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

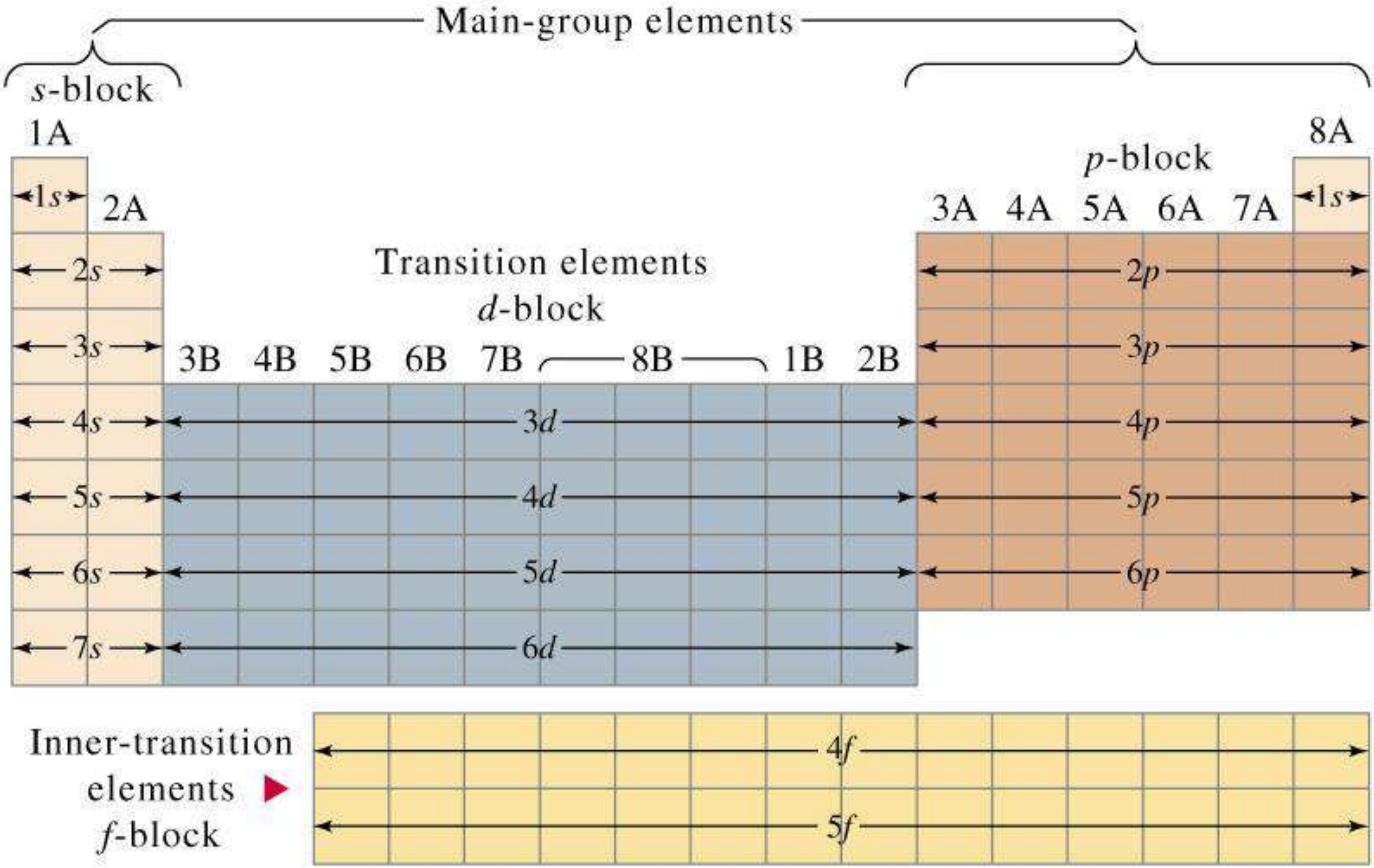
- A **block** of the periodic table is a set of chemical elements having their differentiating electrons predominately in the same type of atomic orbital.
- **A differentiating electron** is the electron that differentiates an element from the previous one.

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.

