HISTORY OF THE PERIODIC TABLE

Antoine Lavoisier (1789)

He made the first ever list of elements (33)

He classified them into metals and non-metals

The list included oxygen, nitrogen, hydrogen, phosphorus, mercury, zinc, Sulphur, and two others, light and caloric which at the time were believed to be elements

Johann Dobereiner (1829)

He categorized the elements into **triads** (groups of 3)

Where the middle element of each group's atomic weight was approx. the average of the other two

He identified a **Halogen triad** (chlorine, bromine, iodine)
He identified an **Alkali metal triad** (lithium, sodium, potassium)

Pros

- He offered an early insight on the relationship between atomic weights and properties
- He categorized the elements based on similarities lead to the modern day foundation

Cons

- His limitations for only 3 elements per triad/group lead to limited explanatory power and inaccuracy

Table 5.1

Group A element	Atomic mass	Group B element	Atomic mass	Group C elements	Atomic mass
N	14.0	Ca	40.1	Cl	35. 5
P	31.0	Sr	87.6	Br	79.9
As	74.9	Ba	137.3	I	126.9

Alexandre-Emile Beguyer (1862)

He put together the first version of the periodic table by ordering elements based on their atomic weight

John Newlands (1864)

He announced the **law of octaves**

All of the elements known at the time were ordered into relative atomic mass, he discovered that elements were similar to other elemants every 8 places later on the new table

In attempt to keep the cycle going, he sometimes mashed multiple elements into the same cell

Flaws

- The mashing lead to inaccuracy

Newlands' Octaves						
H	Li	Ga	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe
Co,Ni	Cu	Zn	Y	In	As	Se
Br	Rb	Sr	Ce,La	Zr	Di,Mo	Ro,Ru
Pd	Ag	Cd	U	Sn	Sb	Te
I	Cs	Ba,V	Ta	W	Nb	Au
Pt,Ir	Tl	Pb	Th	Hg	Bi	Cs

Lothar meyer (1864)

He organized 28 different elements into 6 different families

Each family's elements have similarities in both physical and chemical properties

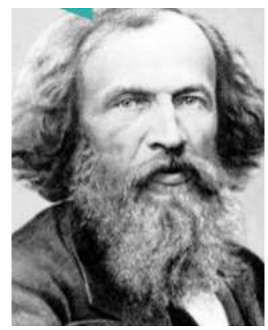
it used valence numbers to order and pattern out stuff

his work did not get published until 1870

Cons

- Most elements weren't discovered yet

Dmitri Mendeleev (1869)



He took 63 elements and arranged them by their atomic mass

He grouped them based on their physical and chemical properties

He left some slots for future elements

He predicted new elements such as silicon and boron

Flaws

- Noble gasses were not found
- Ordering by atomic mass isn't good

I	II	III	IV	V	VI	VII	VIII		
H 1.01									
Li 6.94	Be 9.01	B 10.8	C 12.0	N 14.0	O 16.0	F 19.0			
Na 23.0	Mg 24.3	Al 27.0	Si 28.1	P 31.0	S 32.1	Cl 35.5			
K 39.1	Ca 40.1		Ti 47.9	V 50.9	Cr 52.0	Mn 54.9	Fe 55.9	Co 58.9	Ni 58.7
Cu 63.5	Zn 65.4			As 74.9	Se 79.0	Br 79.9			
Rb 85.5	Sr 87.6	Y 88.9	Zr 91.2	Nb 92.9	Mo 95.9		Ru 101	Rh 103	Pd 106
Ag 108	Cd 112	In 115	Sn 119	Sb 122	Te 128	I 127			
Ce 133	Ba 137	La 139		Ta 181	W 184		Os 194	Ir 192	Pt 195
Au 197	Hg 201	Tl 204	Pb 207	Bi 209					
			Th 232		U 238				

William Ramsay (1894)

He discovered noble gases

Henry Moseley (1913)

He used an electron gun to shoot at elements which will produce x-ray waves

He took the waves and measured the frequency

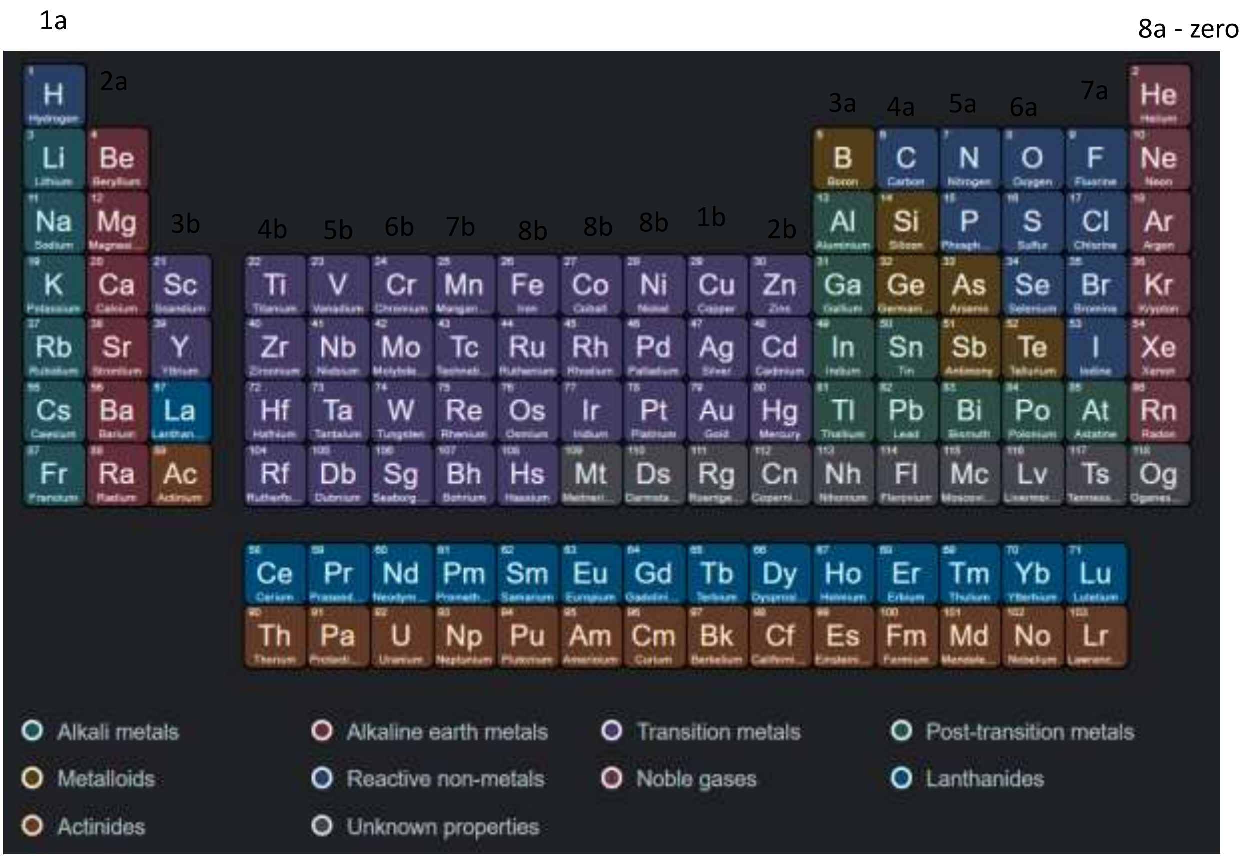
He used the results to organize them by their atomic number

