How elements are arranged in periodic table?

⇒ atomic number.

↳ no. of electrons.

<u>group 1</u> :-

Li
Na
K
Rb
Cs

⇒ elements in a same group shows similar chemical properties.

"Periodicity"

Chemical properties depends upon no. of outermost shells (valence shells)

PHYSICAL PROPERTIES :-

Apperances :-

Metals :- lustre (metallic lustre)

Non Metals :- dull
more matte
sometimes transparent too (gases)

Density :-

↳ respre by ρ → rho (greek letter)

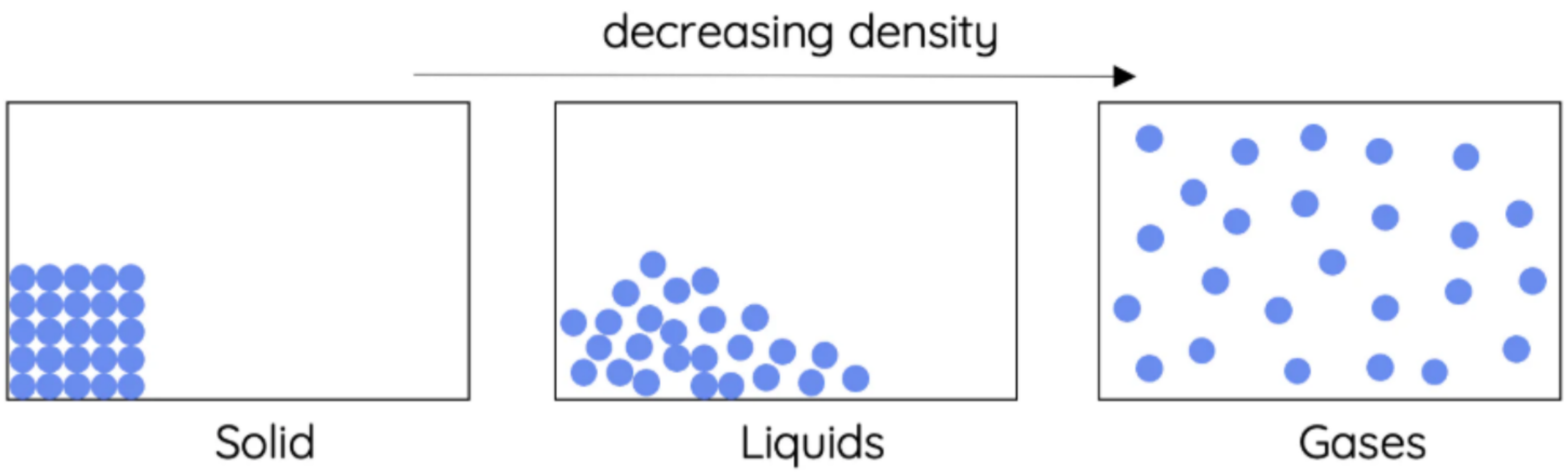
$$\rho = \frac{Mass}{Volume}$$

Mass → kg

Volume → m^3

$$\rho = kg/m^3 \quad -①$$

$M = g \quad V = ml \quad \rho = g/ml$



- Metals have variable densities but collectively tend to have higher densities than semi-metals and non-metals. For example, Tungsten (W) has a very high density of 19.3 g/cm^3 whereas aluminium has a lower density of 2.6 g/cm^3.
- Non-metals have variable densities. Non-metals in covalent molecular structures have lower densities, whereas those in network structures have higher densities. For example, Hydrogen gas has a very lower density of 0.09 g/cm^3 whereas diamond (allotrope of carbon) has much higher density of 3.5 g/cm^3.
- Semi-metals tend to have consistently densities in between the high end of metals and majority of non-metals. For example, germanium's density is 5.2 g/cm^3.

Electrical Conductivity :-

↳ may be it has free electrons. (delocalized)

↳ may be free ions (charged atoms)

- Metals are known for their excellent electrical conductivity. The metallic structure contains a sea of delocalised electrons surrounding metal atoms. These electrons are mobile and account for metals' relatively high electrical conductivity.
- Most non-metals are poor conductors because their electrons are more tightly bound and not free to move. Graphite, a form of carbon, is a notable exception due to its unique structure that allows electrons to move freely within layers.
- Metalloids have variable conductivity, which can often be enhanced by adding impurities in a process known as doping. Silicon (Si), for example, is a semiconductor widely used in electronic devices.

