Chapter 5
The Periodic Table of Elements

1-The History of the Periodic Table

The history of the periodic table began in the 19th century when there were approximately 50 discovered elements. Several attempts were made to systematically organize them in a way that would highlight their similarities and differences. Some of the most significant of these attempts were:

Mendeleev's Periodic Table (1869) Dmitri Mendeleev arranged the elements in ascending order of their atomic weights in the form of horizontal rows and vertical groups. Elements with similar properties were placed together in vertical groups, and he left spaces for elements that were assumed to exist but had not yet been discovered. One of its drawbacks was that it treated isotopes of the same element as separate elements because of their differing atomic weights, and he sometimes placed multiple elements in one slot due to their significant similarities in properties.

Reihen	Gruppo I. — R^2O	Gruppo II. — RO	Gruppo III. — R^2O^3	Gruppo IV. RH^4 RO^2	Gruppo V. RH^3 R^2O^5	Gruppo VI. RH^2 RO^3	Gruppo VII. RH R^2O^7	Gruppo VIII. — RO^4
1	H=1							
2	Li=7	Bo=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fo=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	So=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	To=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

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Mosley's Periodic Table (1913): Henry Moseley arranged the elements in ascending order based on their atomic numbers and added the zero group to the table, which includes the inert gases.

Group 0	I a	I b	II a	II b	III a	III b	IV a	IV b	V a	V b	VI a	VI b	VII a	VII b	VIII
	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	Cu 29	Ca 20	Zn 30	Sc 21	Ga 31	Ti 22	Ge 32	V 23	As 33	Cr 24	Se 34	Mn 25	Br 35	Fe 26, Co 27, Ni 28
Kr 36	Rb 37	Ag 47	Sr 38	Cd 48	Y 39	In 49	Zr 40	Sn 50	Nb 41	Sb 51	Mo 42	Te 52	–	I 53	Ru 44, Rh 45, Pd 46
Xe 54	Cs 55	Au 79	Ba 56	Hg 80	57-71*	Tl 81	Hf 72	Pb 82	Ta 73	Bi 83	W 74	Po 84	Re 75	–	Os 76, Ir 77, Pt 78
Rn 86	–		Ra 88		Ac 89		Th 90		Pa 91		U 92				

The Modern Periodic Table (Seaborg's Table, 1945): Chemical elements in the modern periodic table are arranged in ascending order of their atomic numbers and follow the filling of electron energy levels. The modern periodic table consists of 7 rows and 18 columns and includes 118 elements, with 92 of them being naturally occurring and the rest being synthetic.

2- Study of Periods (Rows)

These are the horizontal rows in the periodic table, represented by an Arabic number from 1 to 7. Each period corresponds to a specific value of the principal quantum number, n. Each period starts with an element that has one electron in the outermost valence shell and ends with an element that has a full valence shell, known as the inert gas (noble gas). There is a simple relationship that allows us to determine the number of elements in each period based on n.

n=1,3,5,7 $x = \frac{(n+1)^2}{2}$: odd n

n=2,4,6 $x = \frac{(n+2)^2}{2}$: even n

We distinguish the lanthanide series, represented by the element lanthanum (La), and the actinide series, represented by the element actinium (Ac), as they share similar chemical properties. They are placed separately below the periodic table to avoid making it too wide.

3- Columns

These are the vertical rows in the periodic table, consisting of 18 columns, referred to as groups. Elements within each group have the same valence electron configuration, which imparts similar chemical properties to these elements. As you move from one column to another, the number of electrons in the outermost energy level increases. There are three systems for numbering the groups: the first uses Arabic numerals, the second uses Roman numerals, and the third combines Roman numerals and Latin letters. The Arabic numeral system has been adopted by the International Union of Pure and Applied Chemistry (IUPAC) to replace Roman numerals and Latin letters.

The periodic table is divided into two types of groups:

- **Groups A:** This category includes 8 groups where the group number corresponds to the number of electrons in the valence shell. Its elements are called representative elements because their outermost electron shells represent them in all chemical reactions.