Group 0	I a b	II a b	III a b	IV a b	V a b	VI a b	VII a 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

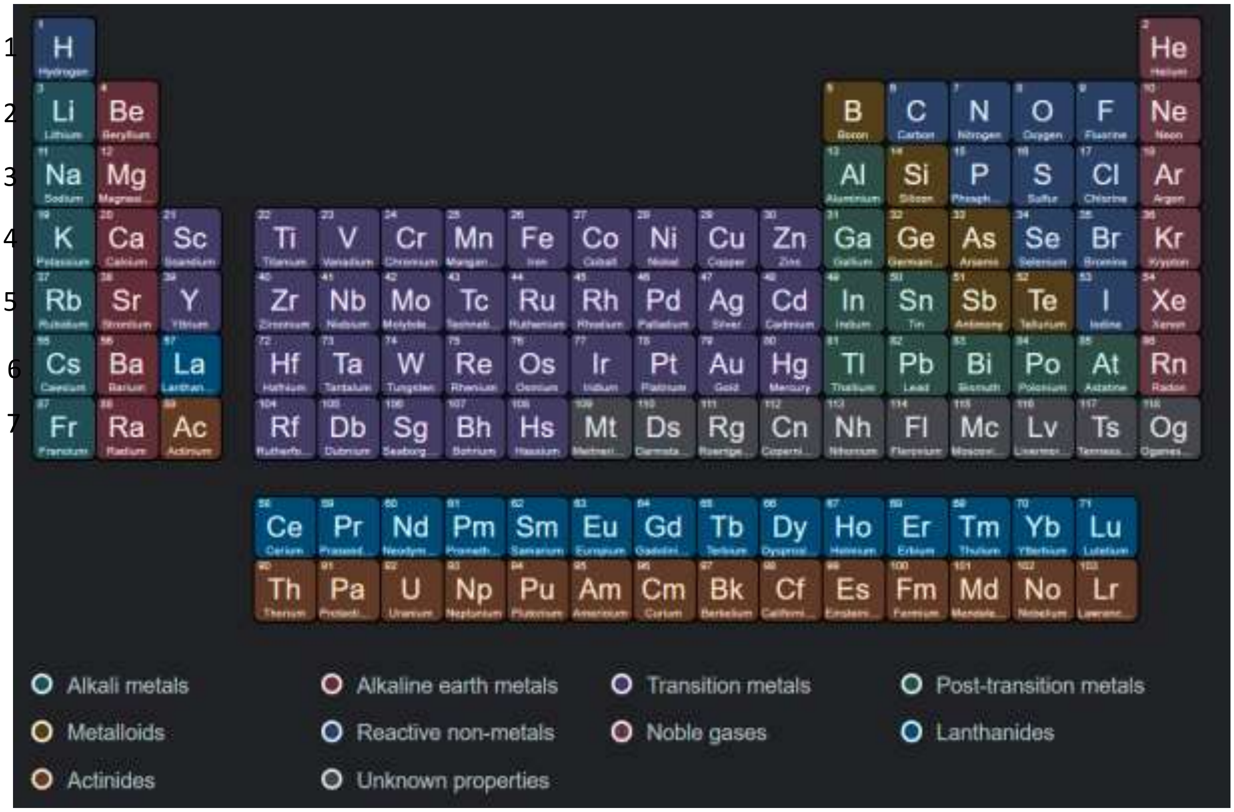
We classify the periodic table in 3 ways

Groups | Columns (18)



