جامـــعـــة المـــســتـقـبـل
AL MUSTAQBAL UNIVERSITY
كليـــــة العـلـــــوم

Analytical Chemistry

1st stage

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Lecture 3: Periodic Table

Department of Medical Biotechnology

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1.1 Classification of Elements

One of the simplest methods for classifying elements is to divide them into three categories:

Metals

- Good conductors of heat and electricity.
- Malleable and ductile with a shiny appearance.
- Found on the left and center of the periodic table (e.g., Fe, Cu).

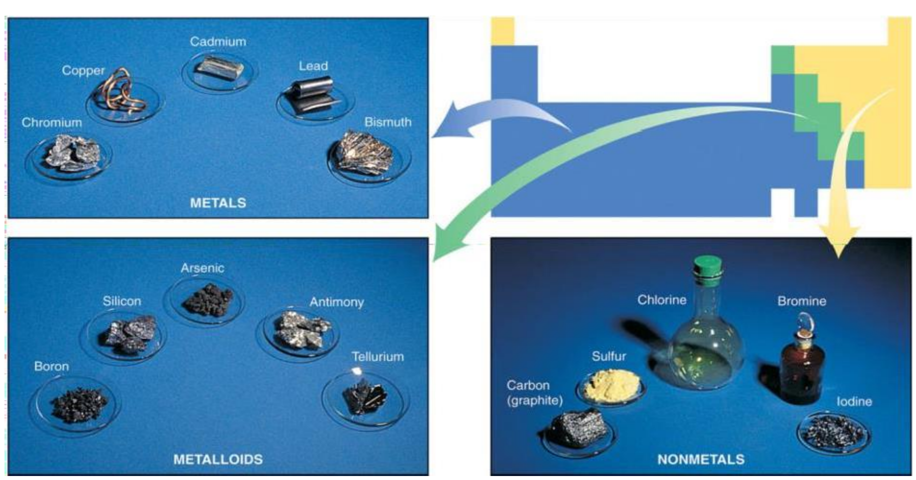
Non-Metals

- Poor conductors of heat and electricity.
- Brittle in solid form and often exist as gases or dull solids.
- Found on the right side of the periodic table (e.g., O, N).

Metalloids

- Exhibit properties of both metals and non-metals.
- Found along the "stair-step" line in the periodic table (e.g., Si, Br).

