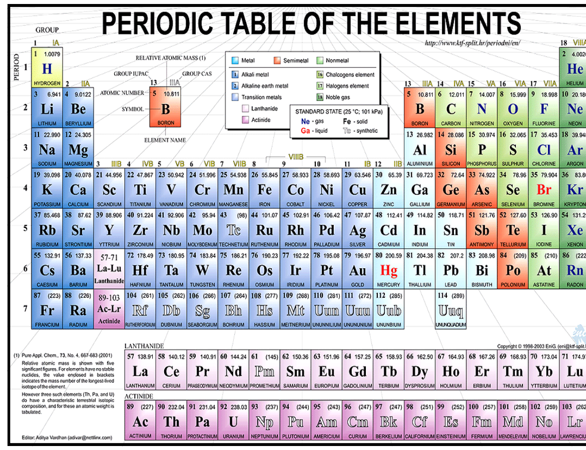
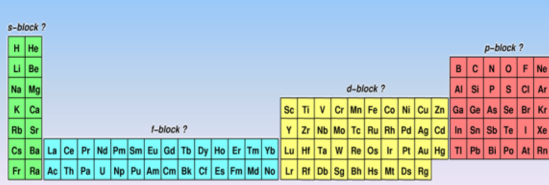
Periodic Table of Elements



Periodic Law

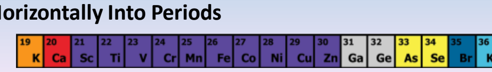
When the elements are listed in order of increasing <u>atomic number</u> similar physical and chemical properties recur periodically.



Elements are arranged

Vertically into Groups

Horizontally Into Periods



GROUP

Look at electronic configuration of element in each group what would you see?

4	Be
12	Mg
20	Ca
38	Sr
56	Ba
88	Ra

Each atom has the same number of electrons in it's outermost shell.

Be (Beryllium) Atom **Mg (Magnesium) Atom**

Note similarity in the configurations of the alkali metals

1 H Hydrogen 1.00794	$1s^1$
3 Li Lithium 6.941	$1s^2 2s^1$
11 Na Sodium 22.989770	$1s^2 2s^2 2p^6 3s^1$
19 K Potassium 39.0983	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
37 Rb Rubidium 85.4678	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$
55 Cs Cesium 132.90545	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$
87 Fr Francium (223)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^1$

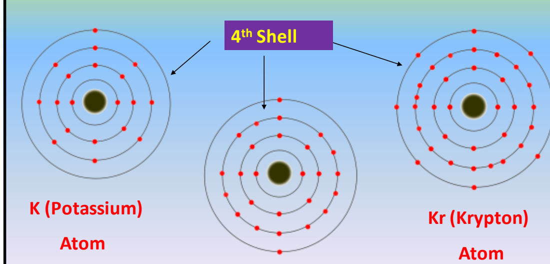
Note the similarity in the configurations of the noble gases

$1s^2$	2 He Helium 4.002602
$1s^2 2s^2 2p^6$	10 Ne Neon 20.1797
$1s^2 2s^2 2p^6 3s^2 3p^6$	18 Ar Argon 39.948
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$	36 Kr Krypton 83.798
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$	54 Xe Xenon 131.293
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$	86 Rn Radon (222)

PERIOD

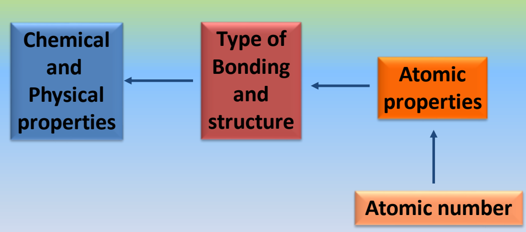
The period 4 atoms each have 4 electron containing shells

19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr



K (Potassium) Atom **Fe (Iron) Atom** **Kr (Krypton) Atom**

Number of electron in the outer most shell
Do most of Chemistry of Elements



Some Important Definitions

- Atomic Number
- Atomic Mass
- Valence electrons
- Shielding (screening) or Core electrons
- Nuclear Charge (Actual Nuclear Charge) (Z)
- Effective Nuclear Charge (Z_{eff})
- Shielding (Screening) effect

Periodic Trends of Elements

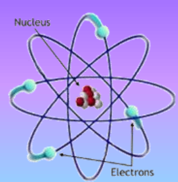
Periodic Trends

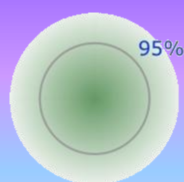
Describe *trends* among the elements for:

- Atomic size (radius),
- Ionic size (radius),
- Ionization energy,
- Electron affinity,
- Electronegativity,
- Metallic and non-metallic character.

