*Chapter 8: Electron Configurations and Periodicity:

*8.1: Electron Spin and the Pauli Exclusion Principle:

*electron configuration: a particular distribution of electrons among the available subshells, which consists a group of orbitals with same (n, L).

*orbital diagram: a diagram to show how the orbitals of a subshell are occupied by electrons.

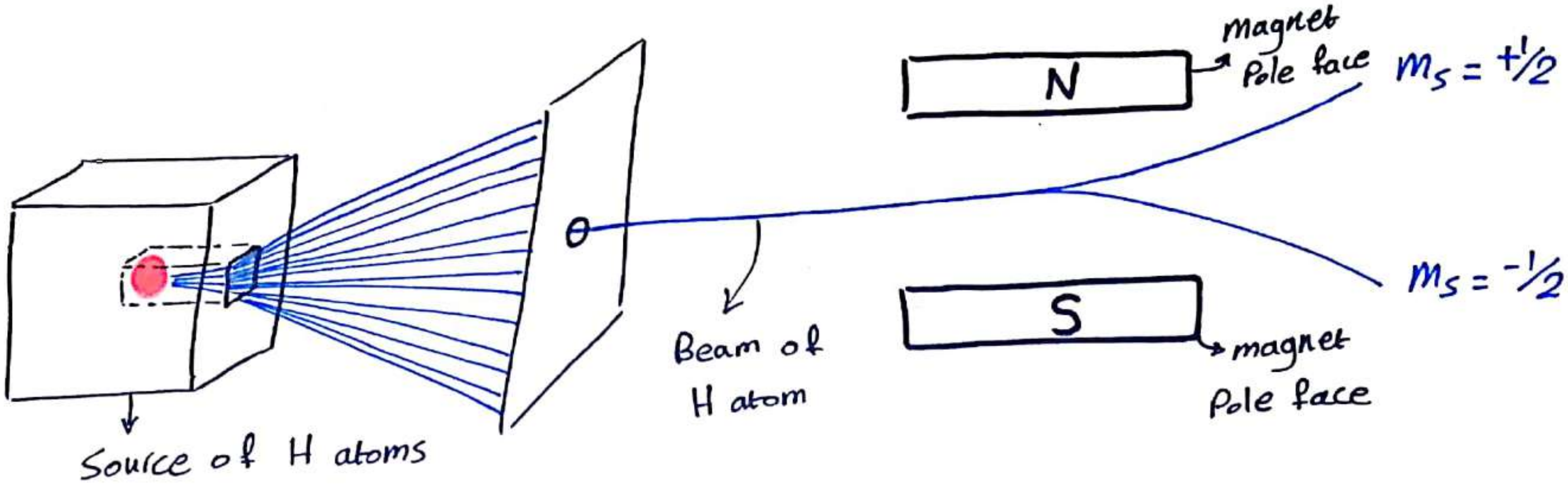
*Pauli exclusion principle: no two electrons in an atom can have the same four quantum numbers.

→ An orbital can hold at most two electrons, and then only if the electrons have opposite spins.

$m_s = +\frac{1}{2} \rightarrow \uparrow$

$m_s = -\frac{1}{2} \rightarrow \downarrow$

Subshell	Number of Orbitals	Maximum number of Electrons
s(L=0)	1	2
p(L=1)	3	6
d(L=2)	5	10
f(L=3)	7	14



* The beam of H atoms is split into two by the magnatic field.

* Half of the atoms are bent in one direction and half in the other.

* the atoms themselves act like magnets.

