

# 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



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## 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 elements 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



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## 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

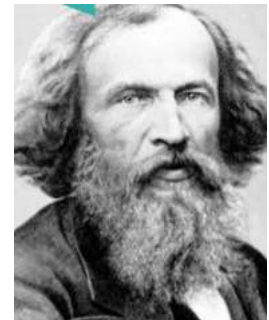
## Dmitri Mendeleev (1869)

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			
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	VIII		
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				



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## 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	II	III	IV	V	VI	VII	VIII
	a b	a b	a b	a b	a b	a b	a b	
	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	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26, Co 27, Ni 28
Kr 36	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Ru 44, Rh 45, Pd 46
	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	
Xe 54	Cs 55	Ba 56	Sr 38	Hf 72	Ta 73	W 74	Re 75	Os 76, Ir 77, Pt 78
	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84		
Rn 86	-	Ra 88	Ac 89	Th 90	Pa 91	U 92		



## The Modern Periodic Table

Legend:

- Alkali metals
- Alkaline earth metals
- Transition metals
- Post-transition metals
- Metalloids
- Reactive non-metals
- Noble gases
- Lanthanides
- Actinides
- Unknown properties



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We classify the periodic table in 3 ways

## Groups | Columns (18)

1a 2a 3a 4a 5a 6a 7a 8a - zero

The periodic table is organized into 18 groups, labeled 1a through 8a-zero. The elements are color-coded according to their properties:

- Alkali metals (Blue)
- Alkaline earth metals (Red)
- Transition metals (Purple)
- Post-transition metals (Green)
- Metalloids (Orange)
- Reactive non-metals (Light blue)
- Noble gases (Pink)
- Lanthanides (Dark blue)
- Actinides (Dark red)
- Unknown properties (Grey)

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

## Periods | Rows (7)

Legend:

- Alkali metals
- Alkaline earth metals
- Transition metals
- Post-transition metals
- Metalloids
- Reactive non-metals
- Noble gases
- Lanthanides
- Actinides
- Unknown properties

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



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## S Blocks (4)

1 H Hydrogen																	2 He Helium	
3 Li Lithium	4 Be Beryllium																	10 Ne Neon
11 Na Sodium	12 Mg Magnesium																	18 Ar Argon
19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton	
37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon	
55 Cs Cesium	56 Ba Barium	57 La Lanthanum	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Re Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon	
87 Fr Francium	88 Ra Radium	89 Ac Actinium	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	
		58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium			
		90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium			

Alkali metals

Metalloids

Actinides

Alkaline earth metals

Reactive non-metals

Unknown properties

Transition metals

Noble gases

Post-transition metals

Lanthanides

They are classified by their last filled sublevel  
 so if you're something like Na  $\rightarrow 1s^2 2s^2 2p^6 3s^1 \rightarrow$  your last sublevel is s  $\rightarrow$  you're in s block

S  $\rightarrow$  1A, 2A

P  $\rightarrow$  3A, 4A, 5A, 6A, 7A, 8A/zero

D  $\rightarrow$  3:12 "transitional elements"

F  $\rightarrow$  contains all elements under the table

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



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# 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 conductors
- 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 conductors of heat and electricity)

## Non-Metals

Are hydrogen, carbon, nitrogen, phosphorus, oxygen, Sulphur, selenium, fluorine, chlorine, iodine and astatine

- They are brittle and poor conductors



6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne
14 Si Silicon	15 P Phosphorus	16 S Sulfur	17 Cl Chlorine	18 Ar
32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr
50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe
82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn



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The melting points of **sulphur 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

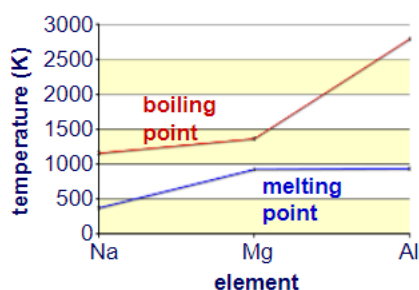
As for **corrosion**, a metal's resistance to corrosion is related to its **reactivity series**, **gold** is an unreactive materials and does not corrode easily



