

General chemistry

Chapter 5 Periodic table

Periodic table

- **Periodic table**—a chart in which elements are arranged by atomic number (shown above the element symbol) in horizontal rows called *periods* and in vertical columns known as *groups or families*.
- The periodic table is a handy tool that correlates the properties of the elements in a systematic way and helps us to make predictions about chemical behavior.

Notes

- A **block** of the periodic table is a set of chemical elements having their differentiating electrons predominately in the same type of atomic orbital.
- A **differentiating electron** is the electron that differentiates an element from the previous one.

Periodic Classification of the Elements

- Elements are grouped according to their outer-shell electron configurations, which account for their similar chemical behavior.
- According to the type of subshell being filled by the differentiating electron, the elements can be divided into categories:
 1. **The Representative elements** (also called main group elements) the elements in Groups 1A through 7A, in which the differentiating electron is filled in the s subshells (s-block) or in the p subshell (p-block) of the highest principal quantum number.

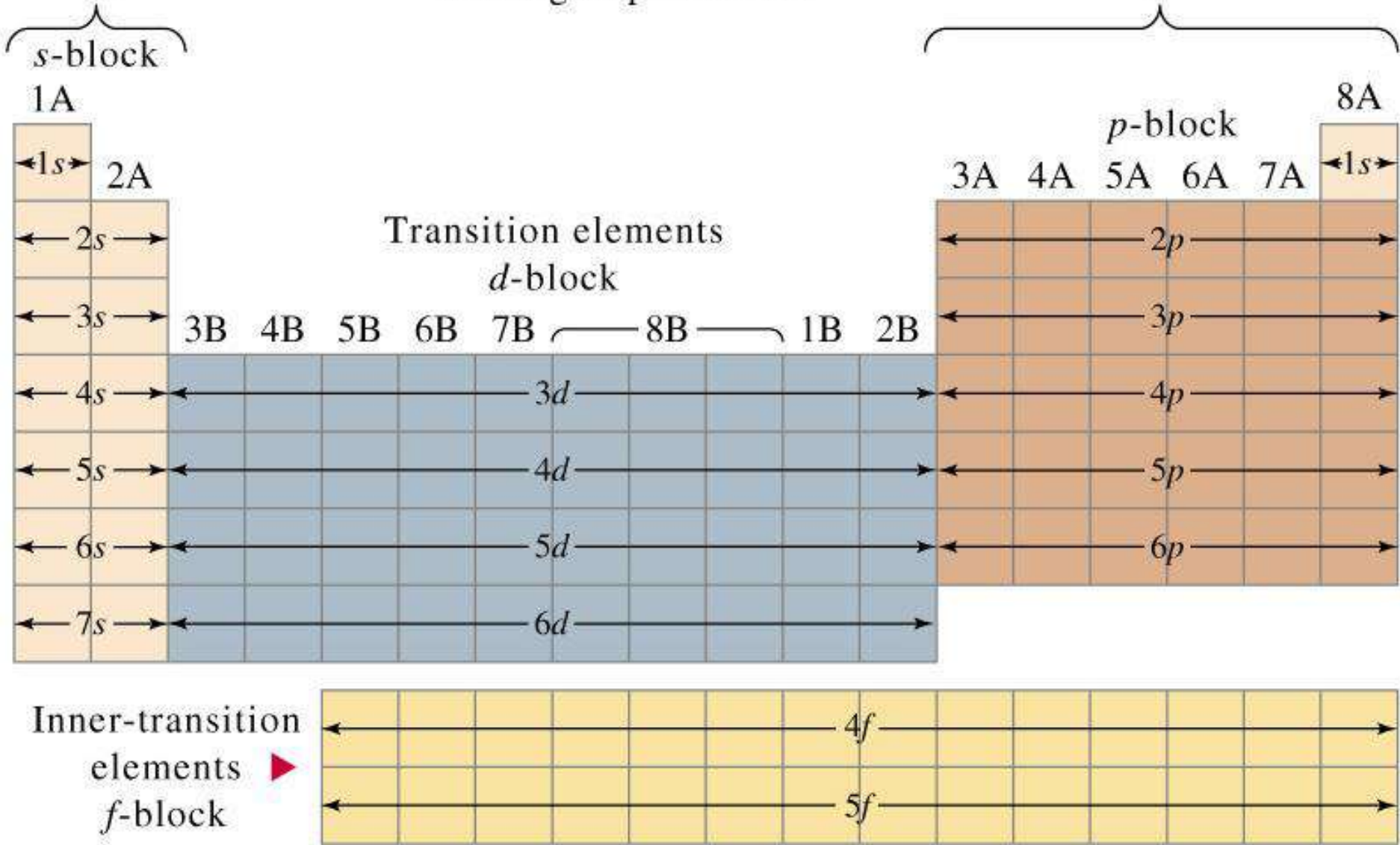
2- The noble gases, with the exception of helium, the noble gases (the Group 8A elements) all have a completely filled s and p subshells of the highest principal quantum number. The electron configurations are $1s^2$ for helium and ns^2np^6 for the other noble gases, in which n is the principal quantum number for the outermost shell.