① Explain why metals conduct electricity (2-3)

② How metalloids can conduct electricity. (2)

Ductile & Malleability: -

Metals

Most metals are both ductile and malleable. This is because the structure of most metals (repeating atoms layered on top of each other) allows for relatively easy sliding movement between them without breaking the attraction between the metallic atoms and surrounding electrons.

Most metals deform well, whether being stretched (exhibiting ductility) or being compressed (which is malleability). Generally if a metal is ductile, it is also malleable (at the same temperature). However, lead is an example of a metal that is malleable (deforms under compressive stress), but not ductile (easily fractures under tensile stress).

Semi-metals

Metalloids are less malleable and ductile than metals. They are more likely to break or shatter when subjected to stress, similar to nonmetals.

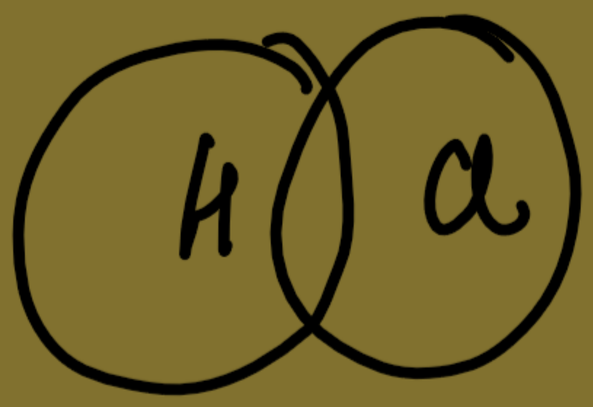
Non-metals

Non-metals are generally brittle in their solid forms and lack the malleability and ductility of metals. When force is applied, covalent bonds between atoms are more likely to break, leading to fracture.

(Key points)

Properties	Metals	Metalloids	Non-metals
Appearance	Lustrous	Low sheen	Dull
Melting & Boiling Point	Variable, mostly high and solids at room temperature	High	Variable, mostly low and gases at room temperature
Electrical conductivity	High	Intermediate (semi-conductors)	Mostly poor
Thermal conductivity	High	Intermediate	Mostly poor
Malleability and ductility	High	Intermediate	Nil (brittle)
Density	High	Intermediate	Low
Tensile Strength	High	Variable	Low
Hardness	Variable	High	Variable
Solubility	Insoluble	Insoluble	Variable

Electronegativity :-



→ covalent bond
↳ bonds formed by sharing of electron pairs.

ideally H —•—•— Cl

some element have more tendency to attract shared pair of electron.

Cl > H EN.

$\delta+$ $\delta-$
H —•—•— Cl

δ → delta (greek)
↳ really small amount (partial charge)

Electronegativity Scale :-

Pauling's Scale :-

Fluorine (4)
Oxygen
Nitrogen
] Learn)

Electronegativity Trends in the Periodic Table

Increasing electronegativity (→)

Decreasing electronegativity (↓)

1 H 1.008 Hydrogen																	2 He 4.003 Helium
3 Li 6.941 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 Aluminium	14 Si 28.09 Silicon	15 P 30.97 Phosphorus	16 S 32.07 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 72.64 Germanium	33 As 74.92 Arsenic	34 Se 78.96 Selenium	35 Br 79.90 Bromine	36 Kr 83.80 Krypton
37 Rb 85.47 Rubidium	38 Sr 87.61 Strontium	39 Y 88.91 Yttrium	40 Zr 91.22 Zirconium	41 Nb 92.91 Niobium	42 Mo 95.96 Molybdenum	43 Tc 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 Caesium	56 Ba 137.3 Barium	57–71 Lanthanoids	72 Hf 178.5 Hafnium	73 Ta 180.9 Tantalum	74 W 183.9 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 Polonium	85 At Astatine	86 Rn Radon
87 Fr Francium	88 Ra Radium	89–103 Actinoids	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds Darmstadtium	111 Rg Roentgenium	112 Cn Copernicium	113 Nh Nihonium	114 Fl Flerovium	115 Mc Moscovium	116 Lv Livermorium	117 Ts Tennessine	118 Og Oganesson

Period :- As we move left to right in a period EN increases (upto group 17)

group :- As we move down the group EN decreases.

IMP :-

- **Across a Period**: Electronegativity increases from left to right across a period. This is because atoms have more protons (increasing the nuclear charge) and the same number of core electrons (electrons in shells closer to the nucleus), pulling the valence electrons closer and making the atoms more effective at attracting bonding electrons.
- **Down a Group**: Electronegativity decreases down a group. As the atomic number increases, so does the number of electron shells. This increases the distance between the nucleus and the valence electrons, reducing the nuclear attraction experienced by bonding electrons.