1. **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.
2. **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).
3. **The Group 2B elements are Zn, Cd, and Hg**, which are neither representative elements nor transition metals.
4. **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
s-block
1A
2A
p-block
8A
3A 4A 5A 6A 7A
1s
2s
3s
4s
5s
6s
7s
Transition elements
d-block
3B 4B 5B 6B 7B 8B 1B 2B
3d
4d
5d
6d
2p
3p
4p
5p
6p
Inner-transition elements
f-block
4f
5f

Representative elements

Noble gases

Transition metals

Zinc
Cadmium
Mercury

Lanthanides

Actinides

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

58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$
90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$	103 **Lr** $7s^25f^{14}6d^1$

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 7Aare known as ***halogens.***
- The Group 8A are called ***noble gases*** (or ***inert*** or ***rare gases***).
- **Lanthanides,** are the ***rare earth series or metals.***

- ***To be continued***

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.

Problems

- Discuss the electron configurations of the following atoms :P^{15}, Mn^{25}, I^{53}, W^{74}, Sm^{62}, Pb^{82}

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:

Na: $[Ne]3s^1$	Na^+: $[Ne]$
Mg: $[Ne]\,3s^2$	Mg^{2+}: $[Ne]$
Al: $[Ne]3s^23p^1$	Al^{3+}: $[Ne]$

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:

H: $1s^1$	H^-: $1s^2$ or [He]
F: $1s^2 2s^2 2p^5$	F^-: $1s^2 2s^2 2p^6$ or [Ne]
O: $1s^2 2s^2 2p^4$	O^{2-}: $1s^2 2s^2 2p^6$ or [Ne]
N: $1s^2 2s^2 2p^3$	N^{3-}: $1s^2 2s^2 2p^6$ or [Ne]

- 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^2np^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

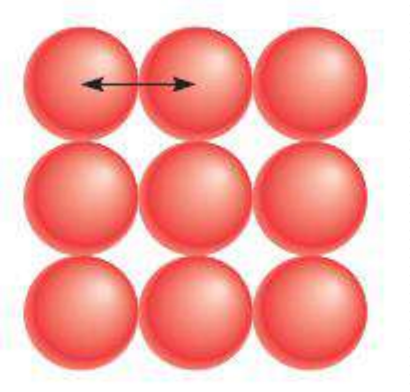
- When a cation is formed from an atom of a transition metal, electrons are always removed first from the *ns* orbital and then from the (*n*-1)*d* orbitals (although the 4*s* orbital is always filled before the 3*d* orbital).
- This is because the 3d orbital is more stable than the 4s orbital in transition 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.

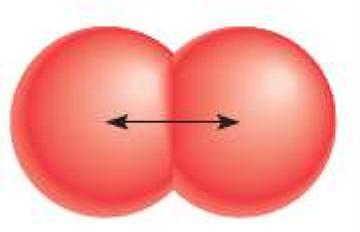
Periodic Variation in Physical Properties