**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 metallic bonds, because more energy is needed to break them



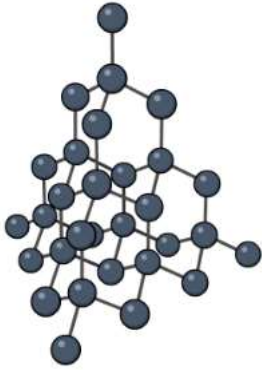
This metallic bond strength comes from

- **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 in 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 element in the universe and nearly a fifth of the human body is **carbon**

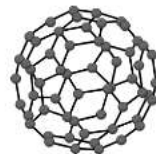
It can exist purely in 3 forms



Diamond, which is the hardest substance known.



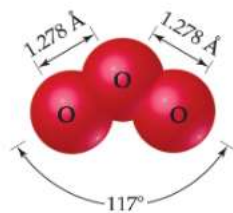
Graphite, which you will find in your pencil.



Fullerenes, which are football-shaped molecules.

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

## Oxygen

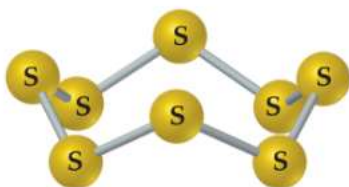


- There are two allotropes of oxygen:
  - $O_2$
  - $O_3$ , ozone
- There can be three anions:
  - $O^{2-}$ , oxide
  - $O_2^{2-}$ , peroxide
  - $O_2^{1-}$ , superoxide
- It tends to take electrons from other elements (oxidation).



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## Sulfur



- Sulfur is a weaker oxidizer than oxygen.
- The most stable allotrope is  $S_8$ , a ringed molecule.

## 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**



9	F	Fluorine
17	Cl	Chlorine
35	Br	Bromine
53	I	Iodine
85	At	Astatine

## 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

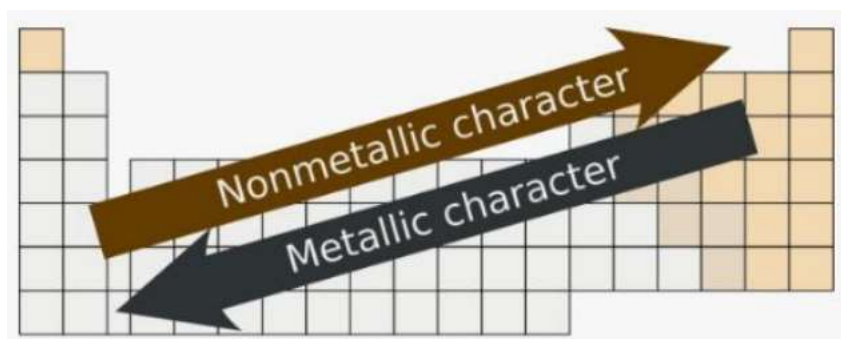
They are “boron, silicon, boron, germanium, arsenic, antimony and tellurium”

