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## **Analytical Chemistry**

**1<sup>st</sup> stage**

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**Lecture 2: Periodic Table**

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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

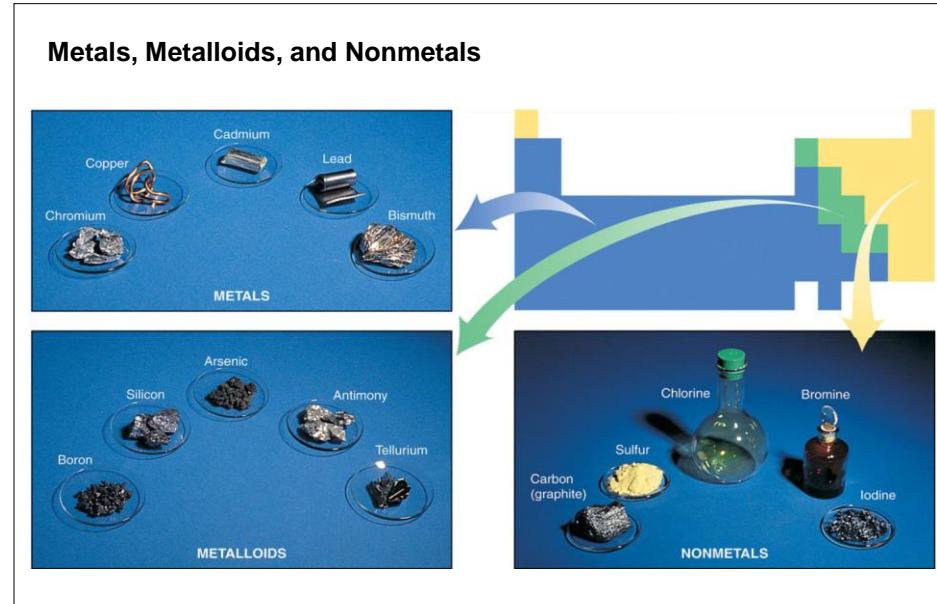
- Good conductors of heat and electricity.
- Malleable and ductile with a shiny appearance.
- Found on the left and center of the periodic table (e.g., Fe, Cu).

### 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).



## 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 been considered by many chemists, but either the data to support the idea were 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.

Period

Group

1

1

2

2

3

3

4

4

5

5

6

6

7

7

1

2

3

4

5

6

7

8

9

10

11

12

13

14

15

16

17

18

## Periodic Table of the Elements

1	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 1.008 hydrogen																	He 4.003 helium
2	Li 6.94 lithium	Be 9.012 beryllium																
3	Na 22.99 sodium	Mg 24.31 magnesium																
4	K 39.10 potassium	Ca 40.08 calcium	Sc 44.96 scandium	Ti 47.87 titanium	V 50.94 vanadium	Cr 52.00 chromium	Mn 54.94 manganese	Fe 55.85 iron	Co 58.93 cobalt	Ni 58.69 nickel	Cu 63.55 copper	Zn 65.38 zinc	Al 26.98 aluminum	Si 28.09 silicon	P 30.97 phosphorus	S 32.06 sulfur	Cl 35.45 chlorine	Ar 39.95 argon
5	Rb 85.47 rubidium	Sr 87.62 strontium	Y 88.91 yttrium	Zr 91.22 zirconium	Nb 92.91 niobium	Mo 95.95 molybdenum	Tc [97] technetium	Ru 101.1 ruthenium	Rh 102.9 rhodium	Pd 106.4 palladium	Ag 107.9 silver	Cd 112.4 cadmium	In 114.8 indium	Sn 118.7 tin	Ge 126.3 germanium	As 124.92 arsenic	Se 78.97 selenium	Kr 83.80 krypton
6	Cs 132.9 cesium	Ba 137.3 barium	La 57-71 lanthanum	Hf 178.5 hafnium	Ta 180.9 tantalum	W 183.8 tungsten	Re 186.2 rhenium	Os 190.2 osmium	Ir 192.2 iridium	Pt 195.1 platinum	Au 197.0 gold	Hg 200.6 mercury	Tl 204.4 thallium	Pb 207.2 lead	Bi 209.0 bismuth	Po [209] polonium	At [210] astatine	Rn [222] radon
7	Fr [223] francium	Ra [226] radium	Ac 89-103 actinium	Rf [267] rutherfordium	Db [270] dubnium	Sg [271] seaborgium	Bh [270] bohrium	Hs [277] hassium	Mt [276] meitnerium	Ds [281] darmstadtium	Rg [282] roentgenium	Cn [285] copernicium	Uut [289] florovium	Uup [288] ununpotassium	Lv [293] livermorium	Uus [294] ununseptium	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 [251] 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 → 1  
 Symbol → H  
 Name → hydrogen  
 Atomic mass → 1.008