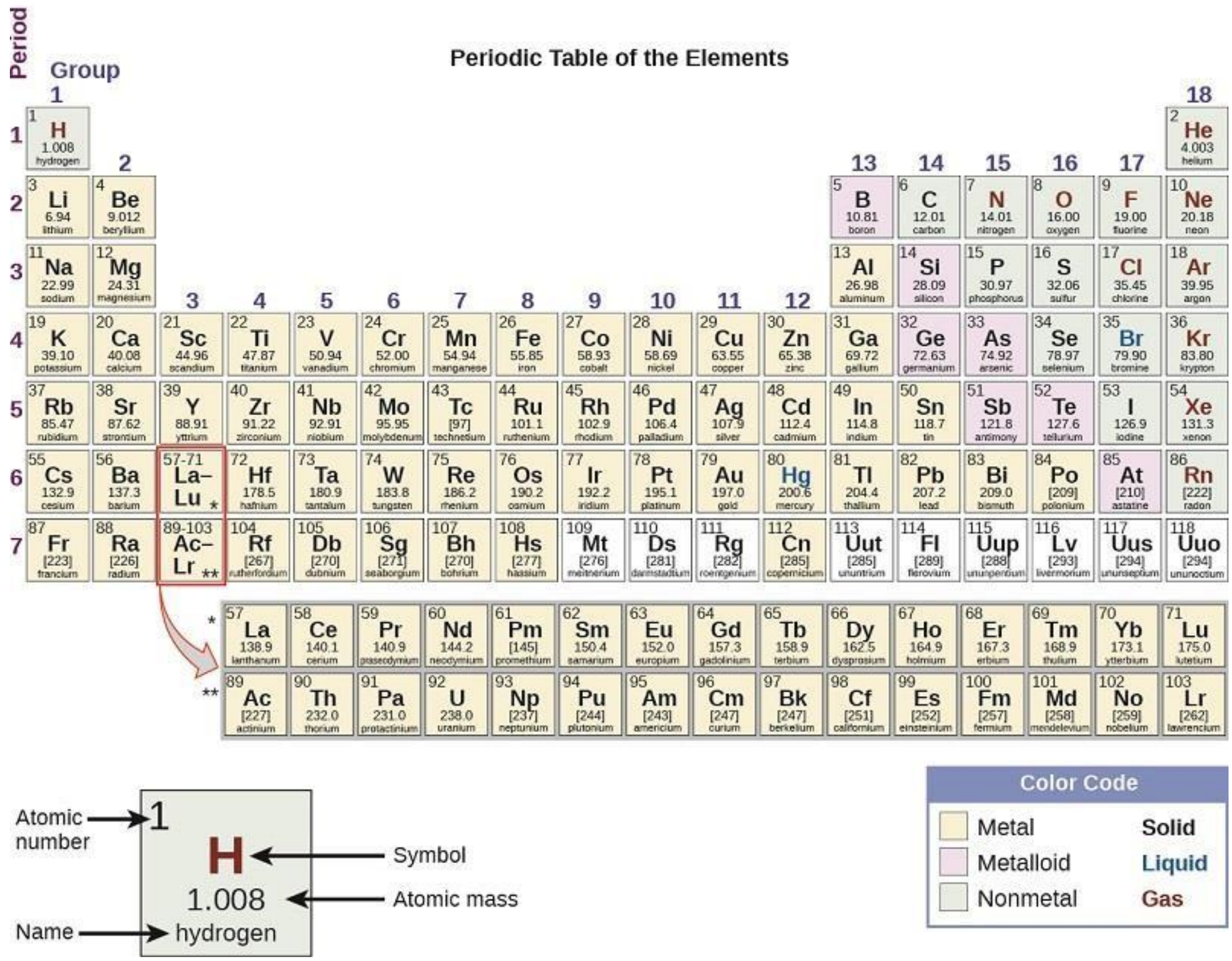
Metals, Metalloids, and Nonmetals



1.2 The Periodic Table

The periodic table is a systematic arrangement of chemical elements organized by their atomic number, electron configuration, and recurring chemical properties. It is a fundamental tool in chemistry, providing a clear framework for understanding the relationships between elements and predicting their behavior.

The idea of arranging the elements into a periodic table had beenconsidered by many chemists, but either the data to support the ideawere insufficient or the classification scheme was incomplete.Mendeleev and Meyer organized the elements in order of atomic weight and then identified families of elements with similar properties.By arranging these families in rows or columns, and by considering similarities in chemical behavior as well as atomic weight,Mendeleev found vacancies in the table and was not able to predict the properties of several elements (Gallium, Scandium, Germanium, Polonium) that had not yet been discovered.