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



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

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





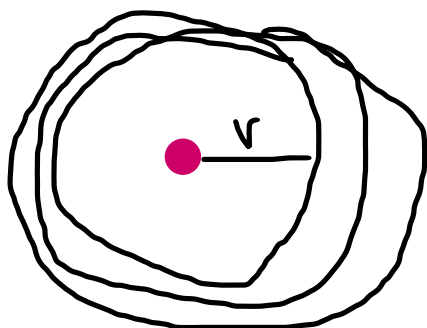
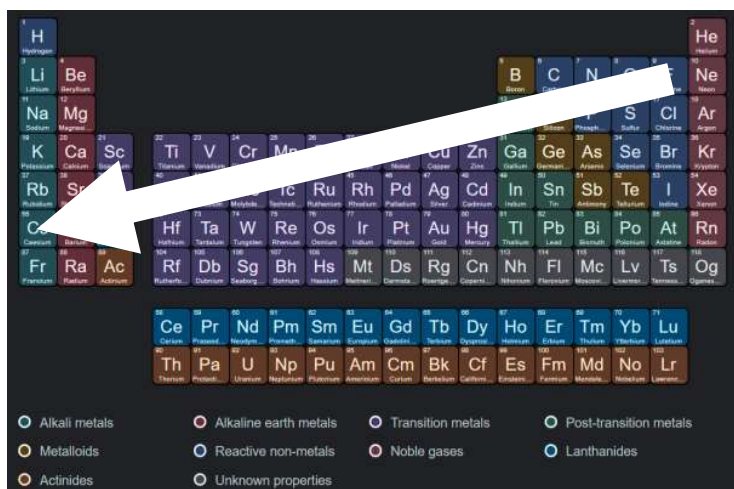
# Periodic Trends

**Periodicity** is a repeating pattern of properties across the periodic table

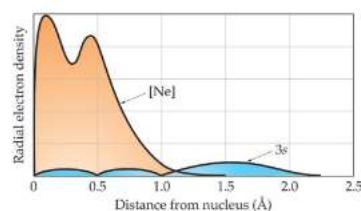
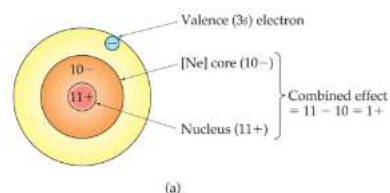
## Atomic Radius

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



the bigger the atom, the smaller it's  $Z_{\text{eff}}$  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 ( $Z_{\text{eff}}$ ) of an electron is determined by

- The atomic number  $Z$
- The screening constant  $S$

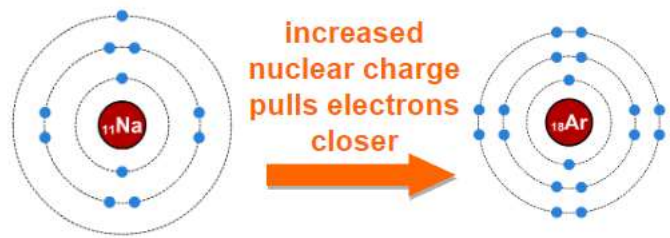
$$Z_{\text{eff}} = Z - S$$



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## LEFT TO RIGHT ( decreasing proton count )

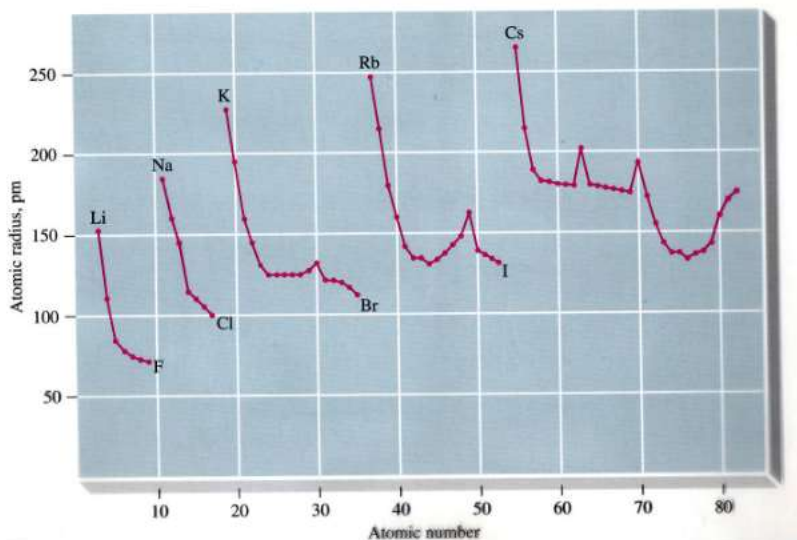
The atomic radius increases when going from the right to the left side of the screen, and that is because moving “right to left” means decreasing the atomic number, aka decreasing the nuclear charge pulling the electrons to the atom, leaving them have some **space**