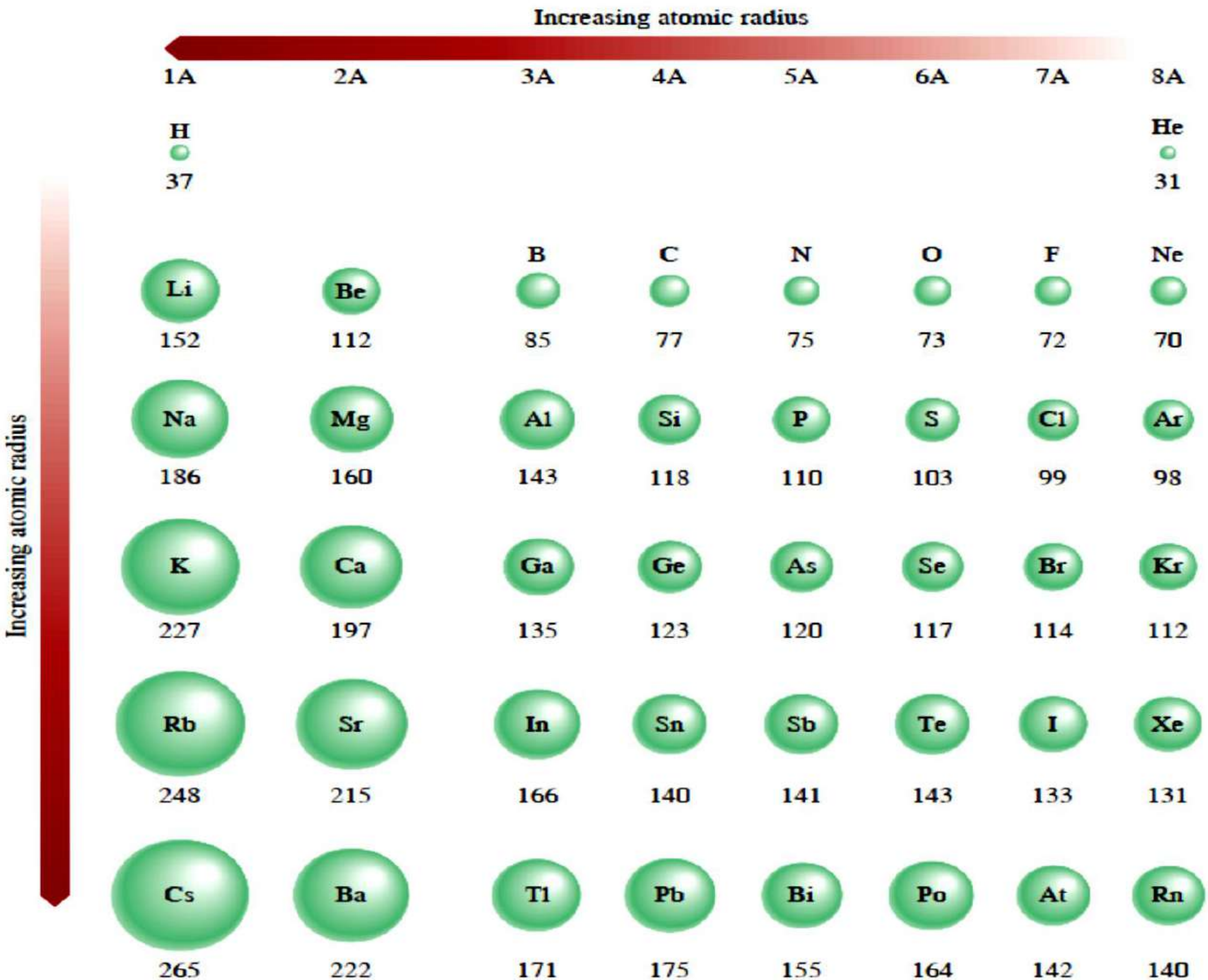
- As we have seen, the electron configurations of the elements show a periodic variation with increasing atomic number. Consequently, there are also periodic variations in physical and chemical behavior.
- With a group, there is similarities in physical and chemical characters as there outer electron 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
1A
2A
3A
4A
5A
6A
7A
8A
H
37
He
31
Li
152
Be
112
B
85
C
77
N
75
O
73
F
72
Ne
70
Na
186
Mg
160
Al
143
Si
118
P
110
S
103
Cl
99
Ar
98
K
227
Ca
197
Ga
135
Ge
123
As
120
Se
117
Br
114
Kr
112
Rb
248
Sr
215
In
166
Sn
140
Sb
141
Te
143
I
133
Xe
131
Cs
265
Ba
222
Tl
171
Pb
175
Bi
155
Po
164
At
142
Rn
140
Increasing atomic radius