Periodic Table of the Elements
Period
Group
1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 16 17 18
1 H 1.008 hydrogen
2 He 4.003 helium
3 Li 6.94 lithium
4 Be 9.012 beryllium
5 B 10.81 boron
6 C 12.01 carbon
7 N 14.01 nitrogen
8 O 16.00 oxygen
9 F 19.00 fluorine
10 Ne 20.18 neon
11 Na 22.99 sodium
12 Mg 24.31 magnesium
13 Al 26.98 aluminum
14 Si 28.09 silicon
15 P 30.97 phosphorus
16 S 32.06 sulfur
17 Cl 35.45 chlorine
18 Ar 39.95 argon
19 K 39.10 potassium
20 Ca 40.08 calcium
21 Sc 44.96 scandium
22 Ti 47.87 titanium
23 V 50.94 vanadium
24 Cr 52.00 chromium
25 Mn 54.94 manganese
26 Fe 55.85 iron
27 Co 58.93 cobalt
28 Ni 58.69 nickel
29 Cu 63.55 copper
30 Zn 65.38 zinc
31 Ga 69.72 gallium
32 Ge 72.63 germanium
33 As 74.92 arsenic
34 Se 78.97 selenium
35 Br 79.90 bromine
36 Kr 83.80 krypton
37 Rb 85.47 rubidium
38 Sr 87.62 strontium
39 Y 88.91 yttrium
40 Zr 91.22 zirconium
41 Nb 92.91 niobium
42 Mo 95.95 molybdenum
43 Tc [97] technetium
44 Ru 101.1 ruthenium
45 Rh 102.9 rhodium
46 Pd 106.4 palladium
47 Ag 107.9 silver
48 Cd 112.4 cadmium
49 In 114.8 indium
50 Sn 118.7 tin
51 Sb 121.8 antimony
52 Te 127.6 tellurium
53 I 126.9 iodine
54 Xe 131.3 xenon
55 Cs 132.9 cesium
56 Ba 137.3 barium
57-71 La– Lu *
72 Hf 178.5 hafnium
73 Ta 180.9 tantalum
74 W 183.8 tungsten
75 Re 186.2 rhenium
76 Os 190.2 osmium
77 Ir 192.2 iridium
78 Pt 195.1 platinum
79 Au 197.0 gold
80 Hg 200.6 mercury
81 Tl 204.4 thallium
82 Pb 207.2 lead
83 Bi 209.0 bismuth
84 Po [209] polonium
85 At [210] astatine
86 Rn [222] radon
87 Fr [223] francium
88 Ra [226] radium
89-103 Ac– Lr **
104 Rf [267] rutherfordium
105 Db [270] dubnium
106 Sg [271] seaborgium
107 Bh [270] bohrium
108 Hs [277] hassium
109 Mt [276] meitnerium
110 Ds [281] darmstadtium
111 Rg [282] roentgenium
112 Cn [285] copernicium
113 Uut [285] ununtrium
114 Fl [289] flerovium
115 Uup [288] ununpentium
116 Lv [293] livermorium
117 Uus [294] ununseptium
118 Uuo [294] ununoctium
*
57 La 138.9 lanthanum
58 Ce 140.1 cerium
59 Pr 140.9 praseodymium
60 Nd 144.2 neodymium
61 Pm [145] promethium
62 Sm 150.4 samarium
63 Eu 152.0 europium
64 Gd 157.3 gadolinium
65 Tb 158.9 terbium
66 Dy 162.5 dysprosium
67 Ho 164.9 holmium
68 Er 167.3 erbium
69 Tm 168.9 thulium
70 Yb 173.1 ytterbium
71 Lu 175.0 lutetium
**
89 Ac [227] actinium
90 Th 232.0 thorium
91 Pa 231.0 protactinium
92 U 238.0 uranium
93 Np [237] neptunium
94 Pu [244] plutonium
95 Am [243] americium
96 Cm [247] curium
97 Bk [247] berkelium
98 Cf [251] californium
99 Es [252] einsteinium
100 Fm [257] fermium
101 Md [258] mendelevium
102 No [259] nobelium
103 Lr [262] lawrencium
Atomic number
Name
1
H
1.008
hydrogen
Symbol
Atomic mass
Color Code
Metal
Metalloid
Nonmetal
Solid
Liquid
Gas

1.3 Structure of the Periodic Table

❖ Groups (Columns):

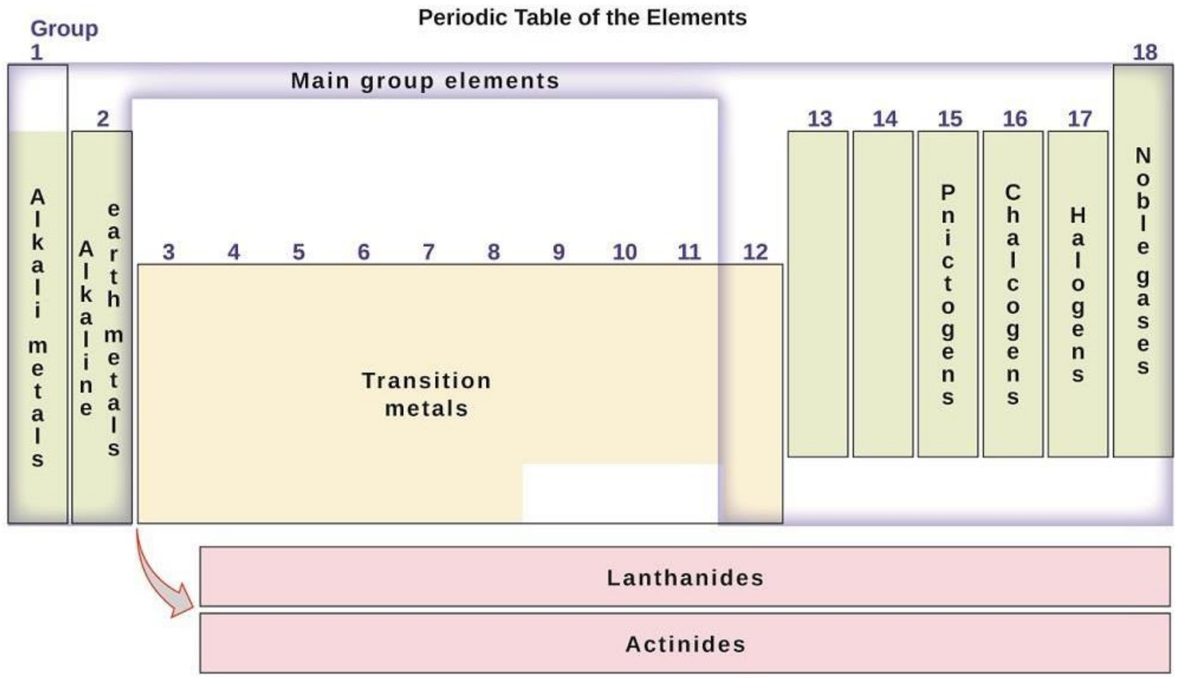
- There are 18 vertical columns called groups.
- Elements in the same group share similar chemical properties because they have the same number of valence electrons.
- Examples:
  - Group 1: Alkali metals (e.g., Lithium - Li).
  - Group 17: Halogens (e.g., Fluorine - F).
  - Group 18: Noble gases (e.g., Helium - He).