## UP TO DOWN ( increasing energy levels )

The more energy levels, the more the radius, simple

Atomic Radii



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



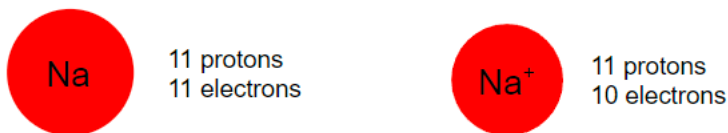
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## ionic Radius

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

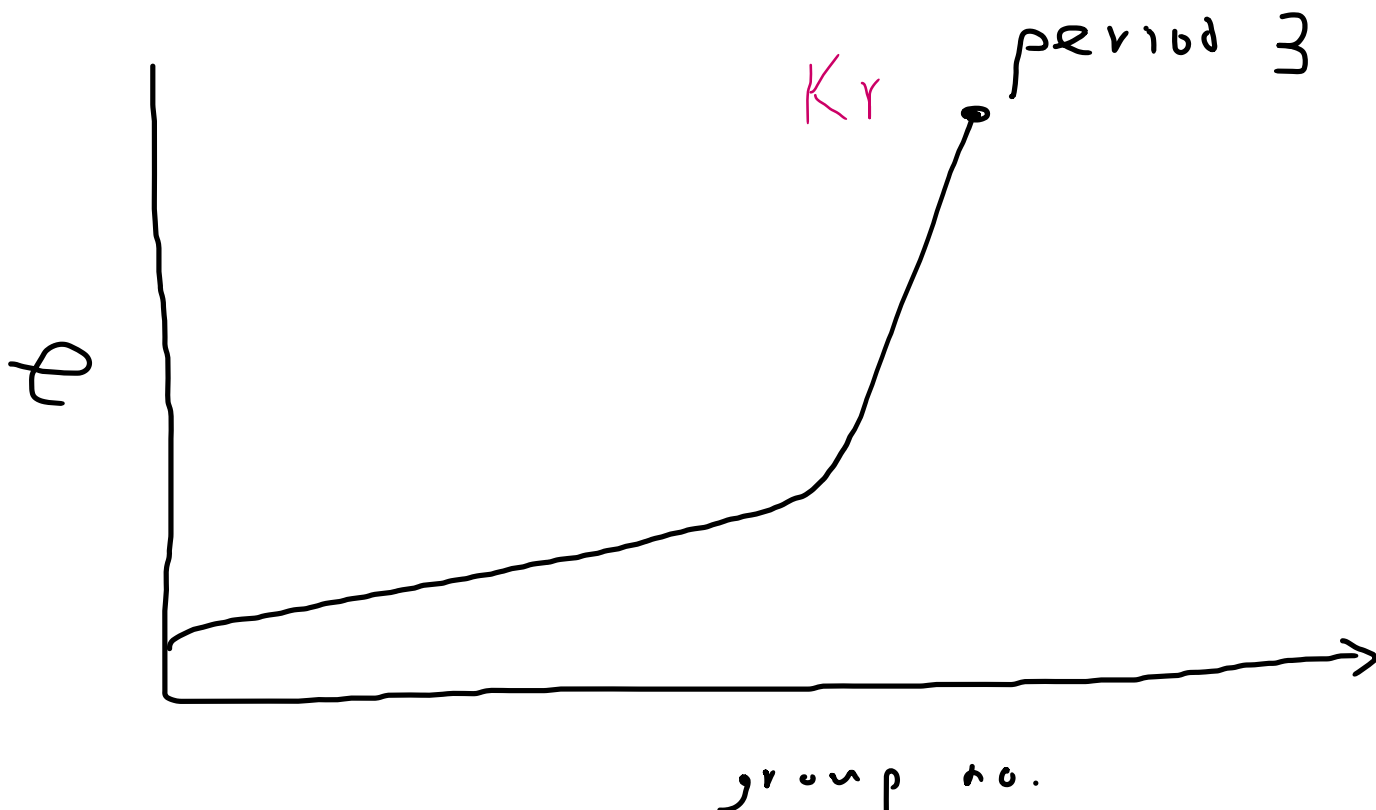


Disregarding the noble gasses, fluorine is the largest in ionization energy and cesium is the least, not francium because it decays quickly

