

Analytical Chemistry

1st 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

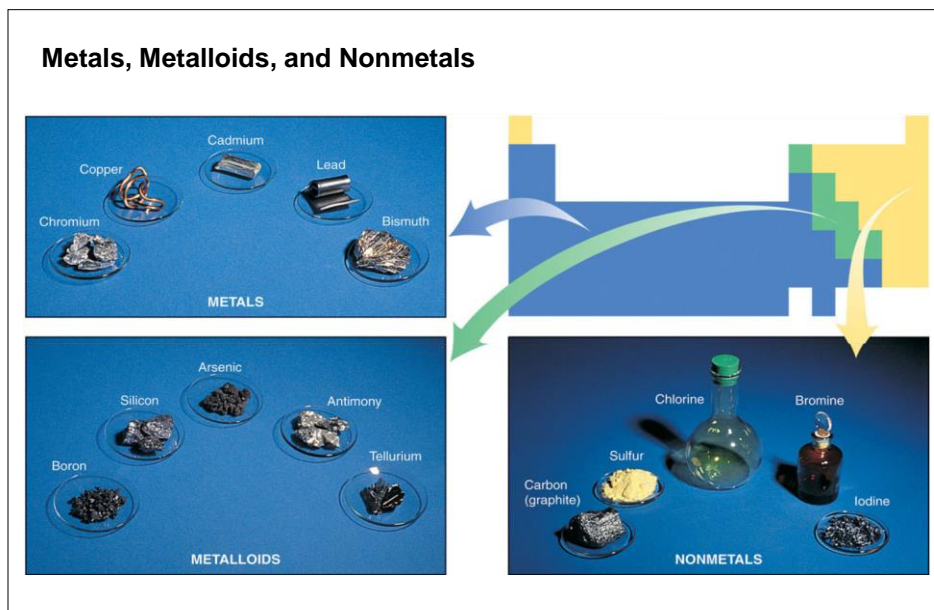
- Good conductors of heat and electricity.
- Malleable and ductile with a shiny appearance.
- Found on the left and center of the periodic table (e.g., Fe, Cu).

⚡ 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

Periodic Table of the Elements

1																	18					
1	1 H 1.008 hydrogen																	2 He 4.003 helium				
2	3 Li 6.94 lithium	4 Be 9.012 beryllium															13 B 10.81 boron	14 C 12.01 carbon	15 N 14.01 nitrogen	16 O 16.00 oxygen	17 F 19.00 fluorine	18 Ne 20.18 neon
3	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
4	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				
5	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				
6	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				
7	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 → 1

Symbol → **H**

Atomic mass → 1.008

Name → hydrogen

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.

Periodic Table of the Elements

Group 1	Main group elements																18
2												13	14	15	16	17	
Alkali metals	Alkaline earth metals	3	4	5	6	7	8	9	10	11	12			Pnictogens	Chalcogens	Halogens	Noble gases
		Transition metals															
		Lanthanides															
		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:

H₁: 1S¹

Li₃: 1S² 2S¹

Na₁₁: 1S² 2S² 2P⁶ 3S¹

K₁₉: 1S² 2S² 2P⁶ 3S² 3P⁶ 4S¹

Rh₃₇: 1S² 2S² 2P⁶ 3S² 3P⁶ 4S² 3d¹⁰ 4P⁶ 5S¹

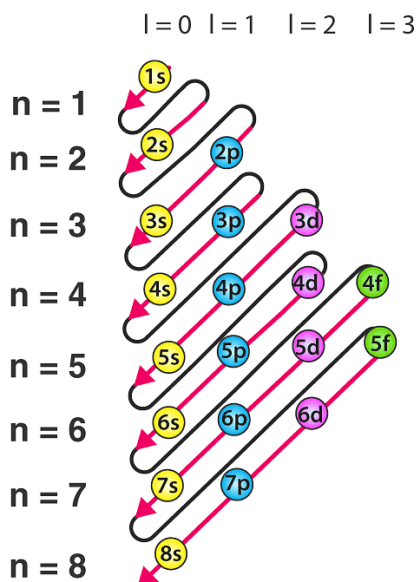
Cs₅₅: 1S² 2S² 2P⁶ 3S² 3P⁶ 4S² 3d¹⁰ 4P⁶ 5S² 4d¹⁰ 5P⁶ 6S¹

Fr₈₇: 1S² 2S² 2P⁶ 3S² 3P⁶ 4S² 3d¹⁰ 4P⁶ 5S² 4d¹⁰ 5P⁶ 6S² 5d¹⁰ 4f¹⁴ 6P⁶ 7S¹

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: $1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, \dots$

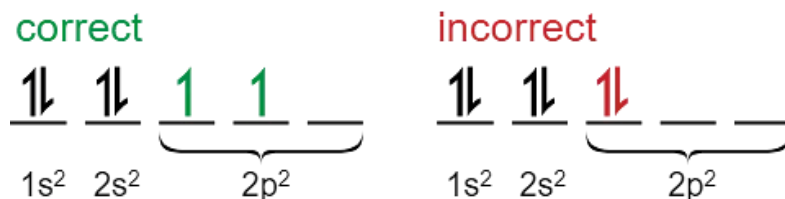


2. Pauli Exclusion Principle:

- 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**.

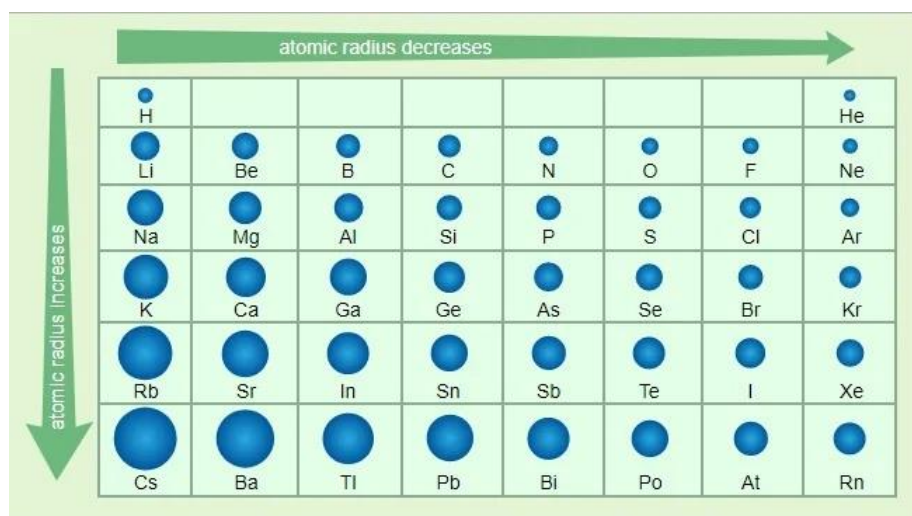


1.6 The Periodic table properties

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.

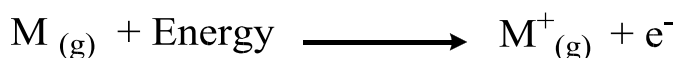
- **Decreases** across a **period** (left to right) due to increasing nuclear charge.
- **Increases** down a **group** as more electron shells are added.



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, 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: $\text{Cs} > \text{Sr} > \text{Ar} > \text{F}$

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:

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

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

$$= 8 - 3.45$$

$$= 4.55$$