❖ Periods (Rows):

- There are 7 horizontal rows called periods.
- The period number corresponds to the number of electron shells in the elements of that row.

❖ Blocks:

- The table is divided into four main blocks based on the electron configuration of the elements:
  - **s-block**: Groups 1-2.
  - **p-block**: Groups 13-18.
  - **d-block**: Transition metals.
  - **f-block**: Lanthanides and actinides.



1.4 Electron Configuration

Electron configuration refers to the arrangement of electrons in an atom's orbitals around its nucleus. It determines the chemical and physical properties of an element. The electrons are distributed in energy levels or shells, which are further divided into subshells (s, p, d, and f).

Examples:

H_1: $1S^1$

Li_3: $1S^2$ $2S^1$

Na_{11}: $1S^2$ $2S^2$ $2P^6$ $3S^1$

K_{19}: $1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ $4S^1$

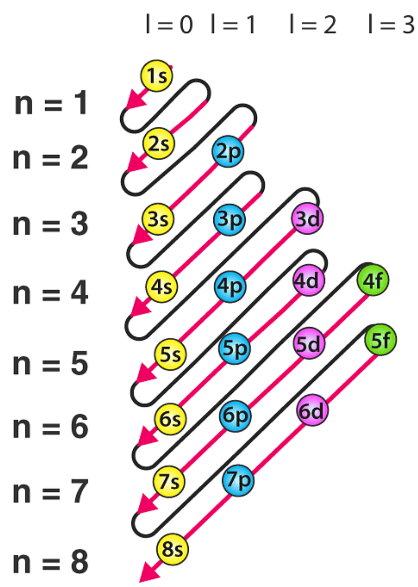
Rh_{37}: $1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ $4S^2$ $3d^{10}4P^6$ $5S^1$

Cs_{55}: $1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ $4S^2$ $3d^{10}4P^6$ $5S^2$ $4d^{10}5P^6$ $6S^1$

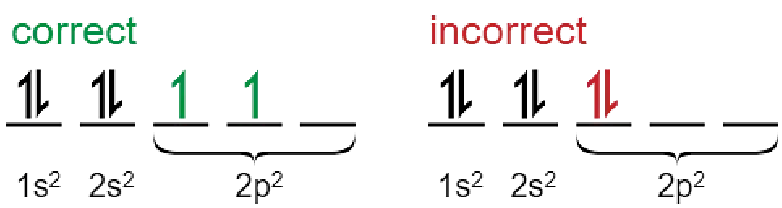
Fr_{87}: $1S^2$ $2S^2$ $2P^6$ $3S^2$ $3P^6$ $4S^2$ $3d^{10}4P^6$ $5S^2$ $4d^{10}5P^6$ $6S^2$ $5d^{10}4f^{14}6P^6$ $7S^1$

1.5 Principles Governing Electron Configuration

1. **Aufbau Principle**:
   - Electrons fill the lowest energy orbitals first before moving to higher energy levels.
   - Orbital filling order: 1s,2s,2p,3s,3p,4s,3d,4p,5s,…



2. **Pauli Exclusion Principle**:
   - Each orbital can hold a maximum of 2 electrons with opposite spins.
3. **Hund's Rule**:
   - In orbitals of equal energy (like the three p-orbitals), electrons fill each orbital singly before pairing up.