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 gas), so it will spend more energy to make it stay

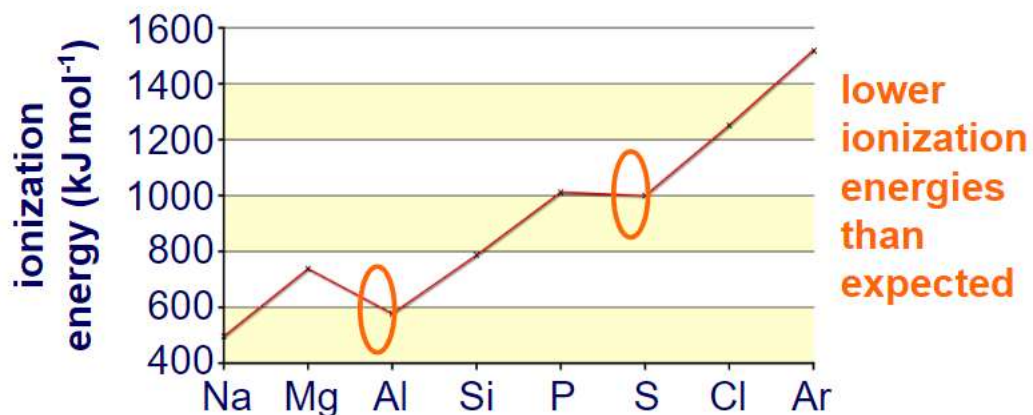
Periodic table of elements showing ionization energy trends. A white arrow points from the bottom-left (Francium) towards the top-right (Fluorine), indicating the direction of increasing ionization energy.



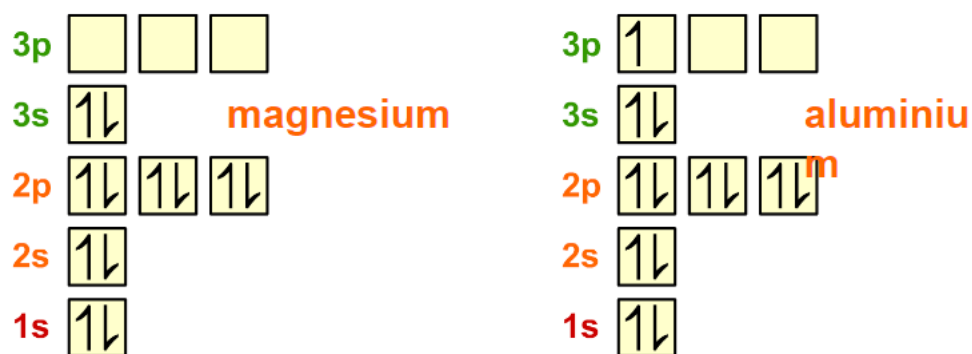
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## Exceptions in ionization energy