3- The transition elements (or transition metals), are the elements in Groups 1B and 3B through 8B, in which the differentiating electron is filled in the d subshells (d-block).

4- The Group 2B elements are Zn, Cd, and Hg, which are neither representative elements nor transition metals.

5- The inner transition elements (the lanthanides, and the actinides), in which the differentiating electron is filled in the f subshells (f-block).

Main-group elements



1 1A																	18 8A		
1 H	2 2A													13 3A	14 4A	15 5A	16 6A	17 7A	2 He
3 Li	4 Be													5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B		10	11 1B	12 2B	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar		
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr		
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe		
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn		
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)		

Representative elements

Noble gases

Transition metals

Zinc
Cadmium
Mercury

Lanthanides

Actinides

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

	1 1A																18 8A	
1	1 H $1s^1$	2 2A											13 3A	14 4A	15 5A	16 6A	17 7A	2 He $1s^2$
2	3 Li $2s^1$	4 Be $2s^2$											5 B $2s^2 2p^1$	6 C $2s^2 2p^2$	7 N $2s^2 2p^3$	8 O $2s^2 2p^4$	9 F $2s^2 2p^5$	10 Ne $2s^2 2p^6$
3	11 Na $3s^1$	12 Mg $3s^2$	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B		10 10B	11 11B	12 12B	13 Al $3s^2 3p^1$	14 Si $3s^2 3p^2$	15 P $3s^2 3p^3$	16 S $3s^2 3p^4$	17 Cl $3s^2 3p^5$	18 Ar $3s^2 3p^6$
4	19 K $4s^1$	20 Ca $4s^2$	21 Sc $4s^2 3d^1$	22 Ti $4s^2 3d^2$	23 V $4s^2 3d^3$	24 Cr $4s^1 3d^5$	25 Mn $4s^2 3d^5$	26 Fe $4s^2 3d^6$	27 Co $4s^2 3d^7$	28 Ni $4s^2 3d^8$	29 Cu $4s^1 3d^{10}$	30 Zn $4s^2 3d^{10}$	31 Ga $4s^2 4p^1$	32 Ge $4s^2 4p^2$	33 As $4s^2 4p^3$	34 Se $4s^2 4p^4$	35 Br $4s^2 4p^5$	36 Kr $4s^2 4p^6$
5	37 Rb $5s^1$	38 Sr $5s^2$	39 Y $5s^2 4d^1$	40 Zr $5s^2 4d^2$	41 Nb $5s^1 4d^4$	42 Mo $5s^1 4d^5$	43 Tc $5s^2 4d^5$	44 Ru $5s^1 4d^7$	45 Rh $5s^1 4d^8$	46 Pd $4d^{10}$	47 Ag $5s^1 4d^{10}$	48 Cd $5s^2 4d^{10}$	49 In $5s^2 5p^1$	50 Sn $5s^2 5p^2$	51 Sb $5s^2 5p^3$	52 Te $5s^2 5p^4$	53 I $5s^2 5p^5$	54 Xe $5s^2 5p^6$
6	55 Cs $6s^1$	56 Ba $6s^2$	57 La $6s^2 5d^1$	72 Hf $6s^2 5d^2$	73 Ta $6s^2 5d^3$	74 W $6s^2 5d^4$	75 Re $6s^2 5d^5$	76 Os $6s^2 5d^6$	77 Ir $6s^2 5d^7$	78 Pt $6s^1 5d^9$	79 Au $6s^1 5d^{10}$	80 Hg $6s^2 5d^{10}$	81 Tl $6s^2 6p^1$	82 Pb $6s^2 6p^2$	83 Bi $6s^2 6p^3$	84 Po $6s^2 6p^4$	85 At $6s^2 6p^5$	86 Rn $6s^2 6p^6$