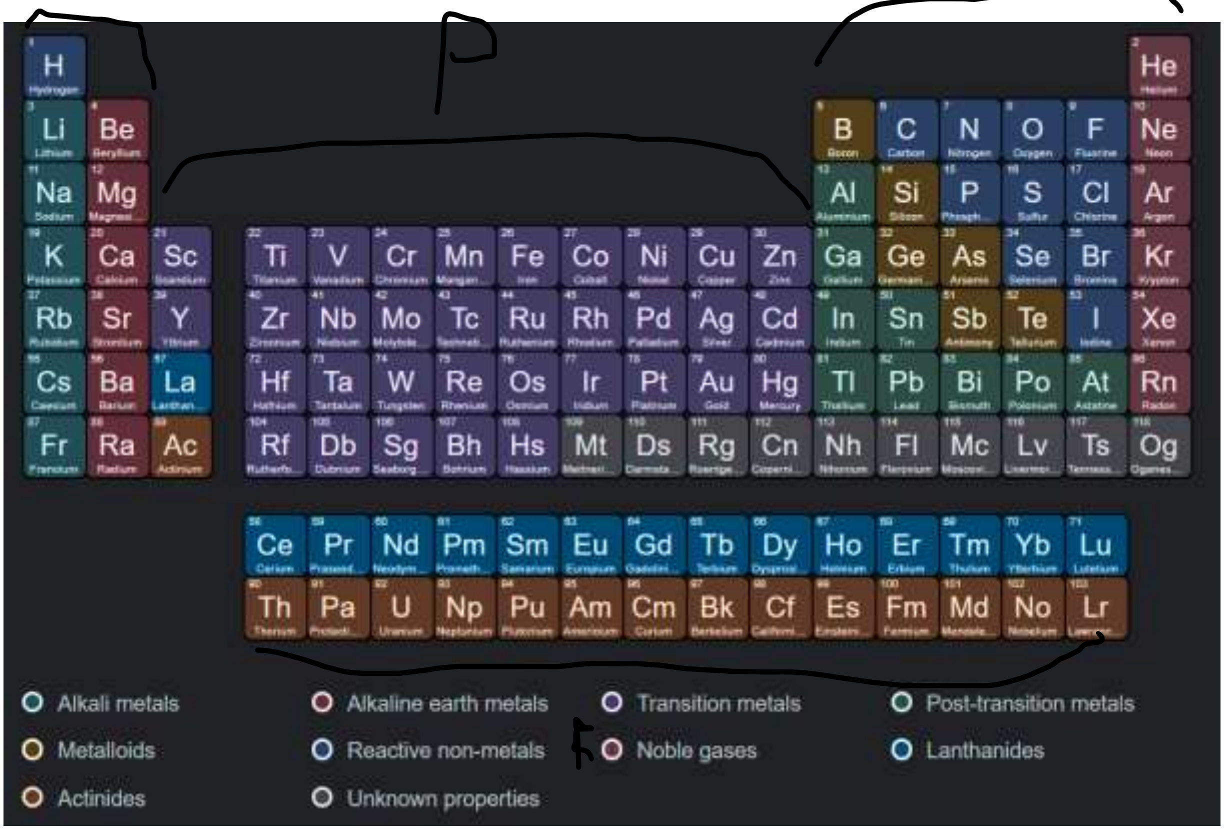
Electrons in the same group have the same number of valence electrons, therefor have similar chemical properties

Periods | Rows (7)



Electrons in the same group have the same number of energy levels

Blocks (4)



They are classified by their last filled sublevel
so if you're something like Na -> 1s2 2s2 2p6 3s1 -> your last sublevel is s -> you're in s block

S -> 1A, 2A

P -> 3A, 4A, 5A, 6A, 7A, 8A/zero

D -> 3:12 "transitional elements"

F -> contains all elements under the table

- **Actinides** are the most radioactive elements in the periodic table
- **Lanthanides** are known as inner transitional elements

Elements

Alkali Metals

Are found in **1A** except hydrogen

- They have 1 electron in their outer shell
- They react with water
- They are extremely reactive, that’s why you can’t find them in their pure state
- They are good malleable, ductile
- They are good heat, electricity conducters
- They explode when in water
- Cesium and francium are their heaviest

Alkaline Earth Metals

Are found in **2A**

- They are reactive but not quite as reactive as alkali metals

Metals

Are found in groups 3 to 12 or the whole **D** block

- They are solid at room temperature except mercury
- They are very malleable
- They are ductile (good conducters of heat and electricity)

Non-Metals

Are hydrogen, carbon, nitrogen, phosphorus,

oxygen, Sulphur, selenium, fluorine, chlorine, iodine and

astatine

- They are brittle and poor conducters

The melting points of **sulfer and phosphorus** are much lower than those of **silicon**

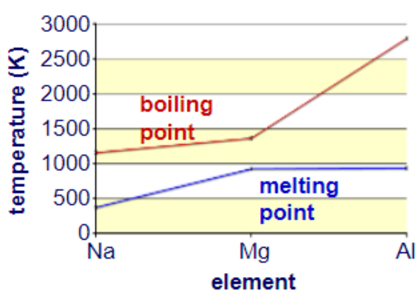
Because they have a simple molecular structure with **weak van der Waals** forces holding the molecules together, so they break quickly

As for **corrosion**, a metal's resistance to corrosion is related to its **reactivity series**, **gold** is an unreactive materials and does not corrode easily

Rusting is the corrosion of **iron**

FACTS FROM THE RELATION BETWEEN Na, Mg, Al

The melting and boiling points increase because of the strength of the metalic bonds, because more energy is need to break them



This metallic bond strength comes from

- **Charge density is** the size of an ion's charge to its size, the more the charge, the more they are attracted to delocalized electrons
- **Number of free electrons** Sodium has 1 free electron per metal ion, this leads to more attraction that must be broken jin electron

Silicon has a **macromolecular structure** similar to diamonds

Each silicon atom is bonded to 4 neighboring atoms, that leads to them being hard to melt, it requires a lot of energy so the melting and boiling points are high

Carbon is a very common non-metal as it's the fourth most common elemnt in the universe and nearly a fifth of the human body is **carbon**