Atomic Size

- The electron cloud doesn't have a definite edge, So





Atomic Size

Measure the Radius

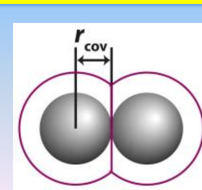
There are several types of radius:

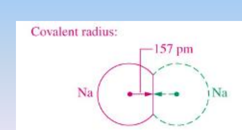
Types of Radii

- Covalent Radius
- Ionic Radius
- Metallic Radius
- Vander Waals radius

Covalent Radius

It is half of experimentally determined distance between two nuclei of similar atoms bonded together by single covalent bond.





Examples for Covalent Radius:

- **Two like atoms (A,A)**

  The experimentally determined distance of C atoms in diamond is 1.54 Å

  $r_c = 1.54/2 = 0.77$ Å

- **Two Different atoms (A,B)**

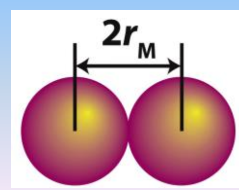
$d_{AB} = r_A + r_B$

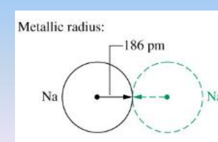
In ICl molecule (d_{I-Cl} = 2.32 Å, r_{Cl} = 0.99 Å)

$r_I = ?$

Metallic Radius

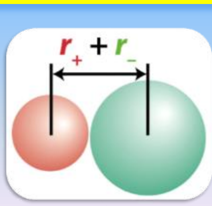
The half of the experimentally determined distance between the nuclei of two adjacent metal atoms in the crystalline solid material.

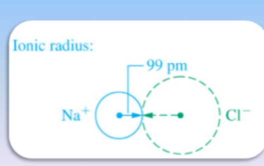




Ionic Radius

- The **IONIC RADIUS** **It is half distance between two nuclei of two atoms that bonded together by an ionic bond.**





Periodic Trends of Size of atoms and Ions

ALL Periodic Table Trends

Influenced by three factors:

1. **Energy Level**
   - **Higher energy levels are further away from the nucleus.**

3. **Shielding effect**
   (blocking effect?)

2. **Charge on nucleus (# protons)**
   - **More charge pulls electrons in closer. (+) and (–) attract each other)**

What do they influence?

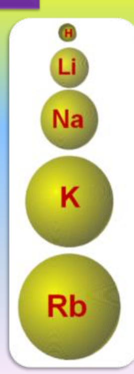
- **Energy levels** and **Shielding** have an effect on the ***GROUP***
- **Nuclear charge** has an effect on a ***PERIOD***

Atomic Size

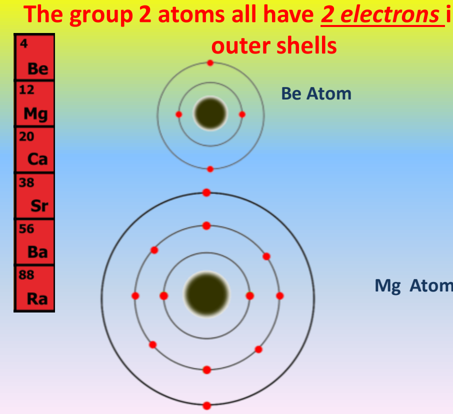
Group Trends (top to bottom)

Atomic radius increases by increasing the atomic number moving down the group.

1. Energy levels (number of shells) increase,
2. Shielding effect increase as a result of increasing the repulsion forces between electrons.



The group 2 atoms all have *2 electrons* in their outer shells

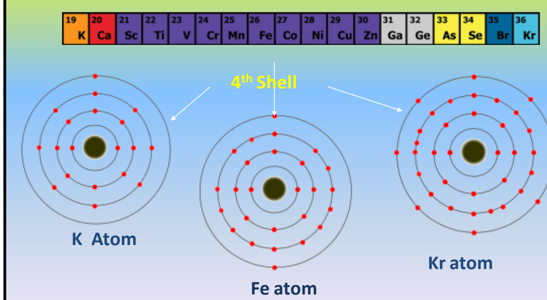


Trends in Atomic Size

- **Period Trends** (left to right)
- **The size gets smaller:**

1. Energy level (number of shells) remain constant.
2. Z increases (by one unit), electrons are added in the same energy level.
3. Nuclear charge increases; electrons are pulled close to the nucleus by the increased Z_{eff}

The **period 4** atoms each have **4 electron containing shells**



Ions

Cations: Positively charged particles.