* **Example:** which of the following are possible and which are impossible:

a] ↑↓ (1s) ✓ ↑↓ (2s) ✓ ↑ _ _ (2p) ✓ — Possible

b] ↑↓ (1s) ✓ ↑↓↑ (2s) ✗ ↑ _ _ (2p) — impossible

c] ↑↓ (1s) ✓ ↑ (2s) ✓ ↑↑ [✗, ≈] _ _ (2p) — impossible

d] $1s^{3}$ ✗ $2s^{1}$ ✓ — impossible

e] $1s^{2} 2s^{1} 2p^{7}$ — impossible

f] $1s^{2} 2s^{2} 2p^{6} 3s^{2} 3p^{6} 3d^{8} 4s^{2}$ — Possible

*8.2: Bulding-Up Principle and the Periodic Table:

⊛ Building-Up Principle (Aufbau Principle)

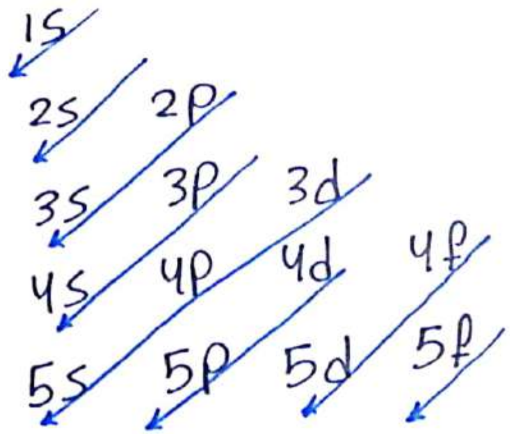
⇒ a scheme used to reproduce the electron configurations of the ground states of atoms by successively filling subshells with electrons in a specific order (the building-up order).

* lowest energy orbitals are filled first.

* following this principle, you obtain the electron configuration of an atom by successively filling subshells in the following order:

⇒ 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f

⊛ Electron Configurations and Periodic Table:



* noble-gas core: an inner-shell configuration corresponding to one of the noble gases.

* Valence electron: An electron in an atom outside the noble-gas or pseudo-noble-gas core

H $1s^1$

He $1s^2$

Li: $1s^2 2s^1 \rightarrow [He]2s^1$

B $1s^2 2s^2 2p^1 \rightarrow [He]2s^2 2p^1$

Ne $1s^2 2s^2 2p^6$

Mg $1s^2 2s^2 2p^6 3s^2 \rightarrow [Ne]3s^2$

Ar $1s^2 2s^2 2p^6 3s^2 3p^6$

Ca $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 \rightarrow [Ar]4s^2$

Cr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4 \rightarrow 4s^1 3d^5$

Cu: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9 \rightarrow 4s^1 3d^{10}$

* Exceptions to the Building-Up Principle:

→ chromium

(Cr)(Z=24): $[Ar]3d^4 4s^2$

$[Ar]3d^5 4s^1$

→ copper

(Cu)(Z=29): $[Ar]3d^9 4s^2$

$[Ar]3d^{10} 4s^1$

*8.3 writing Electron Configuration Using the Periodic Table:

* in the main group elements the valence configuration is:

$ns^a\ np^b$

n: period of element

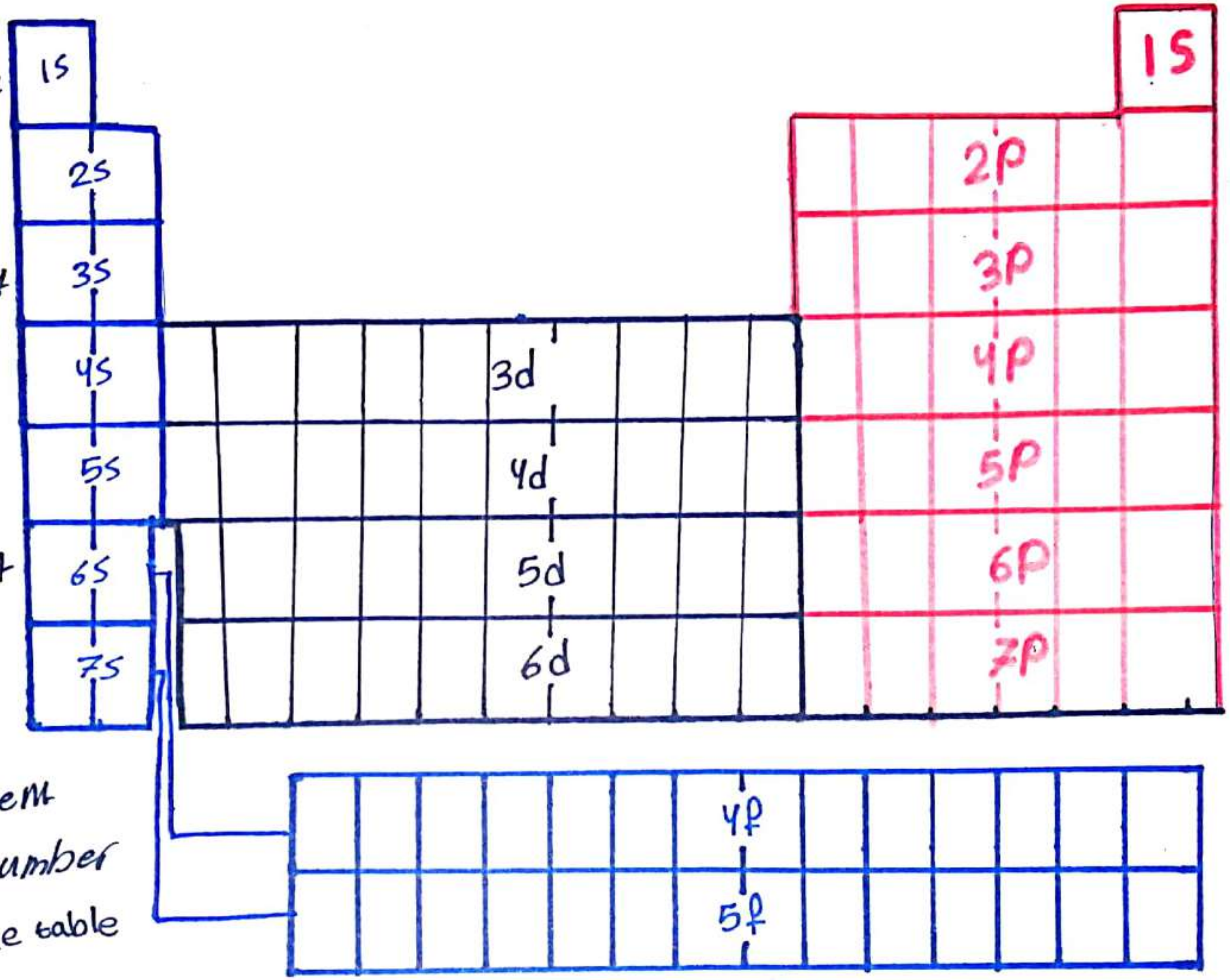
a+b: group number, and equals to valence electrons

* for transition element the outer shell configuration is:

$(n-1)d^{a-2}\ ns^2$

n: period of element

a: is the group number in B side of the table



*** Example:** Use the building-Up principle to obtain the configuration for the ground state of the gallium atom (Z=31). Give the configuration in complete form (do not appreviate for the core). what is the <u>valance-shell configuration?</u>

Z= 31 $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$

valence shell configuration: $4s^2\ 4p^1$ valence electrons = $3\bar{e}$

*** Example:** what are the configurations for the outer electrons of.

a. tellurium Z=52 : $[Kr]\ 5s^2\ 4d^{10}\ 5p^4 \longrightarrow 5s^2\ 5p^4$

b. nickel Z=28 : $[Ar]\ 4s^2\ 3d^8 \longrightarrow 4s^2\ 3d^8$

* 8.4: Orbital Diagrams of Atoms; Hund's Rule.

* Hund's rule: states that the lowest-energy arrangement of electrons in a subshell is obtained by putting electrons into separate orbitals of the subshell with the same spin before pairing electrons

↑↓ (1s) ↑↓ (2s) ↑↓ ↑ _ (2p) ✗

↑↓ (1s) ↑↓ (2s) ↑ ↑ ↑ (2p)

* **Example:** write an orbital diagram for the ground state of the iron atom?