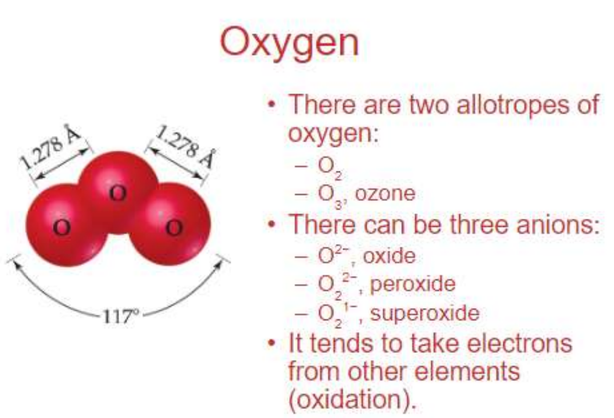
It can exist purely in 3 forms

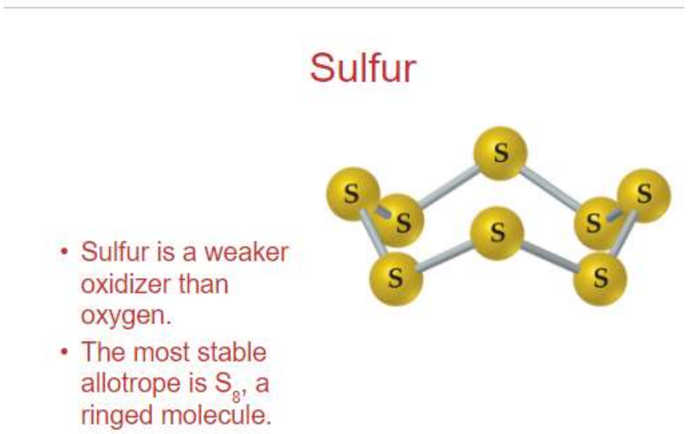
Diamond, which is the hardest substance known.

Graphite, which you will find in your pencil.

Fullerenes, which are football-shaped molecules.

Aluminum does not corrode because there is a thin layer of **aluminum oxide**





Halogens

Are non-metals found in **7A** except **Tennessine & Astatine**

- They very reactive non-metals
- Often found colorful and corrosive
- They are named halogen meaning **salt formers**

Noble Gasses

They are found in 8A or group zero (helium, neon,

argon, krypton, xenon and radon).

- They are colorless unreactive gasses
- They have the maximum number of electrons in their outermost energy level (2 in helium, 8 in others)

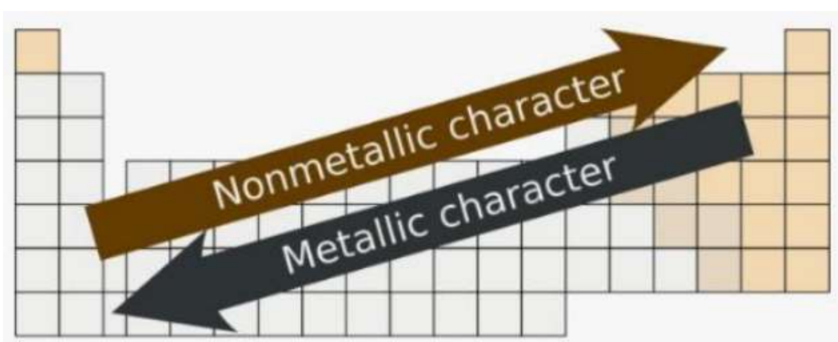
Metalloids

They are "boron, silicon, boron, germanium, arsenic, antimony and tellurium"

- They are semi-conductive
- They carry charge at special conditions
- They can vary from 3-6 valence electrons

Property	Metals	Non-Metals	Metalloids
Physical State	Solid (except Mercury)	Solid, Liquid, or Gas	Solid
Luster	Shiny	Dull	Variable (can be shiny or dull)
Conductivity	Good conductors of heat and electricity	Poor conductors of heat and electricity	Variable (semi-conductors)
Malleability	Malleable (can be hammered into thin sheets)	Not malleable	Variable (some are malleable)
Ductility	Ductile (can be drawn into wires)	Not ductile	Variable (some are ductile)
Density	Generally high density	Lower density	Variable
Melting/Boiling Points	High melting/boiling points	Lower melting/boiling points	Variable
Electron Configuration	Tend to lose electrons to form positive ions	Tend to gain or share electrons to form negative ions	Can exhibit both characteristics
Examples	Iron, Copper, Gold, Aluminum	Oxygen, Carbon, Nitrogen	Silicon, Germanium, Arsenic

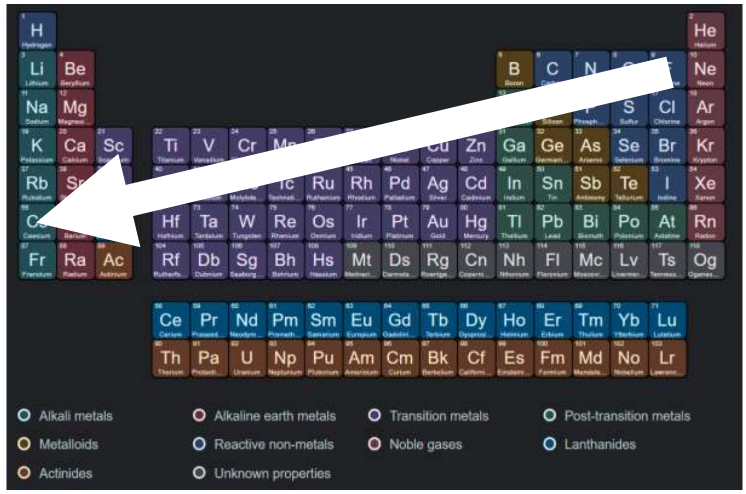
Jons Jacob Berzelius is the one who differentiated between the metals and nonmetals



Periodic Trends

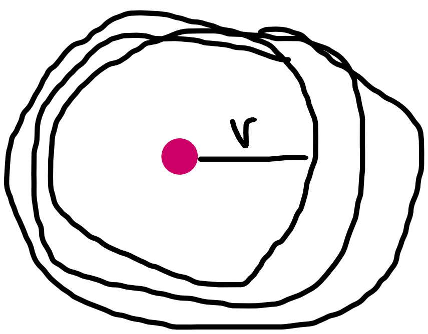
Periodicity is a repeating pattern of properties across the periodic table