7	87 Fr $7s^1$	88 Ra $7s^2$	89 Ac $7s^2 6d^1$	104 Rf $7s^2 6d^2$	105 Db $7s^2 6d^3$	106 Sg $7s^2 6d^4$	107 Bh $7s^2 6d^5$	108 Hs $7s^2 6d^6$	109 Mt $7s^2 6d^7$	110 Ds $7s^2 6d^8$	111 Rg $7s^2 6d^9$	112 Cn $7s^2 6d^{10}$	(113)	114 Fl $7s^2 7p^2$	(115)	116 Lv $7s^2 7p^4$	(117)	(118)
				58 Ce $6s^2 4f^1 5d^1$	59 Pr $6s^2 4f^3$	60 Nd $6s^2 4f^4$	61 Pm $6s^2 4f^5$	62 Sm $6s^2 4f^6$	63 Eu $6s^2 4f^7$	64 Gd $6s^2 4f^7 5d^1$	65 Tb $6s^2 4f^9$	66 Dy $6s^2 4f^{10}$	67 Ho $6s^2 4f^{11}$	68 Er $6s^2 4f^{12}$	69 Tm $6s^2 4f^{13}$	70 Yb $6s^2 4f^{14}$	71 Lu $6s^2 4f^{14} 5d^1$	
				90 Th $7s^2 6d^2$	91 Pa $7s^2 5f^2 6d^2$	92 U $7s^2 5f^3 6d^1$	93 Np $7s^2 5f^4 6d^1$	94 Pu $7s^2 5f^6$	95 Am $7s^2 5f^7$	96 Cm $7s^2 5f^7 6d^1$	97 Bk $7s^2 5f^9$	98 Cf $7s^2 5f^{10}$	99 Es $7s^2 5f^{11}$	100 Fm $7s^2 5f^{12}$	101 Md $7s^2 5f^{13}$	102 No $7s^2 5f^{14}$	103 Lr $7s^2 5f^{14} 6d^1$	

Figure 8.1

The ground-state electron configurations of the elements. For simplicity, only the configurations of the outer electrons are shown.

Special names

- For convenience, some element groups have special names.
- The Group 1A elements are called *alkali metals*.
- The Group 2A elements are called *alkaline earth metals*.
- The Group 7A are known as *halogens*.
- The Group 8A are called *noble gases* (or *inert* or *rare gases*).
- **Lanthanides**, are the *rare earth series or metals*.

Determination of the period number and group number from electron configuration

A- For representative elements (groups A)

- **Period number:** it equals to the number of the highest energy level.
- **Group number:** it equals the sum of electrons in the highest energy level (valance electrons).

B- For transition metals

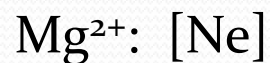
- **Period number:** it equals to the number of the highest energy level.
- **Group number:**
 - If number of electrons of $d < 6$, Group number = number of electrons + 2
 - If number of electrons of $d = 6, 7, 8$, Group number = 8
 - If number of electrons of $d = 9, 10$, Group number = 1, 2 respectively.

Electron Configurations of Cations and Anion

- Because many ionic compounds are made up of monatomic anions and/or cations, it is helpful to know how to write the electron configurations of these ionic species.