Atomic size

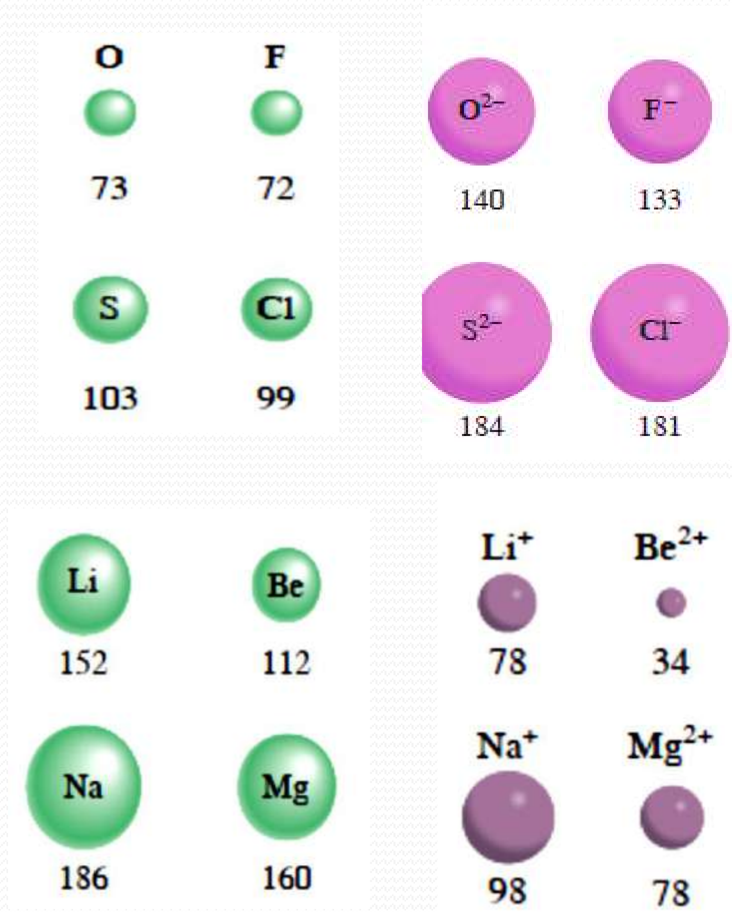
- Periodic **trends** are clearly evident. The atomic radius is determined to a large extent by the strength of the attraction between the nucleus and the ***outer-shell electrons.***
- In the **same period**, **as nuclear charge increases**, the stronger the hold of the nucleus on outer-shell electrons, 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.**