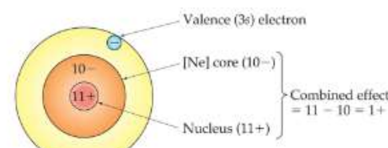
Atomic Radius



Is the size of an atom, it's determined by the boundaries of the valence number

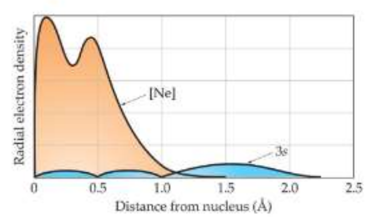
Another definition is that it's half the shortest internuclear distance found in the structure of an element





(a)

the bigger the atom, **the smaller it's Z_{eff} is**



what's an Effective nuclear charge

it's the amount of positive (nuclear) charge experienced pulling an electron in a multi-electron atom

the nuclear effective charge (Z_{eff}) of an electron is determined by

- The atomic number Z
- The screening constant S

$Z_{eff} = Z - S$

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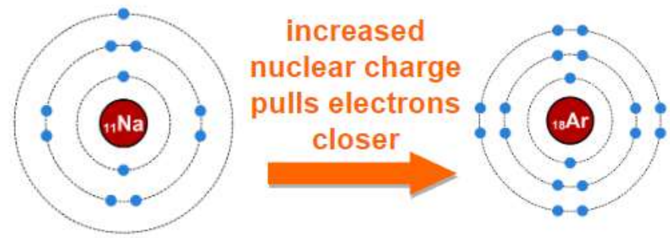
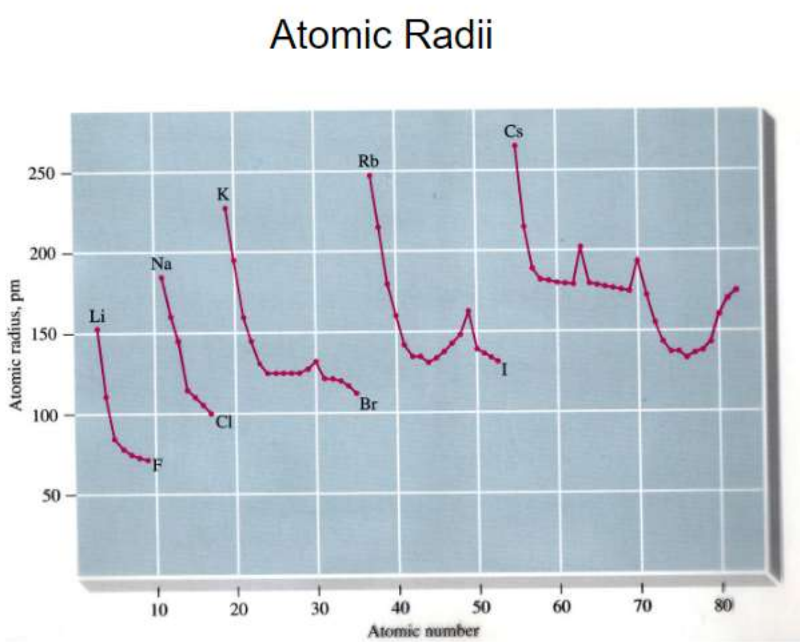
LEFT TO RIGHT (decreasing proton count)

The atomic radius increases when going from the right to the left side of the screen, and that is because moving "right to left" means decreasing the atomic number, aka decreasing the nuclear charge pulling the electrons to the atom, leaving them have some **space**



UP TO DOWN (increasing energy levels)

The more energy levels, the more the radius, simple



for non-metallic elements, instead of an atomic radius, there is a **covalent radius**, it's half the internuclear distance between two identical **bonded** atoms of a single covalent bond

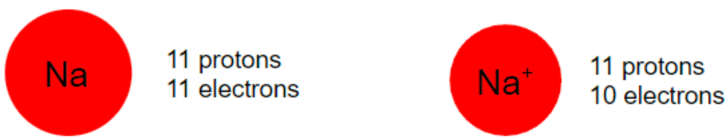
for **non-bonded adjacent atoms** (covalent crystal of a non-metallic element) **van der Waals radius** is used as a value for the atomic radius, it's half the shortest internuclear distance between two **non-bonded** atoms

ionic Radius

It’s the atomic radius but for an ion

When an atom loses one or more electrons, it becomes a **cation**

Cations are smaller than their stable atoms, because of the extra positive charge pulling more



When an atom gains one or more electrons, it becomes an **anion**

Anions are bigger than their original atoms, because of the extra negative charge making the electrons repel further

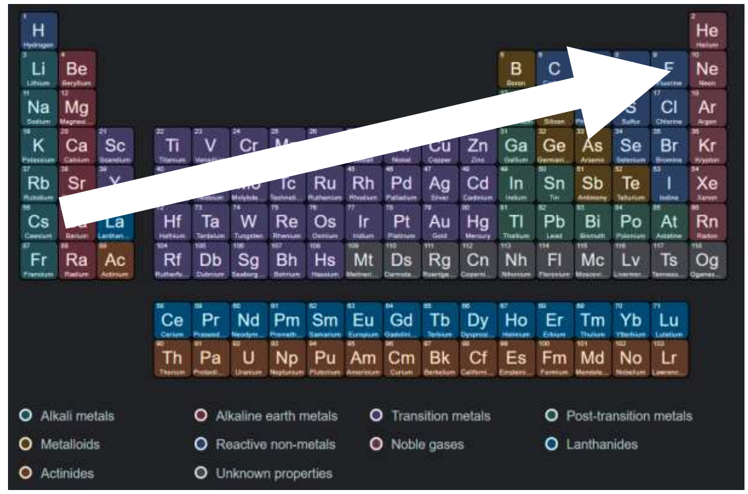
Ionization energy or Ionization potential

Ionization is the process of atoms becoming electrons

ionization energy It’s the energy required to remove a valence electron from the atom