Metals tend to LOSE electrons from their outer energy level

Sodium atom Na,

Sodium ion Na^{1+}

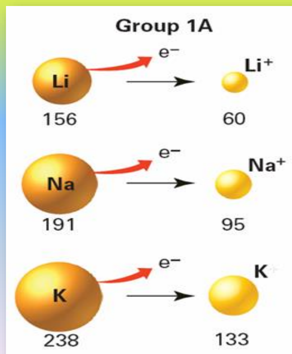
What about the size of Na and Na^{1+}?

Trends in Ionic Size

Cations

- **Cations are smaller than the parent atom – not only do they lose electrons, they lose an *entire energy level.***
- **Cations of representative elements have the noble gas configuration *before* them.**

Cations



Ions

Anions: Negatively charged particles

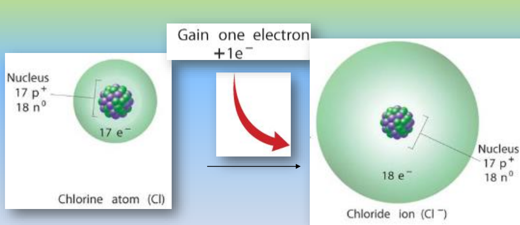
Nonmetals tend to GAIN one or more electrons

Chlorine atom (Cl)

Chloride ion (Cl^{1-})

What about the size?

Ions



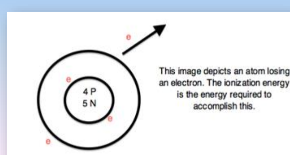
Ionization Energy (Potential)

Ionization Energy (IE)

Ionization Energy (or Ionization Potential)

"It is the amount of energy required to *completely remove an electron* (from a gaseous atom) or ion".

$M(g) \rightarrow M^{+}(g) + e^{-}$



2^{nd} and 3^{rd} Ionization Energy

First ionization energy:

The energy required to remove only the first electron.

Second ionization energy:

Is the energy required to remove the second electron.

Symbol	First	Second	Third
H	1312		
He	2731	5247	
Li	520	7297	11810
Be	900	1757	14840
B	800	2430	3569
C	1086	2352	4619
N	1402	2857	4577
O	1314	3391	5301
F	1681	3375	6045
Ne	2080	3963	6276

Why did these values increase so much?

Comment:

Ionization energy of hydrogen is greater than lithium?

1^{st} IE

H 1312 kJ /mol

Li 520 kJ /mol

1^{st} IE < 2^{nd} IE < 3^{rd} IE

$Mg(g) \rightarrow Mg^{+}(g) + e^{-}$ $I_1 = 738$ kJ/mol

$Mg^{+}(g) \rightarrow Mg^{2+}(g) + e^{-}$ $I_2 = 1451$ kJ/mol

Because it is more difficult to remove an electron from a positively charged particle than from neutral particle.

Variation along a period:
The ionization energy increases with increasing atomic number in a period.

Why?

Answer:
Due to the increased nuclear charge and decreasing in atomic size, the valence electrons are more tightly held by the nucleus. Therefore more energy is needed to remove the electron and hence ionization energy keeps increasing.

Variation down a group
The ionization energy gradually decreases in moving from top to bottom in a group.

Why?

Answer:

1. **Atomic Size (energy level), Shielding effect**
2. **Nuclear Charge**

On moving down a group

1. Nuclear charge increases
2. Number of shells increases, hence atomic size increases, screening is almost constant
3. There is a increase in the number of inner electrons which shield the valence electrons from the nucleus

Thus IE decreases down the group

On moving across a period

1. The atomic size decreases
2. Nuclear charge increases

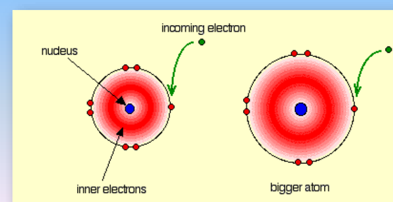
Thus IE increases along a period

Electron affinity

Electron affinity

It is the amount of energy released when an electron is added to an isolated gaseous atom.

$$\text{Atom}_{(g)} + \text{electron} \longrightarrow \text{Anion}_{(g)} + \text{energy}$$



$$\text{Atom}_{(g)} + \text{electron} \longrightarrow \text{Anion}_{(g)} + \text{energy}$$

Energy = Exothermic (heat released) = -ve value
= Endothermic (heat absorbed) = +ve value

If the atom has more tendency to accept an electron then the energy released will be large and consequently the electron affinity will be high.

The amount of energy released when an electron is added to an atom or ion.