1. Ions Derived from Representative Elements

In the formation of a cation from the neutral atom of a representative element, one or more electrons are removed from the highest occupied n shell. Here are the electron configurations of some neutral atoms and their corresponding cations:



Note that each ion has a stable noble gas configuration.

- In the formation of an anion, one or more electrons are added to the highest partially filled n shell. Consider these examples:



- Again, all the anions have stable noble gas configurations. Thus, a characteristic of most representative elements is that ions derived from their neutral atoms have the noble gas outer electron configuration $ns^2 np^6$.

Isoelectronicity

- Ions, or atoms and ions, that have the same number of electrons, and hence the same ground-state electron configuration, are said to be **isoelectronic**.
- Thus, H^- and He are isoelectronic.
- Al^{+3} , Mg^{+2} , Na^+ , N^{-3} , O^{-2} , F^{-1} and Ne are isoelectronic; and so on.

Cations Derived from Transition Metals

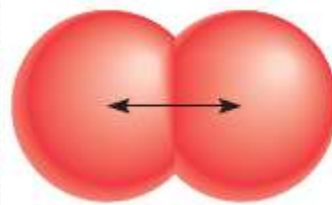
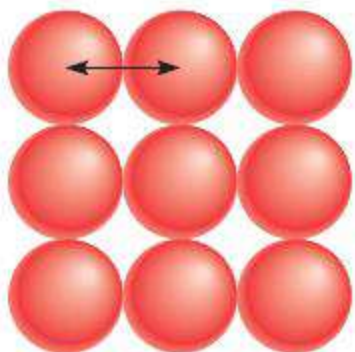
- When a cation is formed from an atom of a transition metal, electrons are always removed first from the ns orbital and then from the $(n-1)d$ orbitals (although the $4s$ orbital is always filled before the $3d$ orbital).
- This is because the $3d$ orbital is more stable than the $4s$ orbital in transition metal *ions*. (*not as the atoms*)
- Keep in mind that most transition metals can form more than one cation and that frequently the cations are not isoelectronic with the preceding noble gases.

Periodic Variation in Physical Properties

- As we have seen, the electron configurations of the elements show a periodic variation with increasing atomic number. Consequently, there are also periodic variations in physical and chemical behavior.
- With a group, there is similarities in physical and chemical characters as there outer electron configuration is similar.
- Physical characters include: atomic sizes, ionization energy, electron affinity and metallic characters.

Atomic size

- A number of physical properties, including density, melting point, and boiling point, are related to the sizes of atoms.
- The size of an atom can be defined in terms of its **atomic radius**, which is *one-half the distance between the two nuclei in two adjacent metal atoms (for mono atoms) or as one-half the distance between the centers of the atoms in the molecule (di atoms).*



Increasing atomic radius

Increasing atomic radius

1A	2A	3A	4A	5A	6A	7A	8A
H 1							He 2
Li 3	Be 4	B 5	C 6	N 7	O 8	F 9	Ne 10
Na 11	Mg 12	Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
K 19	Ca 20	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
Rb 37	Sr 38	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
Cs 55	Ba 56	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86