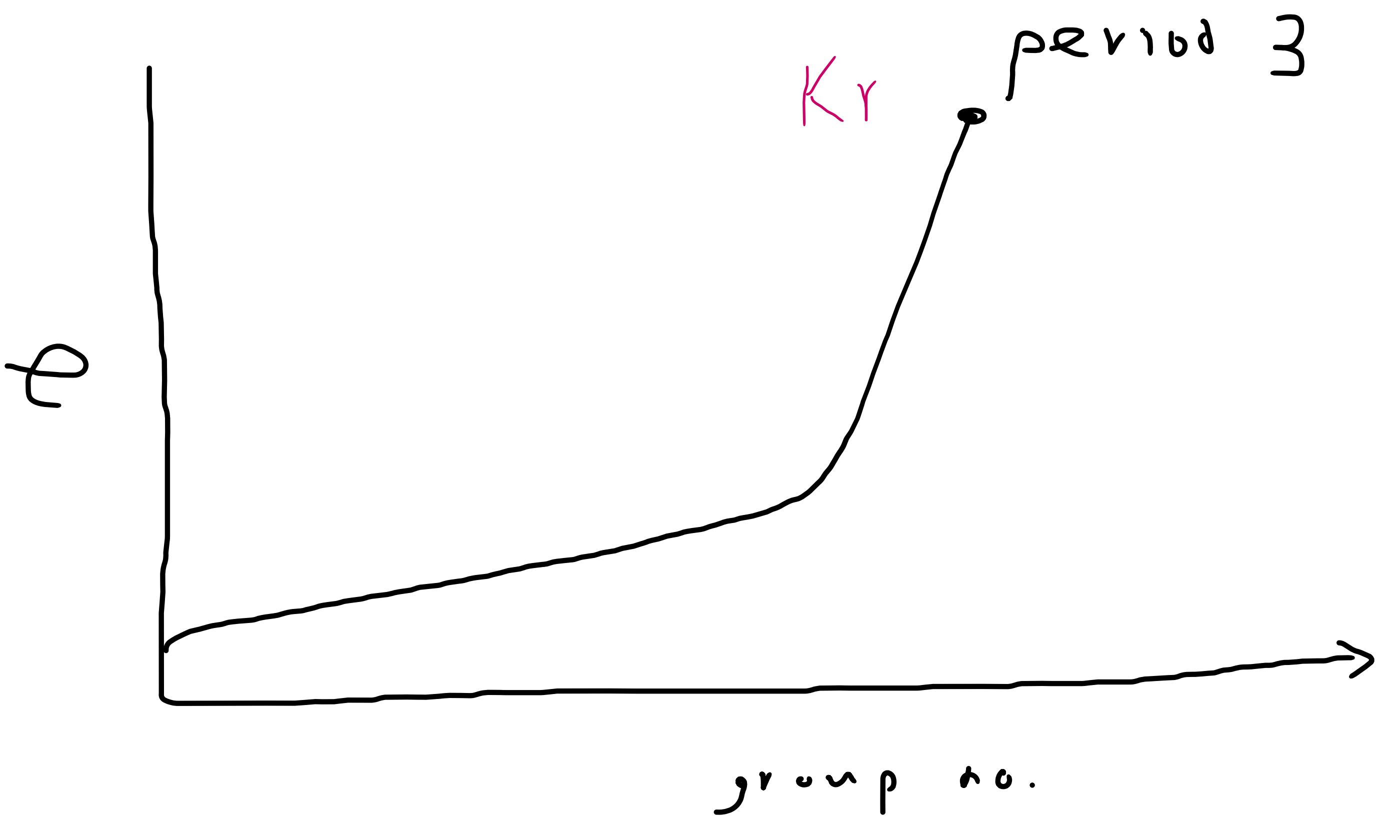
Aka how much the nucleus is holding on to a valence electron
it goes against the atomic radius, because when the atomic radius increases, the distance from the atom to the nucleus increases, and the electromagnetic pull decreases, so the ionization energy decreases

Disregarding the noble gasses, fluorine is the largest in ionization energy and cesium is the least, not francium because it decays quickly

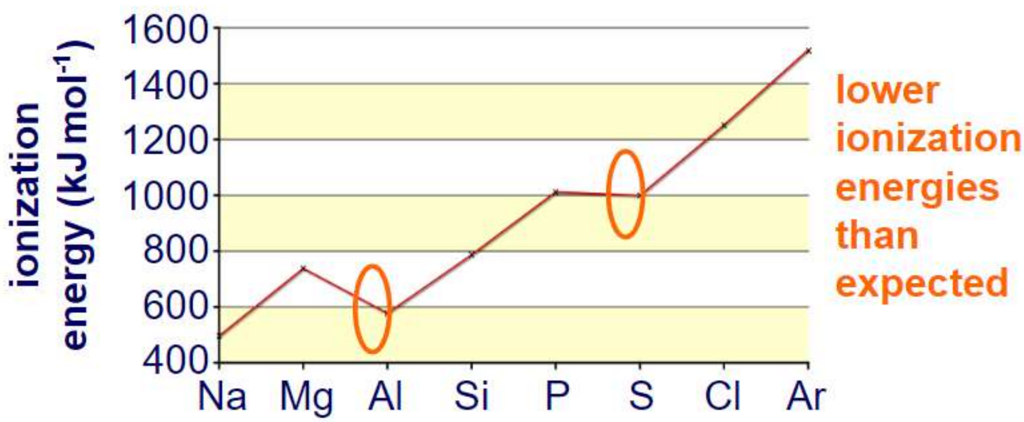


When an element loses more than a valence electron

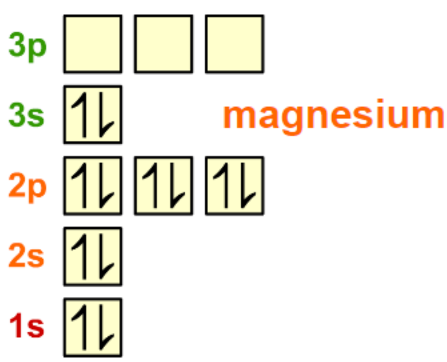
The ionization energy is of the first valence electron, is less than the second electron, until all the valence electrons are lost and you go to the next energy level, there will be a spike in ionization energy, because this atom is stable with electron filling its outermost energy level (noble gass), so it will spend more energy to make it stay