Released 349 J/mol

Released 142 KJ/mol

Addition of the second electron is an endothermic process
Work must be done to force an electron to singly charged particle

Absorbed 710 KJ/mol

Electron affinities of the main-group elements

1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
H −72.8							He (0.0)
Li −59.6	Be (+18)	B −26.7	C −122	N +7	O −141	F −328	Ne (+29)
Na −52.9	Mg (+21)	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+35)
K −48.4	Ca (+186)	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)
Rb −46.9	Sr (+146)	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+41)
Cs −45.5	Ba (+46)	Tl −19.3	Pb −35.1	Bi −91.3	Po −183	At −270	Rn (+41)

Factors affecting electron affinity

- **When the nuclear charge is high there is greater attraction for the incoming electron. Therefore electron affinity increases as the nuclear charge increases.**

- **With the increase in the size of the atom the electron affinity decreases because the distance between the nucleus and the incoming electron increases.**

- **Electron affinities are low or almost zero in elements having stable electronic configurations (half filled and completely filled valence sub-shells) because of the small tendency to accept additional electron.**

Variation along a period

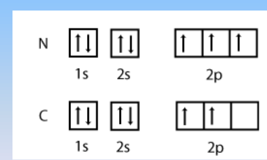
The size of an atom decreases and the nuclear charge increases on moving across a period. This results in greater attraction for the incoming electron. Hence the electron affinity increases in a period from left to right.

Variation down a group

As we move down a group the atomic size and nuclear size increases due to increasing of energy levels. Therefore, the additional electron feels less attracted by the large atom. Consequently the electron affinity decreases.

Carbon has a greater affinity for an electron than nitrogen.

Since a half-filled "p" subshell is more stable, C has a greater affinity to gain an additional electron.



Electronegativity

Electronegativity

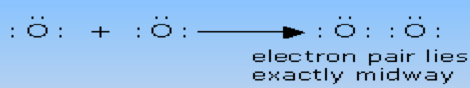
Electronegativity:

Ability of an element to attract electrons toward itself when bonded to another element.

An electronegative element attracts electrons.
An electropositive element releases electrons.

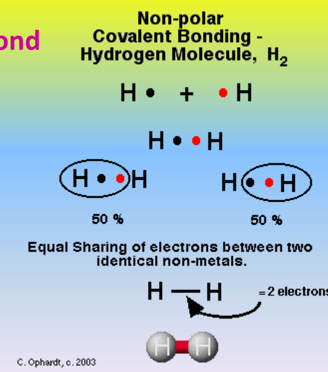
Atoms of the same electronegativity

E.g: H_2, O_2, Cl_2 molecules.



This type of covalent bond is described as ***a non-polar covalent bond.***

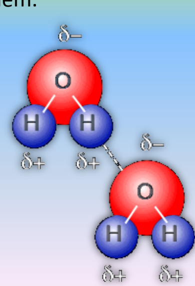
Non-polar covalent bond



Atoms of different electronegativity

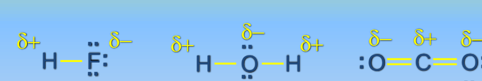
Polar covalent bond

Results when two **different non-metals unequally share** electrons between them.

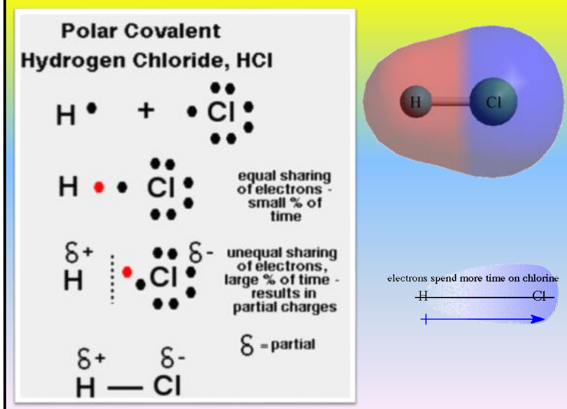


Atoms of different electronegativity

- **The greater the difference in electronegativity between two bonded atoms; the more polar the bond.**



polar bonds connect atoms of different electronegativity



The greater is the difference in the electronegativity values of the combining atoms, greater is the polar character in the bond so formed. For example, in the series H - X (X=F, Cl, Br, I), the electronegativity difference between H and X atom follows the order:

H- F > H - Cl > H - Br > H - I

Electronegativity difference:	(4-2.1)	(3.0-2.1)	(2.8-2.1)	(2.5-2.1)
	1.9	0.9	0.7	0.4

Therefore, the polarity in the H - X bond follows the order H-F > H - Cl > H - Br > H I
i.e., H-F bond is the most polar and H-I bond is the least polar in this series of compounds.

Percent Ionic Character of a polar Covalent Bond

It depends upon two factors:

1. Electronegativity difference of the bonded atoms
2. Dipole moment of the compound

1. Electronegativity Difference

ΔEN > 1.7 ionic bond - transfer

ΔEN < 1.7 covalent bond - sharing

So we have a range of electronegativity difference of 0 to 1.7 for sharing an electron pair.