18	17	16	15	14	13	2	1	The column with the new numbering
0 أو $VIII_A$	VII_A	VI_A	V_A	IV_A	III_A	II_A	I_A	The group or category (old numbering)
ns^2np^6	ns^2np^5	ns^2np^4	ns^2np^3	ns^2np^2	ns^2np^1	ns^2	ns^1	The valence shell
8	7	6	5	4	3	2	1	Valence electron count
Noble gases	Halogens							The family

- Halogens have a high tendency to easily acquire electrons, leading to their electron configuration resembling that of noble gases.
- Noble gases, also known as inert gases, are stable elements. They are named as such due to their chemical inactivity, attributed to the saturation of their outermost energy level with electrons.

- **Groups B**: The outermost electron configuration for elements in Group B is ns(n-1)d.

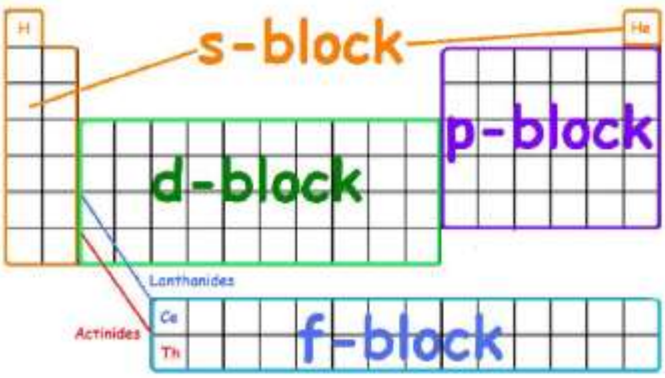
12	11	10	9	8	7	6	5	4	3	The column with the new numbering
II_B	I_B	$VIII_B$			VII_B	VI_B	V_B	IV_B	III_B	The group (old (numbering
$ns^2(n-1)d^{10}$	$ns^1(n-1)d^9$	$ns^2(n-1)d^8$	$ns^2(n-1)d^7$	$ns^2(n-1)d^6$	$ns^2(n-1)d^5$	$ns^2(n-1)d^4$	$ns^2(n-1)d^3$	$ns^2(n-1)d^2$	$ns^2(n-1)d^1$	The valence shell
2	1	10	9	8	7	6	5	4	3	Valence electron count
المجموعات العناصر الانتقالية										The family

The elements of Group VIIIB are called "triads" due to their similarity in properties to elements in the same rows and columns. Additionally, transition elements are classified into two categories.

- Main Transition Elements: These are elements whose electron configuration ends in the sublevel d.
- Inner Transition Elements: These are elements whose electron configuration ends in the sublevel f (the actinides and lanthanides). We fill the inner f sublevel before the outer d sublevel, which has already started filling.

4- Periodic Table Regions

The periodic table is divided into 4 regions: wings, sectors, blocks, and main regions.



- S-Block: Includes elements whose outermost electrons are in the S sublevel in their electron configuration.
- P-Block: Includes elements whose outermost electrons are in the P sublevel in their electron configuration.
- d-Block: Includes elements whose outermost electrons are in the d sublevel in their electron configuration.
- f-Block: Includes elements whose outermost electrons are in the f sublevel in their electron configuration.

5- General Rules for Determining the Position of an Element in the Periodic Table

Determining the Period or Row Number It is the number of the last principal energy level in the electron configuration (the highest principal quantum number, n, in the electron configuration). Example:

Fe_{26} : $[Ar]_{18}$ $4S^2$ $3d^6$ Belongs to Period 4

The block or region is named after the sublevel that receives the last electron in the electron configuration.

For example: Fe_{26} : $[Ar]_{18}$ $4S^2$ $3d^6$ Belongs to d-block

The groups

- Elements x belong to Groups A if the sublevel that receives the last electron in the electron configuration is either S or P.

Example:

- Element x belongs to Groups B if the sublevel that receives the last electron in the electron configuration is either d or f.

Example: Fe_{26} : $[Ar]_{18}$ $4S^2$ $3d^6$ belongs to the groups B.

The group

The group number for an element is the number of valence electrons, and it is typically represented using Roman numerals.

