bottom both the atomic radius and the ionic radius increase within a group.

Periodic Relationships Among the Elements

■ Ionic Radius:

→ If we examine isoelectronic ions, we find that cations are smaller than anions. For example, Na^+ is smaller than F^- . Both ions have the same number of electrons, but Na ($Z=11$) has more protons than F ($Z=9$). The larger effective nuclear charge of Na^+ results in a smaller radius.

→ Focusing on isoelectronic cations, we see that the radii of tripositive ions (ions that bear three positive charges) are smaller than those of dipositive ions (ions that bear two positive charges), which in turn are smaller than unipositive ions (ions that bear one positive charge).

Periodic Relationships Among the Elements

■ Ionic Radius:

→ This trend is nicely illustrated by the sizes of three isoelectronic ions in the third period: Al^{3+} , Mg^{2+} , and Na^{+} . The Al^{3+} ion has the same number of electrons as Mg^{2+} , but it has one more proton. Thus the electron cloud in Al^{3+} is pulled inward more than that in Mg^{2+} . The smaller radius of Mg^{2+} , compared with that of Na^{+} can be similarly explained.

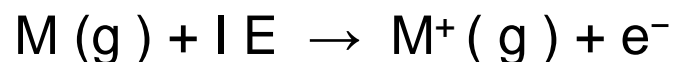
→Turning to isoelectronic anions, we find that the radius increases as we go from ions with uni-negative charge (-) to those with di-negative charge (2-), and so on. Thus, the oxide ion is larger than the fluoride ion because oxygen has one fewer proton than fluorine; the electron cloud is spread out more in O^{2-} .

Periodic Relationships Among the Elements

■ Ionisation Energy:

→ The process of removing an electron from an isolated atom to form a positive ion is called ionisation. Energy will be required to remove an electron from the atom against the force of attraction of the nucleus.

→ **The ionisation energy (IE) of an element is defined as the energy needed to remove a single electron from an atom of the element in the gaseous state.**



→ Since one, two or more electrons may be removed from the same atom, one after the other, we have as many ionisation energies of the element.

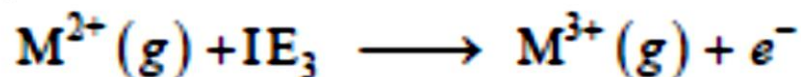
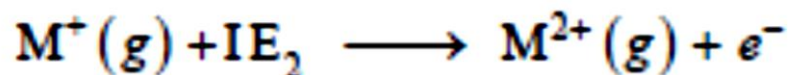
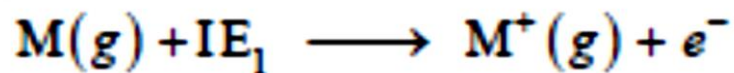
Periodic Relationships Among the Elements

■ Ionisation Energy:

→ The First ionisation energy (IE_1), is the energy needed to remove the first electron from the gaseous atom M to form M^+ ion.

The Second ionisation energy (IE_2), is the energy needed to remove a second electron, from the gaseous M^+ ion to form M^{2+} ion.

Higher ionisation energies can be defined in the same way. We can depict the first, second and third ionisation energies in the form of equations as :



Periodic Relationships Among the Elements

■ Ionisation Energy:

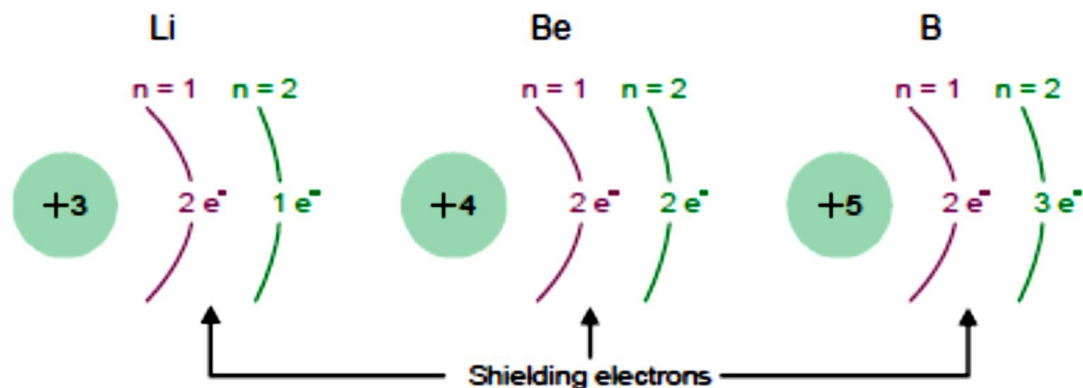
→ Increase Across a Period

As we pass from left to right in a period, the first ionisation energy shows a steady increase. Thus in Period 2 from Li to N, we have;

	Li	Be	B	C	N
$IE_1 (\text{kJ mol}^{-1})$	525	906	805	1090	1400

Explanation

The outer-shell electrons in the elements of the same period are arranged in the same shell. For example, the build up of electrons in Period 2 from Li to B is



Periodic Relationships Among the Elements

■ Ionisation Energy:

Decrease Down a Group:

→ In the elements of a vertical Group of the Periodic table, the number of outer shell electrons is the same. But the following changes are noted from top to bottom.

(1) The principal quantum number n containing the valence electrons increases.

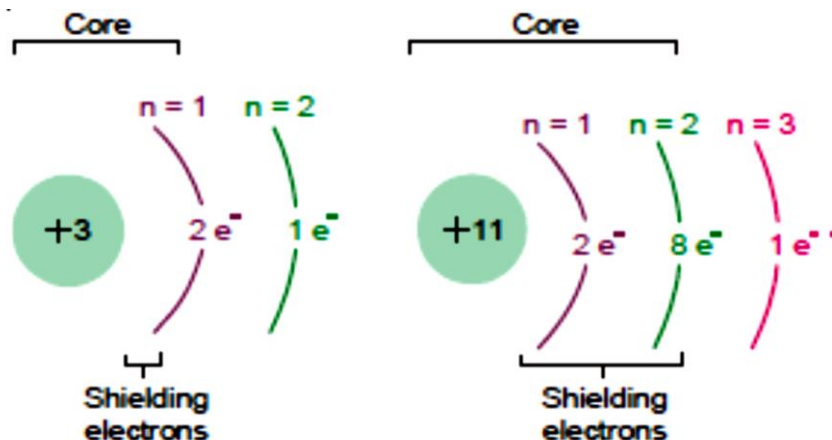
(2) The nuclear charge (At. No.) increases.

(3) The number of electrons in the inner shells (shielding electrons) increases.

The net result of these changes is that the first ionisation energies down a group record a progressive decrease

	Li	Na	K	Rb	Cs
IE_1 (kJ mol ⁻¹)	525	500	424	408	382

Periodic Relationships Among the Elements



Decrease Down a Group:

→ Lithium and sodium both have one outer-shell electron. The number of shielding electrons in sodium is 10 while in lithium it is 2. If we assume that the inner shell electrons provide hundred percent screening, the core charge attracting the outer-shell electron would be:

	Nuclear charge	Shielding electrons	Core charge attracting outer-shell electron
Li	3	2	$3 - 2 = +1$
Na	11	2, 8	$11 - 10 = +1$

Periodic Relationships Among the Elements

■ Ionisation Energy:

Decrease Down a Group:

→ Thus the same net charge (+ 1) attracts the outer-shell electrons to the core. But the distance of the outer electron from the nucleus is greater in Na ($n = 3$) than in Li ($n = 1$). Therefore the force of attraction between the outer electron and the core will be less in Na than in Li. That explains the lower IE of Na compared to Li. By the same line of argument, the decrease trend in IE from element to element while going down a Group can be justified.

→ Regular Discontinuities:

IE_1 of B is less than that of Be and the IE_1 of O is less than that of N.

Periodic Relationships Among the Elements

(a) The electronic configuration of Be and B are :



The 2p orbital electron of B is already higher in energy than the 2s orbital electron. Therefore the removal of electron from B requires less energy and its IE_1 is lower.