## Color Code

Metal	Solid
Metalloid	Liquid
Nonmetal	Gas

## 1.3 Structure of the Periodic Table

### ❖ Groups (Columns):

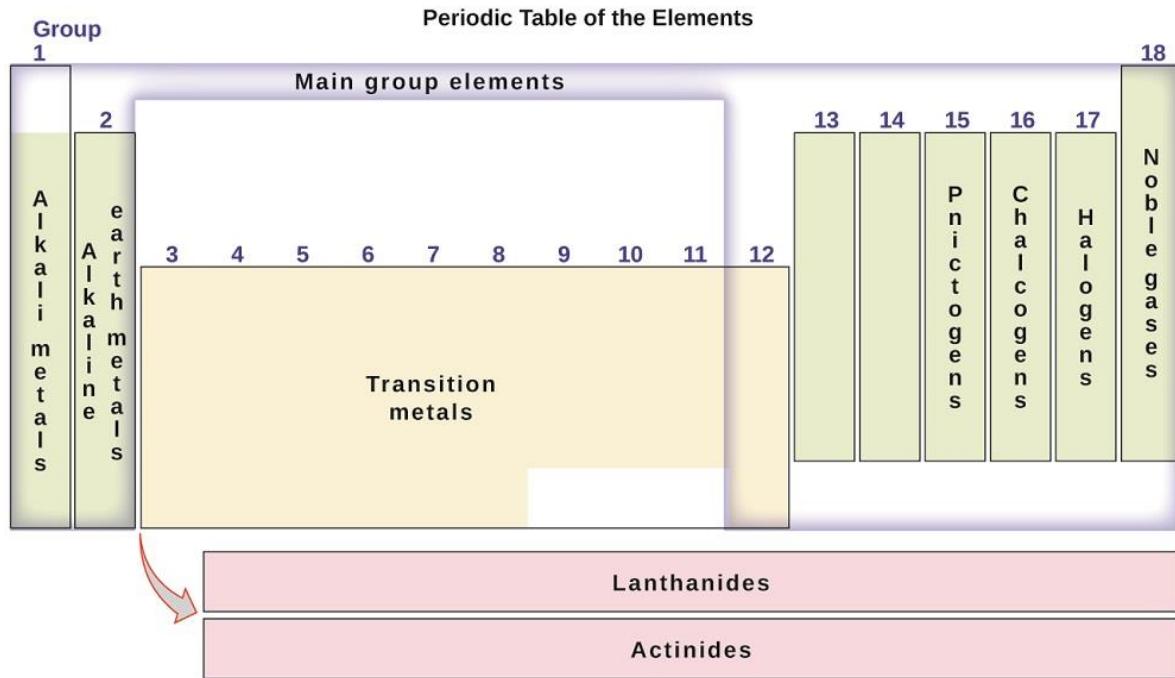
- There are 18 vertical columns called groups.
- Elements in the same group share similar chemical properties because they have the same number of valence electrons.
- Examples:
  - Group 1: Alkali metals (e.g., Lithium - Li).
  - Group 17: Halogens (e.g., Fluorine - F).
  - Group 18: Noble gases (e.g.,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

$\text{H}_1: 1\text{S}^1$

$\text{Li}_3: 1\text{S}^2 2\text{S}^1$

$\text{Na}_{11}: 1\text{S}^2 2\text{S}^2 2\text{P}^6 3\text{S}^1$

$\text{K}_{19}: 1\text{S}^2 2\text{S}^2 2\text{P}^6 3\text{S}^2 3\text{P}^6 4\text{S}^1$