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

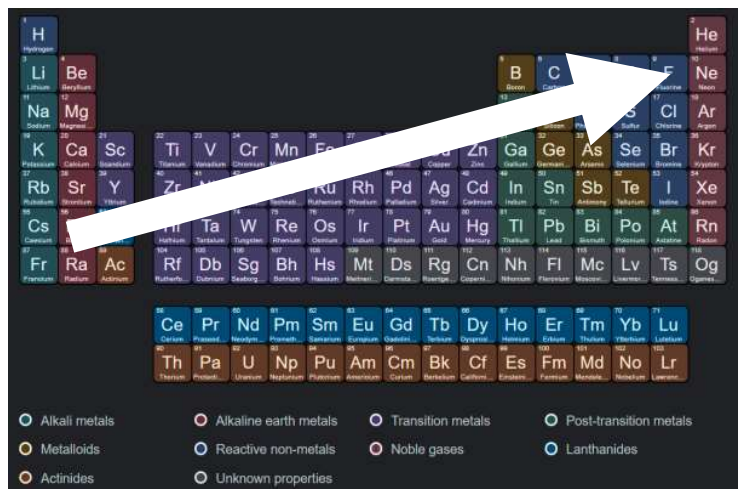


**Mutual Repulsion** between paired electrons in sulfur will mean less energy to remove one of them because they do be kicking each other