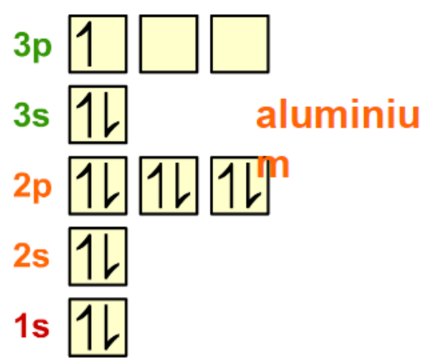


Exceptions in ionization energy



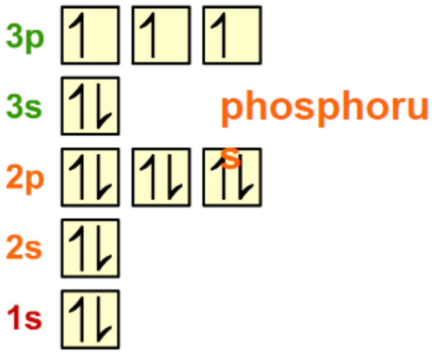
Ionization energy of aluminum is less than magnesium's because the electron removed from aluminum is in 3p while magnesium is 3s, and removing an electron from a higher energy orbital requires less energy

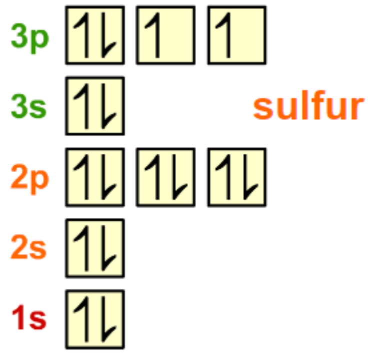




Ionization energy of sulfur is less than that of phosphorus, even though it has a higher positive pull and they are on the same sub-level (3p)

Because an electron in 3p of sulfur is paired, while electrons in 3p in phosphorus are singly occupied

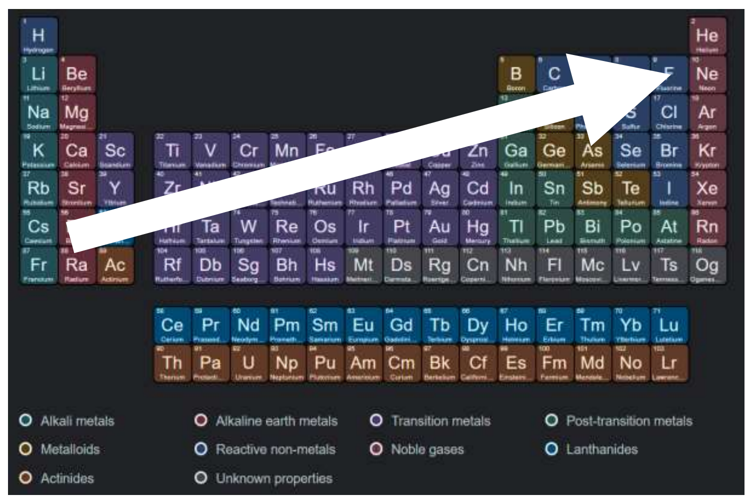




Mutual Repulsion between paired electrons in sulfur will mean less energy to remove one of them because they do be kicking each other

Electron affinity

It's the energy released when an atom gains an electron, it tells us how much an atom wants to gain an electron



- Noble gasses have an electron affinity of 0 because they are stable
- Electronic affinity increases from left to right, because the increase in non-metallic properties
- Disregarding noble gasses, the biggest group in electron affinity is 7A, but the biggest element in electron affinity is chlorine and not fluorine, because chlorine is so small and there is extra pull from protons, the atoms are crammed together and will try to repel any other atom
- The lowest group in electron affinity is 2A, and not 1A, because 2A have special stability, they don't want to gain shit
- The lowest element in electron affinity is radium