Atomic size

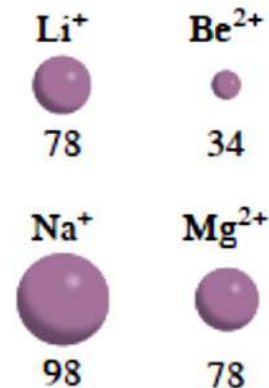
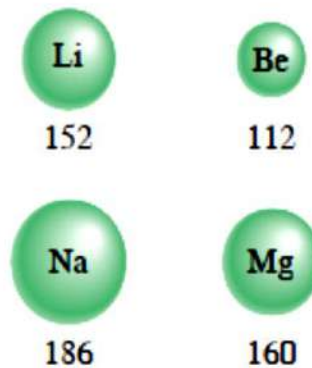
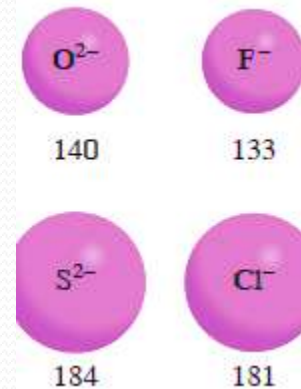
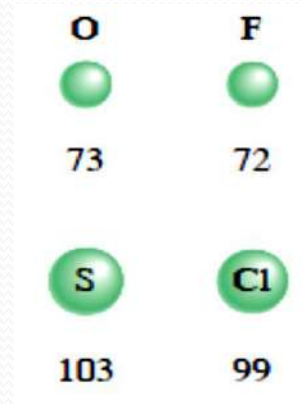
- Periodic **trends** are clearly evident. The atomic radius is determined to a large extent by the strength of the attraction between the nucleus and the *outer-shell electrons*.
- In the **same period**, as **nuclear charge increases**, the stronger the hold of the nucleus on outer-shell electrons, and the smaller the atomic radius. Hence, the atomic radius decreases steadily from left to right.
- Within a **group** of elements we find that atomic radius increases with increasing atomic number. This is because that orbital size increases with the increasing the **number of principal levels**, so does the average distance of a outer shells from the nucleus.

Ionic Radius

- Ionic radius is the radius of a cation or an anion. Ionic radius affects the physical and chemical properties of an ionic compound. 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.
- 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*.