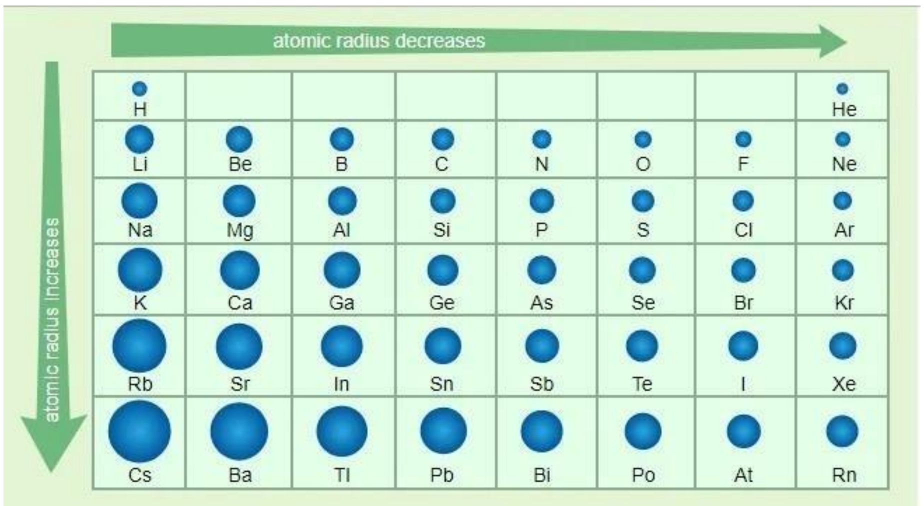


1.6 The Periodic table properties

1.6.1 Atomic Radius

Definition: The atomic radius is the distance from the nucleus to the outermost electron, or it can be described as half the distance between the centers of two adjacent atoms of the same element in a molecule.

- Decreases across a period (left to right) due to increasing nuclear charge.
- Increases down a group as more electron shells are added.



1.6.2 Ionization Energy

Definition: The ionization energy, EI, of an atom/ion is the minimum energy, which is required to remove an electron of an atom. The unit of ionization energy is kJ/mol.

- Increases across a period because atoms hold their electrons more tightly.
- Decreases down a group as electrons are farther from the nucleus.

$$M_{(g)} + \text{Energy} \longrightarrow M^{+}_{(g)} + e^{-}$$

1.6.3 Electronegativity

Definition: The electronegativity, χ, describes the ability of an atom to attract electrons towards.

- Increases across a period as the tendency to attract electrons in a bond strengthens.
- Decreases down a group due to increased distance from the nucleus.

Problem: Put in order of largest to smallest: F, Ar, Sr, and Cs.

Solution: Cs > Sr >Ar > F

1.7 Shielding

Each electron act as a shell for another electron and reducing the attraction between the nucleus and further electrons.

Slater's Rules: In 1930, J. C. Slater proposed a set of empirical rules to semiquantify the concept of effective nuclear charge. He proposed a formula that related

How can we calculate the effective charge of the nucleus?

$$Z_{eff} = Z - S$$

1. The electrons are written in groups as follows:

1s | 2s 2p | 3s 3p | 3d | 4s 4p | 4d | 4f | 5s 5p | 5d |....

2. All electrons in orbitals of greater principal quantum number

(At n+1) contribute zero.

3. for ns or np valence electrons:

a- Electrons in the same ns, np group (same principal quantum number) contributes (0.35).

b- Electrons in the (n-1) principal group contribute (0.85).

c- Electrons in the (n-2) or lower groups contribute (1.00).

4- For nd and nf valence electrons:

a- Electrons in the same nd and nf group contribute (0.35).

b- Electrons in the groups to the left contribute (1.00).

Example: to calculate the effective nuclear charge on one of the 2p

Electrons in the oxygen atom ($1S^2S^2P^4$), we first find the screening (or

Shielding) constant:

$S = (2 \times 0.85) + (5 \times 0.35) = 3.45$

$Z_{eff} = Z - S$

$= 8 - 3.45$

$= 4.55$