(b) The electronic configuration of N and O is :



The 2p orbitals may be represented as:



Whenever two electrons occupy a particular orbital, they repel each other. As a result it is easier to remove one of the paired 2p electrons from O than it is to remove an unpaired electron from N atom.

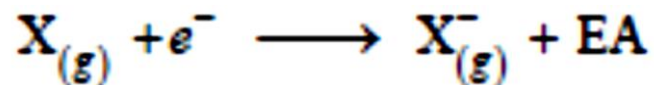
Thus IE_1 of O is lower than that of N.

Periodic Relationships Among the Elements

■ Electron Affinity:

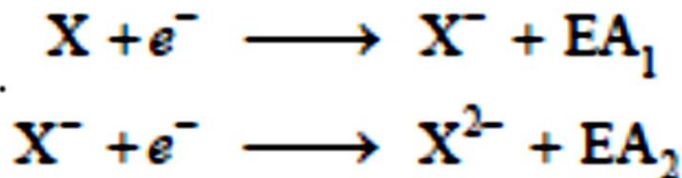
A neutral atom can accept an electron to form negative ion. In this process, in general, energy is released.

Electron affinity (EA) of an element is the amount of energy released when an electron is added to a gaseous atom to form an anion.



The energy involved in the addition of the first electron is called first-electron affinity;

The energy involved in the addition of a second electron is called second-electron affinity; and so on. Thus,



Periodic Relationships Among the Elements

■ Electron Affinities:

Trends in Electron Affinities:

→ The first-electron affinities of elements in the Periodic table are expected to show trends analogous to those of ionisation energies.

→ (a) Increase Across a Period:

As we proceed from left to right, the general trend is the increase of electron affinities. Be, N and Ne are exceptions.

Be	B	C	N	O	F	Ne
-240	23	123	0.00	142	323	-29 kJ mol ⁻¹

Elements having relatively stable electronic configurations find it difficult to accept an electron readily.

Periodic Relationships Among the Elements

■ Electron Affinities:

Trends in Electron Affinities:

The atom of Be has the configuration $1s^2 2s^2$. The $1s$ subshell is completely filled and, therefore, the electron being added must go to a subshell of considerably higher energy. This gives rise to negative electron affinity for Be.

The atom of N ($1s^2 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$) has half-filled $2p$ subshells, a condition of extra stability. Therefore the electron affinity of N would be less than expected.

The electron affinity of Neon is low because it has a stable outer-shell octet. Its atom shows little tendency to start a new shell.

Periodic Relationships Among the Elements

■ Electron Affinities:

Trends in Electron Affinities:

→ (*b*) Decrease Down a Group:

The electron affinities show a general decrease from top to bottom. This is so because the valence shell is progressively farther from the nucleus. The value for fluorine, however, is out of line as it has a smaller atomic size than that of chlorine.

F	Cl	Br	I	
332.6	349	324.7	296	kJ mol^{-1}

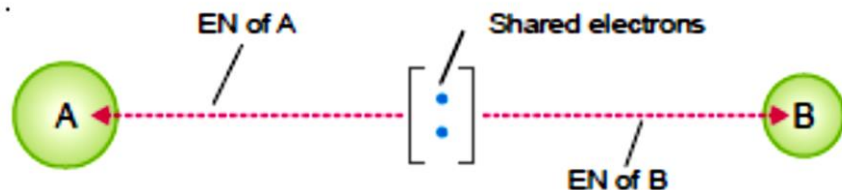
→ The second electron affinity of an element is always negative. This is on account of repulsion between the electron being added and the already negatively charged atom.

Periodic Relationships Among the Elements

■ Electronegativity:

→ In a molecule A – B the electrons forming the covalent bond are attracted by atom A as well as by B. This attraction is measured in terms of what we call electronegativity, EN.

It may be defined as : **The attraction exerted by an atom on the electron pair bonding it to another atom by a covalent bond.**



→ It is evident that an atom of high electronegativity will attract the shared electron pair away from one of lower electronegativity. Thus the former atom will acquire a partial negative charge while the other atom will get a partial positive charge.