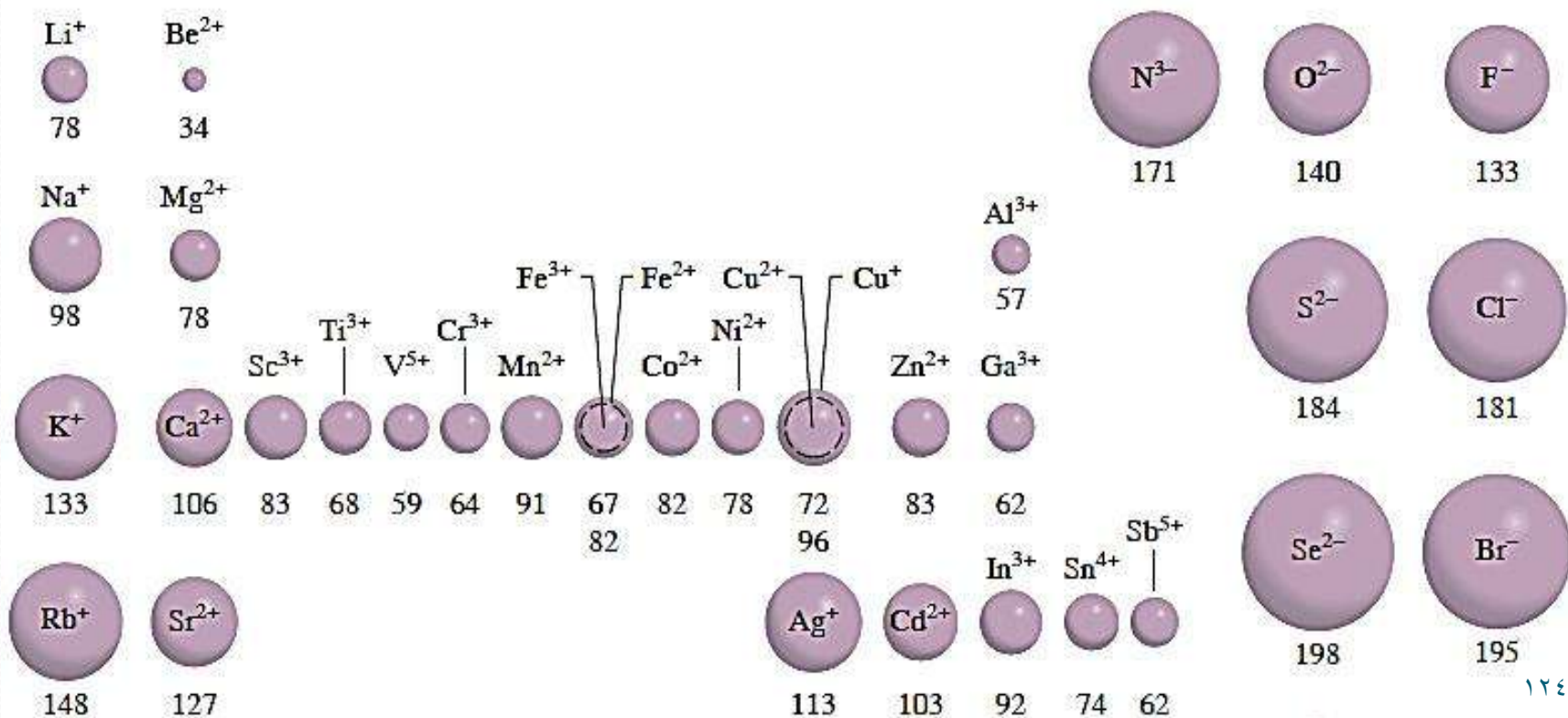
Ionic Radius

- Compare the ionic radius with atomic radius for each element.



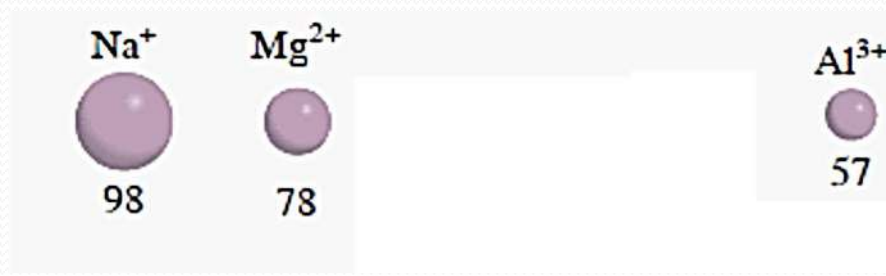
Ionic Radius

- From top to bottom, both the atomic radius and the ionic radius increase within a group.
- In the same period, the size of anions is larger than the size of cations.

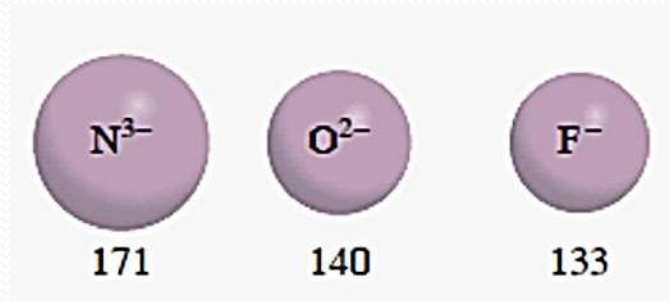


Ionic Radius

- Focusing on isoelectronic cations, we see that the radii of tripositive ions are smaller than those of dipositive which in turn are smaller than unipositive ions.

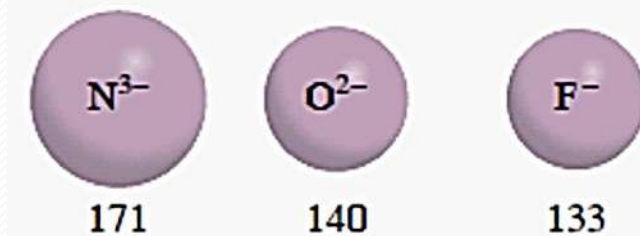


- The radius of isoelectronic anions increases as we go from ions with uninegative charge to those with dinegative charge and so on.



Ionic Radius

- Comparing isoelectronic ions, cations are smaller than anions.