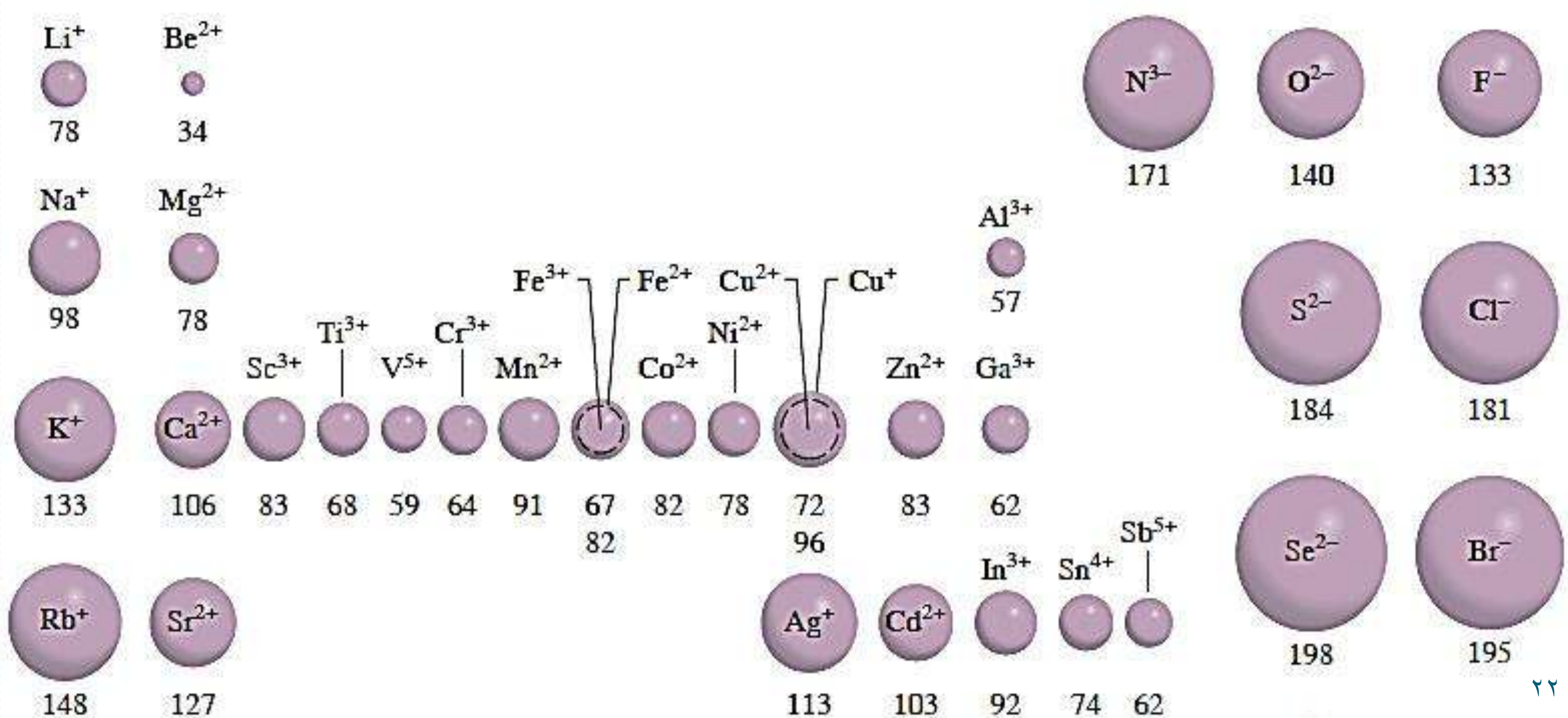
Ionic Radius

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



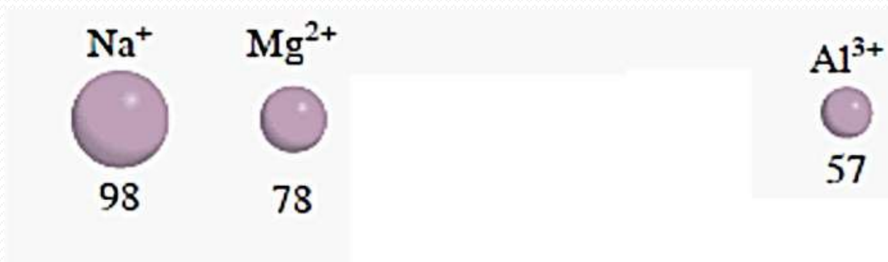
Ionic Radius

- 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.

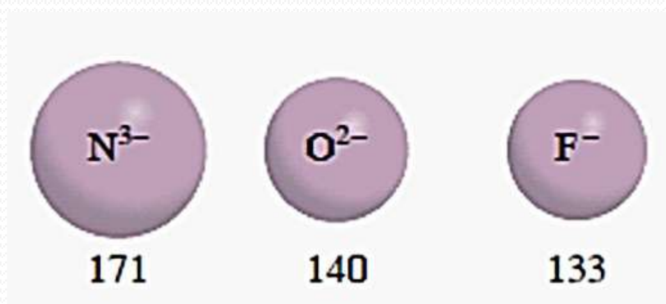


Ionic Radius

- 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.

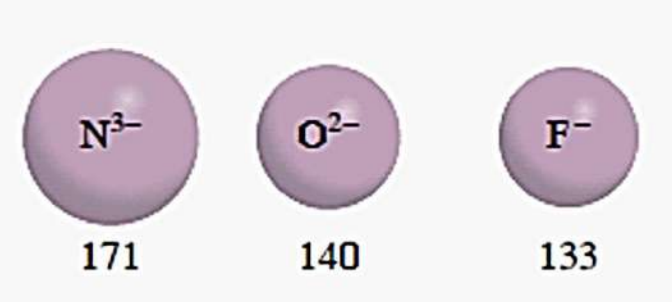


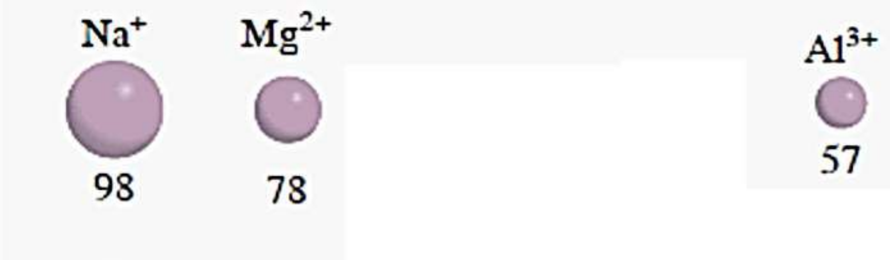
- The radius of isoelectronic anions increases as we go from ions with uninegative charge to those with dinegative charge and so on.