The column

The column number is determined by the number of valence electrons. If the outermost electron belongs to the P sublevel, we add 10 to the number of valence electrons to determine the column number. For the d-block, the column number is the sum of electrons in the sublevels ns(n-1)d.

Example :

$_{30}Zn$: $[Ar]_{18}$ $4S^2$ $3d^{10}$

Belongs to the column 12

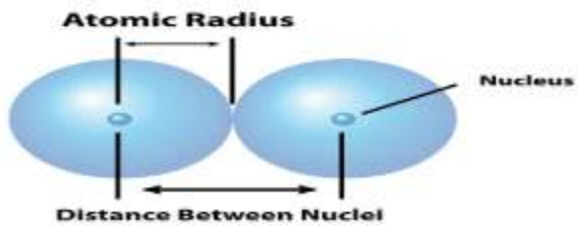
6- The gradient of properties in the periodic table

There are physical and chemical properties that vary (change in a specific direction) within the horizontal periods and vertical groups in the periodic table.

The periodic trend of atomic radius.

The wave theory demonstrated that the exact location of an electron around the nucleus cannot be precisely determined. Therefore, it would be incorrect to define the atomic radius as the distance from the nucleus to the outermost electron. Instead, the atomic radius is defined as:

One-half the distance between the nuclei of identical atoms that are bonded together.

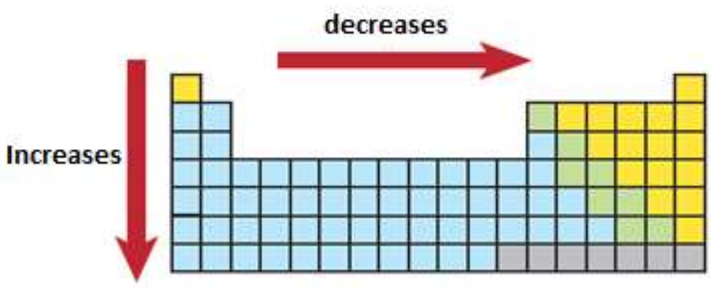


In the horizontal periods (within the same row)

The atomic radius decreases as the atomic number increases across a single period from left to right. The explanation for this: is that with an increase in the atomic number, the positive charges (number of protons) in the nucleus gradually increase while the principal energy levels within the period remain constant. As a result, the attractive force between the nucleus and the valence electrons increases, leading to a decrease in the atomic radius.

In the column

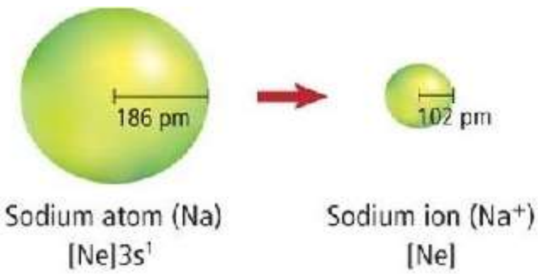
the atomic radius increases as the atomic number increases when moving from top to bottom within a group. The explanation for this is: that as the atomic number increases, the positive charge in the nucleus also increases, which would tend to pull the electrons closer and reduce the atomic radius. However, simultaneously, the number of filled energy levels with electrons also increases. These filled energy levels act as electron shielding, reducing the effective nuclear charge experienced by the outermost electrons. As a result, the nucleus's attraction for these outermost electrons decreases, causing the atomic radius to increase.



The periodic trend of atomic radius

Periodic radius of the positive ion

The positive ion's radius is smaller than half the radius of its neutral atom. The justification lies in the fact that for an atom to become a positively charged ion, it must lose electrons to achieve stability. Consequently, the number of electrons in the electron shells becomes less than the number of protons in the nucleus, leading to an increased attraction force between the nucleus and the remaining electrons.