Ionization Energy

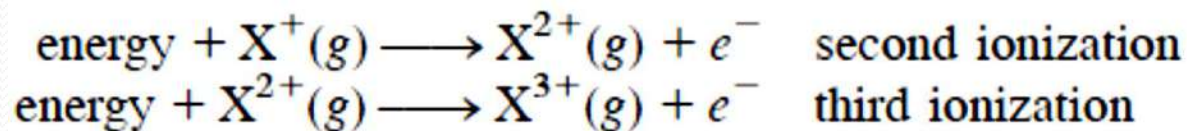
- Ionization energy is the minimum energy (in kJ/mol) required to remove an electron from a gaseous atom in its ground state.
- The magnitude of ionization energy is a measure of how “tightly” the electron is held in the atom.
- The higher the ionization energy, the more difficult it is to remove the electron.
- For a many-electron atom, the amount of energy required to remove the first electron from the atom in its ground state, is called the first ionization energy (I_1) described by the following equation:



X represents an atom of any element and e^- is an electron.

Ionization Energy

- The second ionization energy (I_2) and the third ionization energy (I_3) are shown in the following equations:



- When an electron is removed from an atom, the repulsion among the remaining electrons decreases and the effective nuclear charge increases, so more energy is needed to remove another electron from the positively charged ion.
- Thus, ionization energies always increase in the following order:

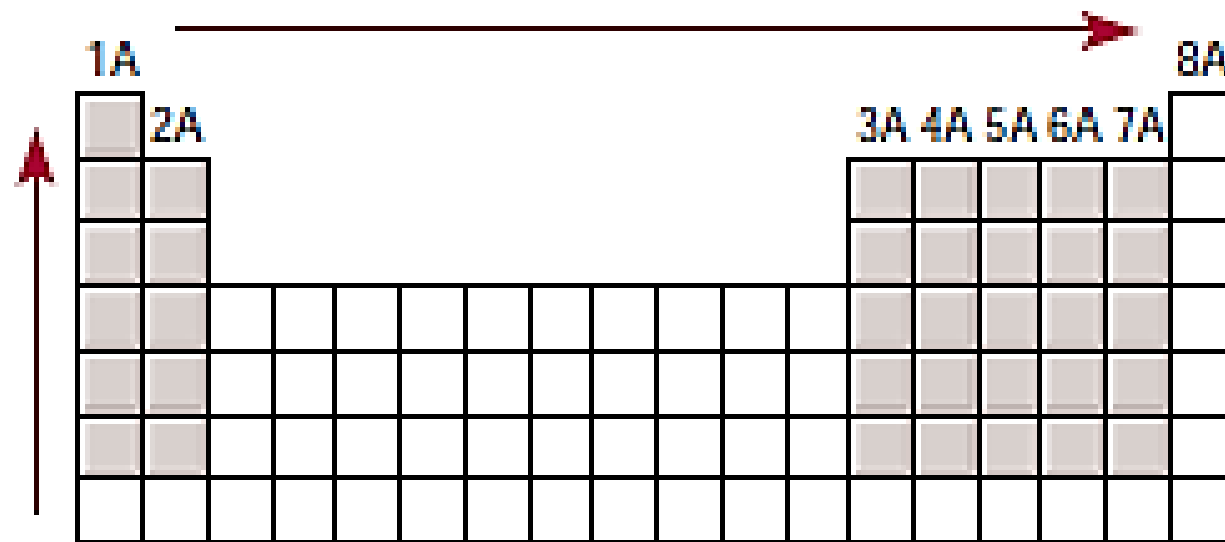
$$I_1 < I_2 < I_3 < \dots$$

Ionization Energy

- Ionization is always an endothermic process, that means the energy absorbed by atoms (or ions) in the ionization process has a positive value.
- Note that, apart from small irregularities, **the first ionization energies of elements in a period increase with increasing atomic number.**
- This trend is due to the increase in effective nuclear charge from left to right (as in the case of atomic radii variation). A larger effective nuclear charge means a more tightly held outer electron, and hence a higher first ionization energy.

Ionization Energy

- For a given group, ionization energy decreases with increasing atomic number (that is, as we move down the group).
- Elements in the same group have similar outer electron configurations. However, as the principal quantum number n increases, so does the average distance of a valence electron from the nucleus. A greater separation between the valence electrons and the nucleus means a weaker attraction (+ shielding effect of the inner subshells), so that it becomes increasingly easier to remove the first electron as we go from element to element down a group.