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:

$$\text{energy} + X(g) \longrightarrow X^{+}(g) + e^{-}$$

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

Ionization Energy

- The second ionization energy (I2) and the third ionization energy (I3) are shown in the following equations:

$$\text{energy} + X^{+}(g) \longrightarrow X^{2+}(g) + e^{-} \quad \text{second ionization}$$
$$\text{energy} + X^{2+}(g) \longrightarrow X^{3+}(g) + e^{-} \quad \text{third ionization}$$

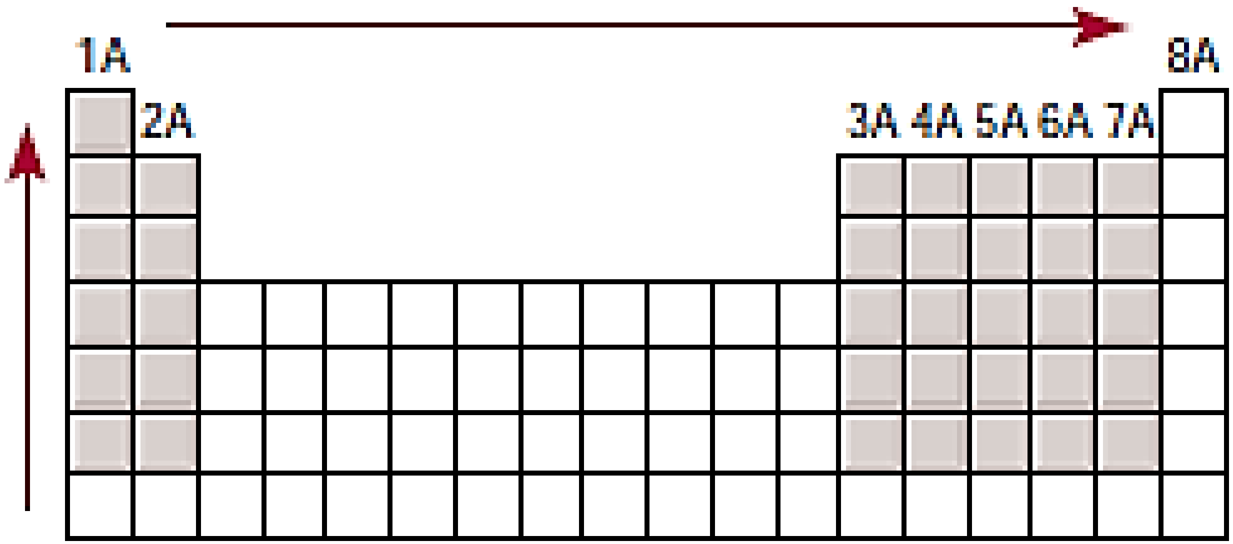
- When an electron is removed from an atom, the repulsion among the remaining electrons decreases and the effective nuclear charge increases, so more energy is needed to remove another electron from the positively charged ion.
- Thus, ionization energies always increase in the following order: $I_1 < I_2 < I_3 < \cdot$

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 valance electrons and the nucleus means a weaker attraction (+ shielding effect of the inner sub-shells), 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.
- It is defined as the energy change that occurs when an electron is accepted by an atom in the gaseous state to form an anion.

$$X(g) + e^{-} \longrightarrow X^{-}(g)$$

- If the energy is released (exothermic reaction), this means the anoin is stable and accepting the electron is favorable.
- If the energy is absorbed (endothermic), this means the accepting the electron is unfavorable.

Electron Affinity

- 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	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											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)