$\text{Rh}_{37}: 1\text{S}^2 2\text{S}^2 2\text{P}^6 3\text{S}^2 3\text{P}^6 4\text{S}^2 3\text{d}^{10} 4\text{P}^6 5\text{S}^1$

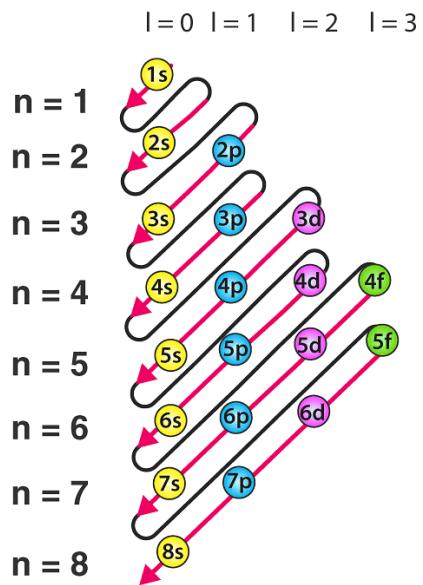
$\text{Cs}_{55}: 1\text{S}^2 2\text{S}^2 2\text{P}^6 3\text{S}^2 3\text{P}^6 4\text{S}^2 3\text{d}^{10} 4\text{P}^6 5\text{S}^2 4\text{d}^{10} 5\text{P}^6 6\text{S}^1$

$\text{Fr}_{87}: 1\text{S}^2 2\text{S}^2 2\text{P}^6 3\text{S}^2 3\text{P}^6 4\text{S}^2 3\text{d}^{10} 4\text{P}^6 5\text{S}^2 4\text{d}^{10} 5\text{P}^6 6\text{S}^2 5\text{d}^{10} 4\text{f}^{14} 6\text{P}^6 7\text{S}^1$

## 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, ...

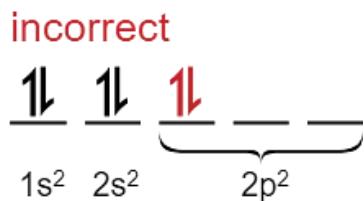
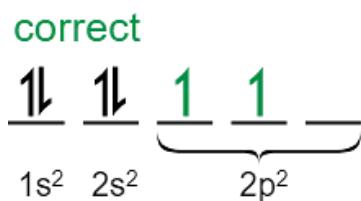


### 2. Pauli Exclusion Principle:

- 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.**

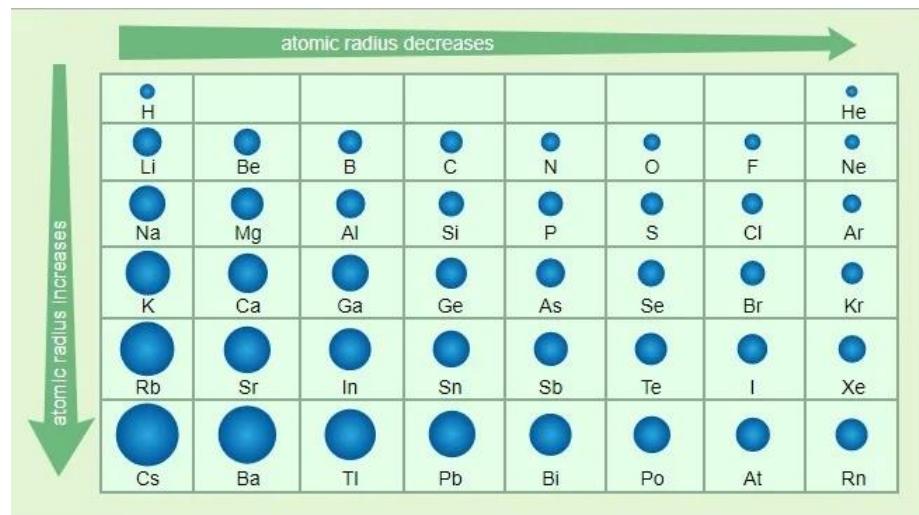


## 1.6 The Periodic table properties

### 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.



### 1.6.3 Electronegativity

**Definition:** The electronegativity,  $\chi$ , 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?

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

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:

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

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

$$= 8 - 3.45$$

$$= 4.55$$