The increase in first ionization energy from left to right across a period and from bottom to top in a group for representative elements.

Electron Affinity

- It is a measure to the ability of atoms to accept one or more electrons.
- There is an increase in the tendency to accept electrons (electron affinity values increase) from left to right across a period.
- The values vary within a given group.
- Halogens (Group 7A) have the highest electron affinity values. This is not surprising when we realize that by accepting an electron, each halogen atom assumes the stable electron configuration of the noble gas immediately to its right.
- Group 1A and 2A have low electron affinity.

Metallic character

- Elements are further classified into metals, non-metals, and metalloids based on their properties, which are correlated with their placement in the periodic table.
- Metallic character of the elements decreases from left to right across a period and increases from top to bottom within a group.
- Group 7A are the most reactive non-metals and group 1A are the most reactive metals.
- Metalloids have properties intermediate between the metals and nonmetals. Metalloids are useful in the semiconductor industry. The physical properties of metalloids tend to be metallic, but their chemical properties tend to be non-metallic.

1 1A											13 3A	14 4A	15 5A	16 6A	17 7A	18 8A	
1 H	2 2A											5 B	6 C	7 N	8 O	9 F	10 Ne
3 Li	4 Be											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
11 Na	12 Mg	3 3B	4 4B	5 5B	6 6B	7 7B	8	9	10	11 1B	12 2B	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	(113)	114	(115)	116	(117)	(118)
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	(113)	114	(115)	116	(117)	(118)

Metals	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
Metalloids	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
Nonmetals														

Metallic characters

Characteristic physical properties of metallic and non-metallic elements

Metallic Elements	Nonmetallic elements
Distinguishing luster (shine)	Non-lustrous, various colors
Metals are solids at room temperature except mercury	Gases (oxygen) and solids (carbon) except bromine
Malleable and ductile (flexible), hard as solids	Brittle, hard or soft
Conduct heat and electricity	Poor conductors
Metals have high melting and boiling points	The melting points of non-metals are generally lower than metals

Metallic characters

Characteristic chemical properties of metallic and non-metallic elements

Metals typically have 1 to 3 electrons in the outermost shell of their atoms	Non-Metals typically have more than 3 electrons in the outermost shell of their atoms
Metals tend to have low ionization energies and low electron affinity and <i>typically lose electrons (i.e. are oxidized)</i> . when reacting with non-metals (tend to become cation) and form ionic bonds	Non-metals tend to have high ionization energies and high electron affinity and <i>typically gain or share electrons with other atoms</i> . They are electronegative in character. Nonmetals, when reacting with metals, tend to gain electrons and become anions and form ionic bonds Nonmetals, when reacting with other non-metals, share electrons and form covalent bonds.
Most metal oxides are basic and ionic. They dissolve in water to form metal hydroxides : $\text{Na}_2\text{O}(s) + \text{H}_2\text{O}(l) \rightarrow 2\text{NaOH}(aq)$ $\text{CaO}(s) + \text{H}_2\text{O}(l) \rightarrow \text{Ca}(\text{OH})_2(aq)$	They generally form covalent, acidic or neutral oxides with oxygen that that dissolve in water to form acids: $\text{P}_4\text{O}_{10}(s) + 6\text{H}_2\text{O}(l) \longrightarrow 4\text{H}_3\text{PO}_4(aq)$ $\text{SO}_3(g) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_2\text{SO}_4(aq)$
Metal oxides exhibit their basic chemical nature by reacting with acids to form metal salts and water: $\text{MgO}(s) + \text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2\text{O}(l)$ $\text{NiO}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{NiSO}_4(aq) + \text{H}_2\text{O}(l)$	Nonmetallic oxides are acidic can combine with bases to form salts. $\text{CO}_2(g) + 2\text{NaOH}(aq) \rightarrow \text{Na}_2\text{CO}_3(aq) + \text{H}_2\text{O}(l)$
Form cations in aqueous solution	Form anions in aqueous solution

Summary