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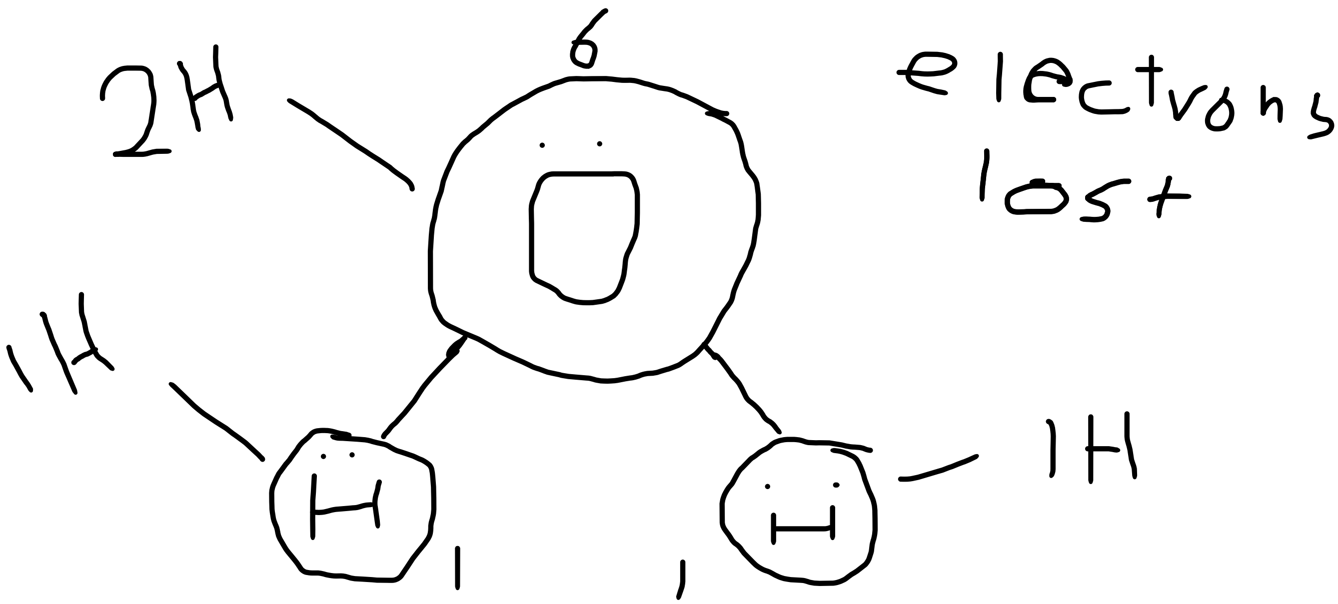
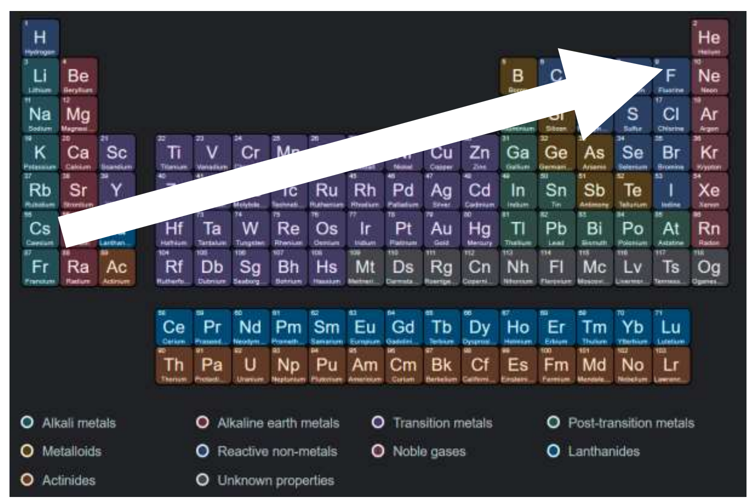
Electronegativity

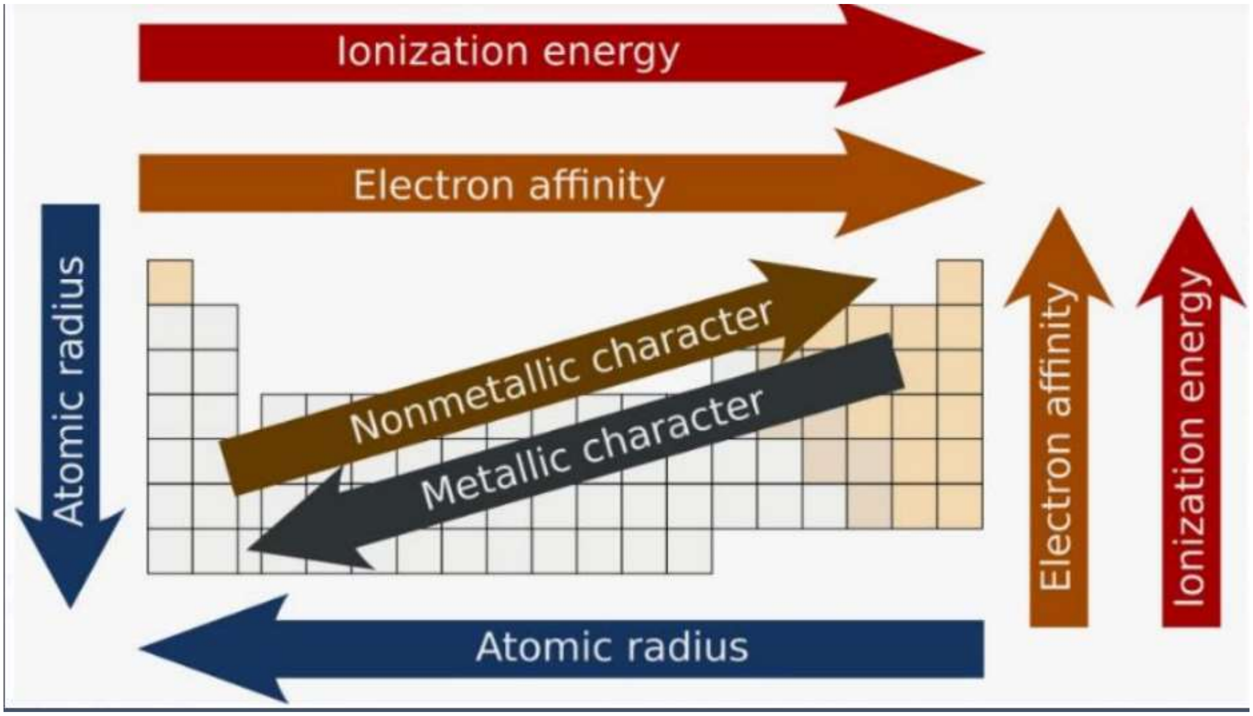
It’s the tendency of an atom to attract a shared electron closer to it from other atoms in a covalent bond

Covalent bond : is a bond where atoms share electrons and shit



- It increases from **left to right** because when the atomic number increases, the nuclear charge increases, and the electromagnetic pull increases
- It increases from **down to up** because when the energy level count decreases, they are closer to the nucleus, so the nuclear charge increases
- Disregarding noble gasses, highest element with electron negativity is fluorine in 7A
- The least element with electronegativity is Francium 1A





Disregarding noble gasses

Because they are stable atoms with their outermost energy level filled

1. Electronegativity of 0, they don't react
2. Electron affinity of 0, they don't want no more electrons
3. Highest ionization energy, they don't want to lose any electrons
4. Not very reactive

What makes an atom reactive?

- Incomplete valence electron level