## 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 cramped 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

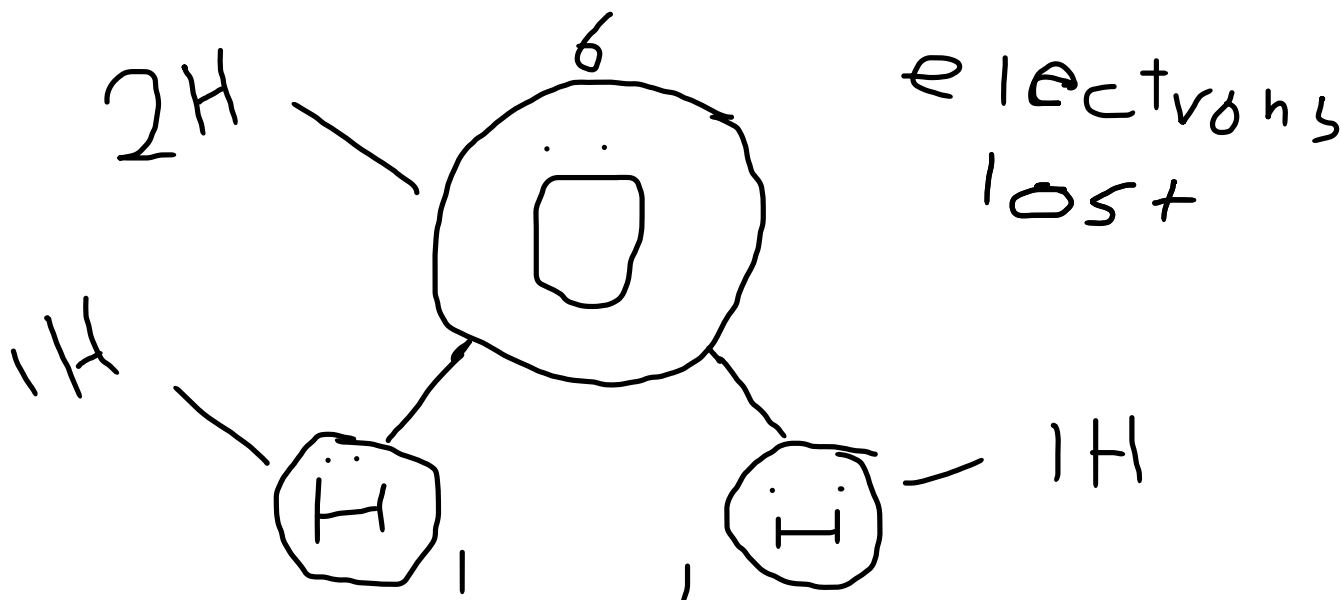


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## Electronegativity

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

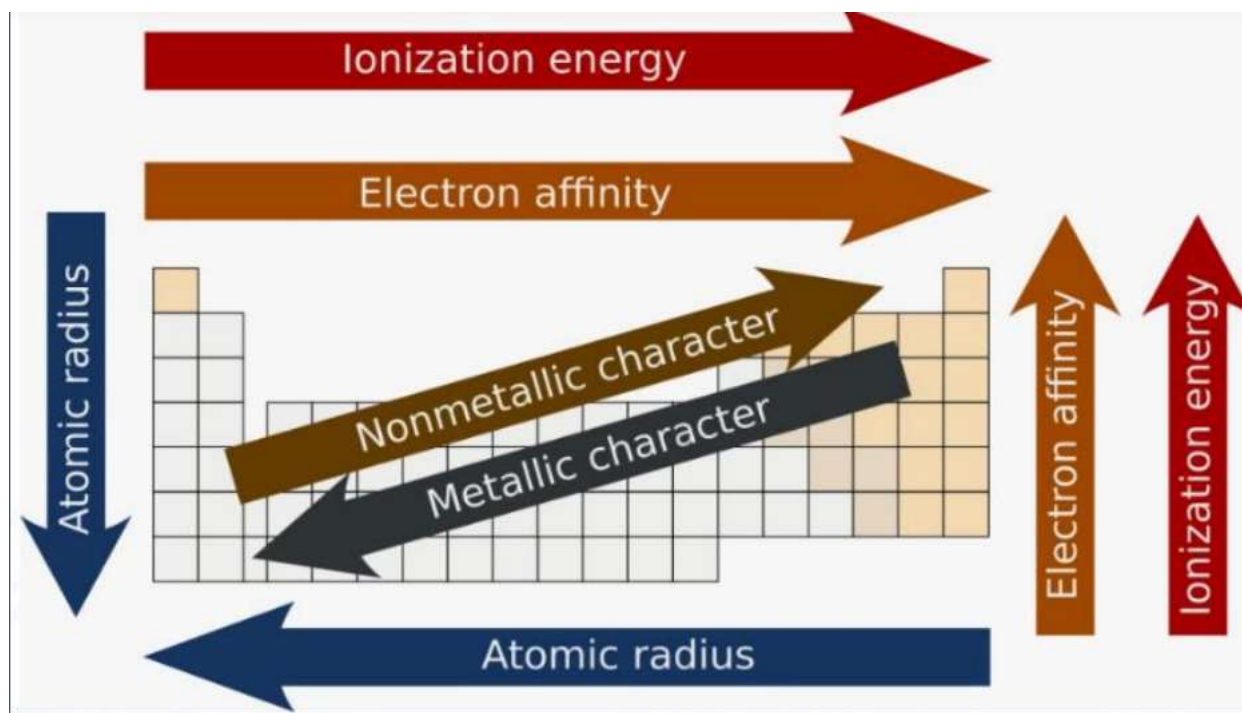
H																	He																															
Li	Be	B	C	N	O	F	Ne	Na	Mg	Al	Si	P	S	Cl	Ar	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Cu	Zn	Ga	Ge	As	Se	Br	Kr																	
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Cu	Zn	Ga	Ge	As	Se	Br	Kr	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe															
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	Fr	Ra	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr														
																																Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu			
																																Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr			

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### Disregarding noble gases

Because they are stable atoms with their outermost energy level filled

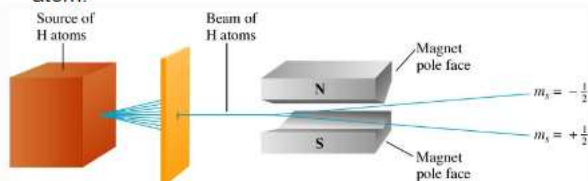
1. Electronegativity of 0, they don't react
2. Electron affinity of 0, they don't want no more electrons
3. Highest ionization energy, they don't want to lose any electrons
4. Not very reactive

### What makes an atom reactive?

- Incomplete valence electron level

Fig. 8.2 Stern-Gerlach Experiment

- Hydrogen atoms split into two beams when passed through magnetic field. Beams correspond to spin on atom.







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