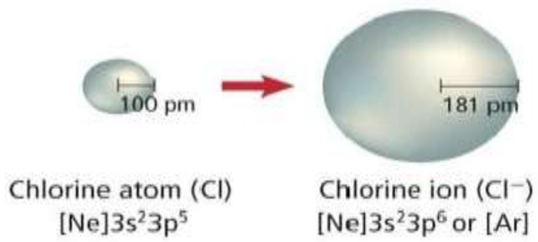


Periodic radius of the positive ion

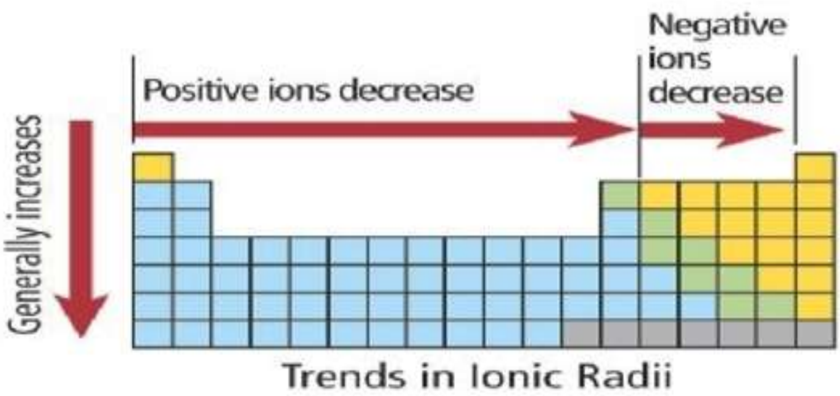
The positive ion's radius is smaller than half the radius of its neutral atom. The justification lies in the fact that for an atom to become a positively charged ion, it must lose electrons to achieve stability. Consequently, the number of electrons in the electron shells becomes less than the number of protons in the nucleus, leading to an increased attraction force between the nucleus and the remaining electrons.

Periodic radius of negative ion

The radius of the negative ion is larger than half the radius of its neutral atom. Justification: This is because for an atom to transform into a negative ion, it must gain electrons, increasing the number of negative charges in the electron shells beyond the number of positive charges in the nucleus. As a result, the attractive force from the nucleus is distributed over a larger number of electrons, reducing the overall nucleus-electron attraction.



The figure below summarizes the trend in the change of ion radii across groups and periods.



Ionization Energy Periodicity

Ionization energy is the amount of energy required to remove an electron from a gaseous atom in its ground state. The energy needed to remove the first electron from an atom is called the first ionization energy (E_{i1} or ΔE_1).

$$M_{(g)} + \Delta E_1 \rightarrow M^{+}_{(g)} + e^{-}$$

The energy required to remove the second electron from an atom is called the second ionization energy (E_{i2} or ΔE_2).

$$M^{+}_{(g)} + \Delta E_2 \rightarrow M^{+2}_{(g)} + e^{-}$$

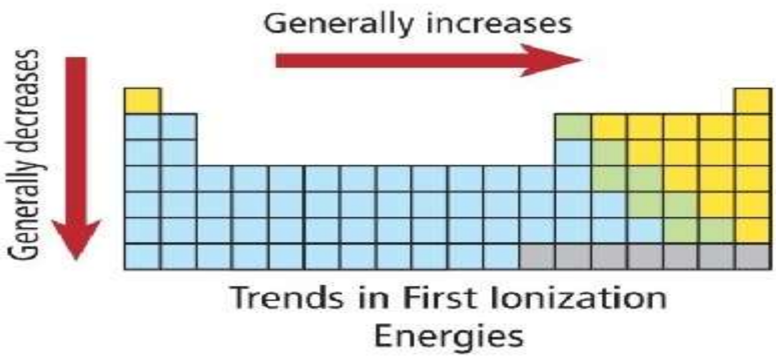
In the horizontal periods

Ionization energy increases as the atomic number increases across the period from left to right. The justification is due to the decreasing atomic radius from left to right in the periodic table, causing the valence electrons to be closer to the nucleus. As a result, more energy is required to separate these electrons from the atom.

In the column

Ionization energy decreases with an increase in atomic number across the group from top to bottom. The justification lies in the fact that the atomic radius increases when moving from top to bottom within the same column, resulting in the outermost electron being farther from the nucleus. Consequently, the attractive force between the outer electron and the nucleus weakens, making it easier to remove the electron, and thus ionization energy decreases from top to bottom within the same column.