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

Metals

Metalloids

Nonmetals

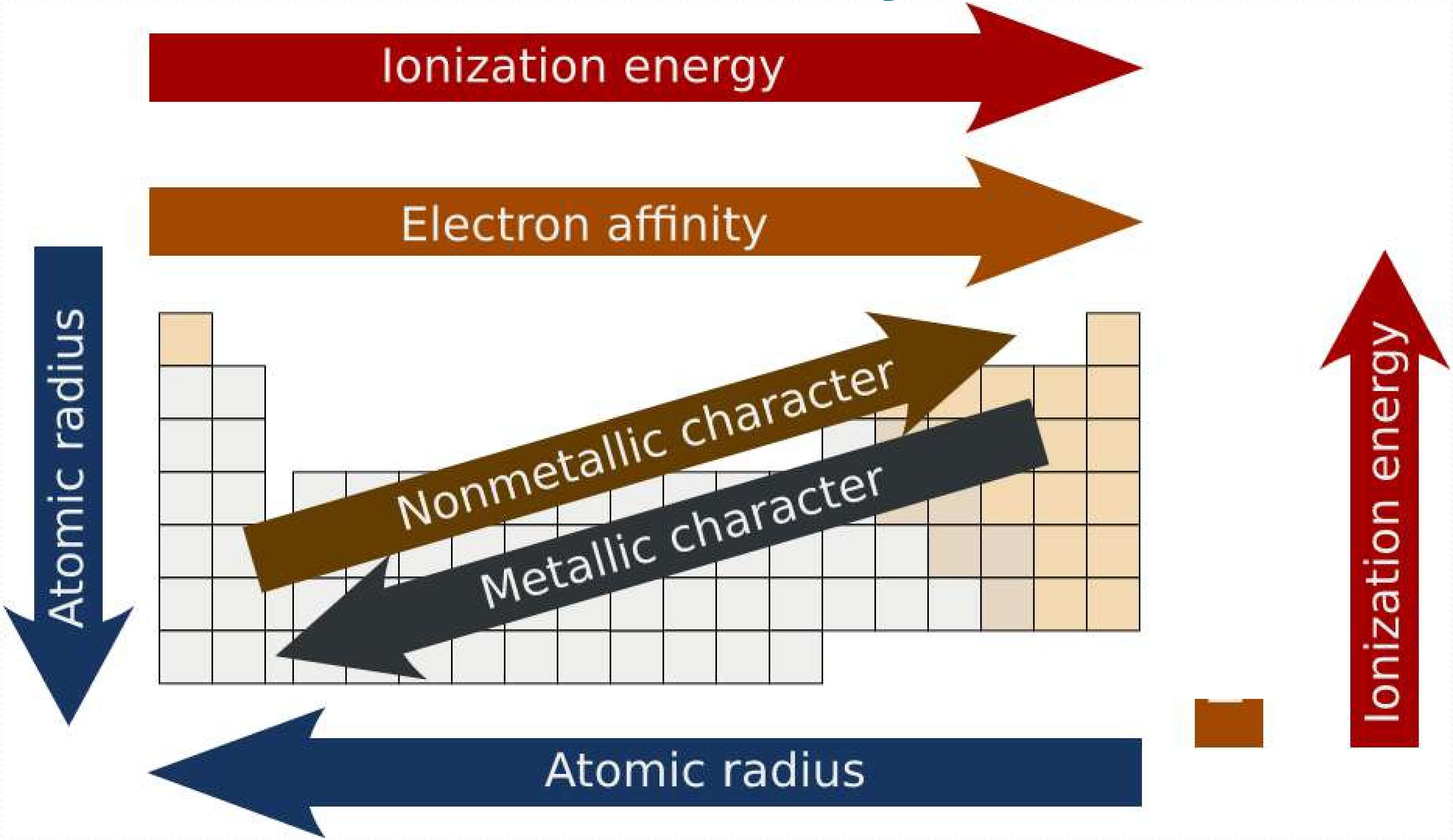
Metallic characters

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 *typically lose electrons (i.e. are **oxidized**)*. 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 *typically* gain or share electrons with other atoms. 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***: $Na_2O(s)+H_2O(l) \rightarrow 2NaOH(aq)$ $CaO(s)+H_2O(l) \rightarrow Ca(OH)_2(aq)$	They generally form covalent, acidic or neutral oxides with oxygen that that dissolve in water to form acids: $P_4O_{10}(s) + 6H_2O(l) \longrightarrow 4H_3PO_4(aq)$ $SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$
Metal oxides exhibit their ***basic*** chemical nature by reacting with ***acids*** to form metal ***salts*** and water: $MgO(s)+HCl(aq) \rightarrow MgCl_2(aq)+H_2O(l)$ $NiO(s)+H_2SO_4(aq) \rightarrow NiSO_4(aq)+H_2O(l)$	Nonmetallic oxides are acidic can combine with bases to form salts. $CO_2(g)+2NaOH(aq) \rightarrow Na_2CO_3(aq)+H_2O(l)$
Form cations in aqueous solution	Form anions in aqueous solution

Summary


Ionization energy
Electron affinity
Atomic radius
Nonmetallic character
Metallic character
Ionization energy
Atomic radius