Periodic Relationships Among the Elements

■ Electronegativity::

Trends in Electronegativity:

→ (1) Increase across a Period

The values of electronegativities increase as we pass from left to right in a Period. Thus for Period 2 we have

Li	Be	B	C	N	O	F
1.0	1.5	2.0	2.5	3.0	3.5	4.0

→ This is so because the attraction of bonding electrons by an atom increases with increase of nuclear charge and decrease of atomic radius. Both these factors operate as we move to the right in a Period.

Periodic Relationships Among the Elements

■ Electronegativity::

Trends in Electronegativity:

→ (2) Decrease down a Group

The electronegativities of elements decrease from top to bottom in a Group. Thus for Group VII;

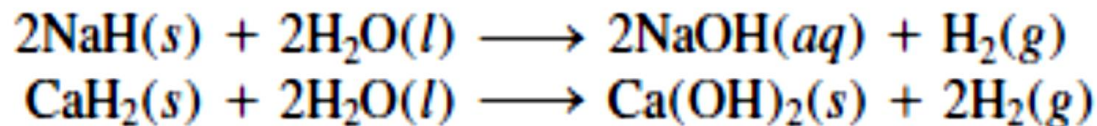
F	Cl	Br	I
4.0	3.0	2.8	2.5

→ The decrease trend is explained by more shielding electrons and larger atomic radius as we travel down a Group.

Periodic Relationships Among the Elements

→ Hydrogen (1s¹):

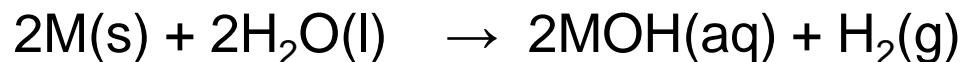
Like the alkali metals, it has a single s valence electron and forms a unipositive ion (H⁺) , which is hydrated in solution. On the other hand, hydrogen also forms the hydride ion (H⁻) in ionic compounds such as NaH and CaH₂ . In this respect, hydrogen resembles the halogens, all of which form uninegative ions (F⁻ , Cl⁻ , Br⁻ , and I⁻) in ionic compounds. Ionic hydrides react with water to produce hydrogen gas and the corresponding metal hydroxides:



Periodic Relationships Among the Elements

→ Group 1 Elements (ns^1 , $n \geq 2$)

All of alkali metal elements have low ionization energies and therefore a great tendency to lose the single valence electron. In fact, in the vast majority of their compounds they are unipositive ions. These metals are so reactive that they are never found in the pure state in nature. They react with water to produce hydrogen gas and the corresponding metal hydroxide:

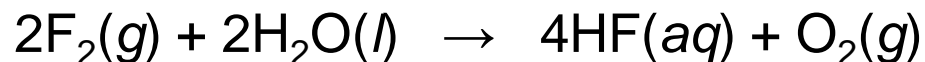


Periodic Relationships Among the Elements

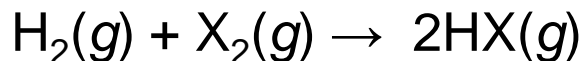
→ **Group 7A or 17 Elements ($ns^2 np^5$, $n \geq 2$)**

All the halogens are non-metals with the general formula X_2 , where X denotes a halogen element. Because of their great reactivity, the halogens are never found in the elemental form in nature.

Fluorine is so reactive that it attacks water to generate oxygen:



The halogens react with hydrogen to form hydrogen halides:



When this reaction involves fluorine, it is explosive, but it becomes less and less violent as we substitute chlorine, bromine, and iodine.

Periodic Relationships Among the Elements

→ Group 18 or 8A Elements ($ns^2 np^6$, $n \geq 2$)

All noble gases exist as monatomic species. Their atoms have completely filled outer ns and np subshells, which give them great stability. The Group 8A ionization energies are among the highest of all elements, and these gases have no tendency to accept extra electrons. For years these elements were called inert gases, and rightly so. Until 1963 no one had been able to prepare a compound containing any of these elements.

Since then, a number of xenon compounds (XeF_4 , XeO_3 , XeO_4 , XeOF_4) and a few krypton compounds (KrF_2 , for example) have been prepared. However, their compounds do not have any major commercial applications, and they are not involved in natural biological processes. No compounds of helium and neon are known.