Note: The second ionization energy is greater than the first, the third ionization energy is greater than the second, and the fourth ionization energy is greater than the third, and so on. The justification for this is that as electrons are removed, the positive charge of the nucleus has a greater effect, increasing its attraction for the remaining electrons. Consequently, more energy is needed to remove these electrons, resulting in an increasing ionization energy.

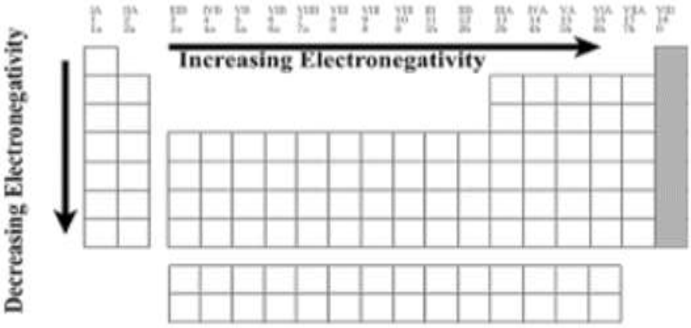


Electronegative periodicity χ

Electronegativity is the atomic ability to attract electrons in a chemical bond, denoted by the symbol χ.

In horizontal rows: Electronegativity increases when moving from left to right due to the increase in atomic number and the decrease in atomic radius. This leads to an increased nuclear attraction force towards the bonding electrons.

In vertical columns: Electronegativity decreases from top to bottom due to the increase in atomic number and atomic radius. This results in a decrease in the nuclear attraction force towards the bonding electrons.



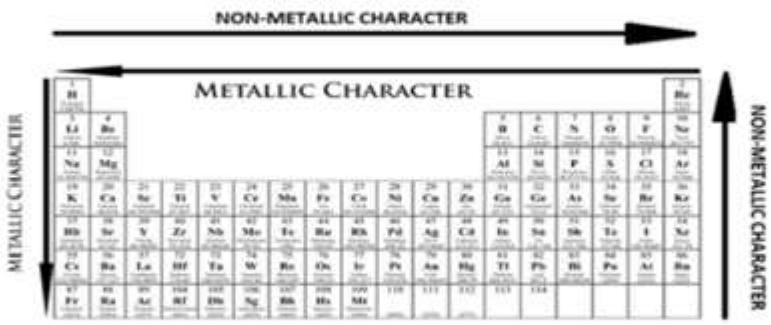
Gradation of Metallic Property

- Metals are chemical elements that lose electrons to form positive ions, acquiring the noble gas configuration.
- Nonmetals are elements that readily gain electrons.
- Metalloids exhibit characteristics of both metals and nonmetals, sometimes tending to lose electrons and at other times to gain electrons.

In horizontal periods: Metallic property decreases within a given period from left to right, starting with strong metallic elements. As the atomic number increases within the same period, the metallic property gradually decreases until reaching metalloids. Then, the nonmetallic property increases until reaching the strongest nonmetals in column 17.

Justification: Due to the decreasing atomic radii, we find that metallic property decreases from, for example, lithium, gradually reaching fluorine.

In vertical columns: Metallic property increases within a given column from top to bottom with the increase in atomic number. Justification: Due to the larger atomic size, such as in column 1. On the other hand, nonmetallic property decreases with the increase in atomic number within the same column. Justification: Due to lower electronegativity values, as seen in column 17.



Sanderson rule

The Sanderson rule states that an element is metallic if the number of electrons in its outermost shell (corresponding to the highest n value) is less than or equal to its group number, with exceptions for H, Ge, and Po.

Example:

Zn_{30} : $[Ar]_{18}$ $4S^2 3d^{10}$ 4>2 , 2 is the number of electrons in the outermost shell, which is 4. Therefore, Zn is metallic.

I_{53} : $[Kr]_{36}$ $5S^2 4d^{10} 5p^5$ 7>5 - Iodine (I) is not metallic as the number of electrons in its outermost shell (7) is greater than its group number.

Note: The exceptions mentioned are H, Ge, and Po.