$Fe_{26}: 1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^2\, 3d^6$

↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑↓ (2p) ↑↓ (3s) ↑↓ ↑↓ ↑↓ (3p) ↑↓ (4s) ↑↓ ↑ ↑ ↑ ↑ (3d)

* Magnatic Properties of Atom:

- Paramagnatic substance: a substance that is weakly attracted by a magnetic field, and this attraction is generally the result of unpaired electrons
  - ↳ at least one unpaired electron.

example:

O : $Z=8$: $1s^2\,2s^2\,2p^4$ ↑↓ (1s) ↑↓ (2s) ↑↓ ↑ ↑ (2p)

- diamagnatic substance: a substance that is not attracted by a magnetic field or is very slightly repelled by such a field, this property generally means that the substance has only paired electrons.
  - ↳ All electrons are paired.

example:

Ne : $Z=10$: $1s^2\,2s^2\,2p^6$ ↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑↓ (2p)

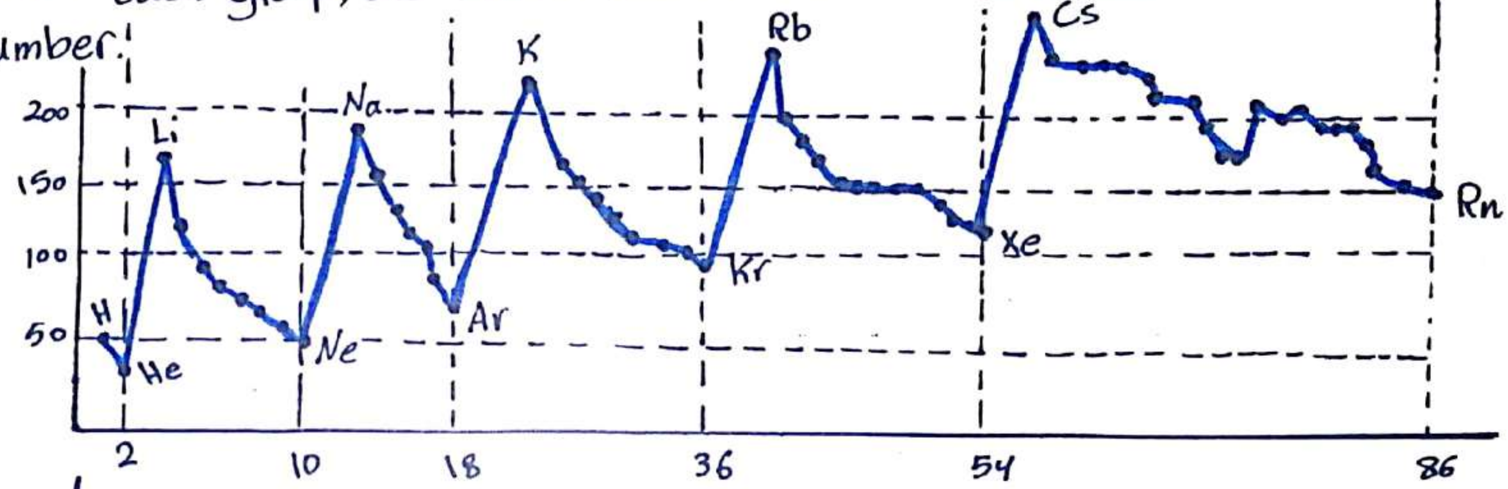
* 8.6: Some Periodic Properties.

* Periodic law: when the elements are arranged by atomic number, their physical and chemical properties vary periodically.

* Atomic ~~.....~~ Radius: the maximum in the radial distribution function of the outer shell of the atom.

→ within each period, the atomic radius tends to decrease with increasing atomic number

→ within each group, the atomic radius tends to increase with the period number.



Atomic Number (Z)

→ Cation (+) smaller than natural atom.

→ anion (-) larger than natural atom.

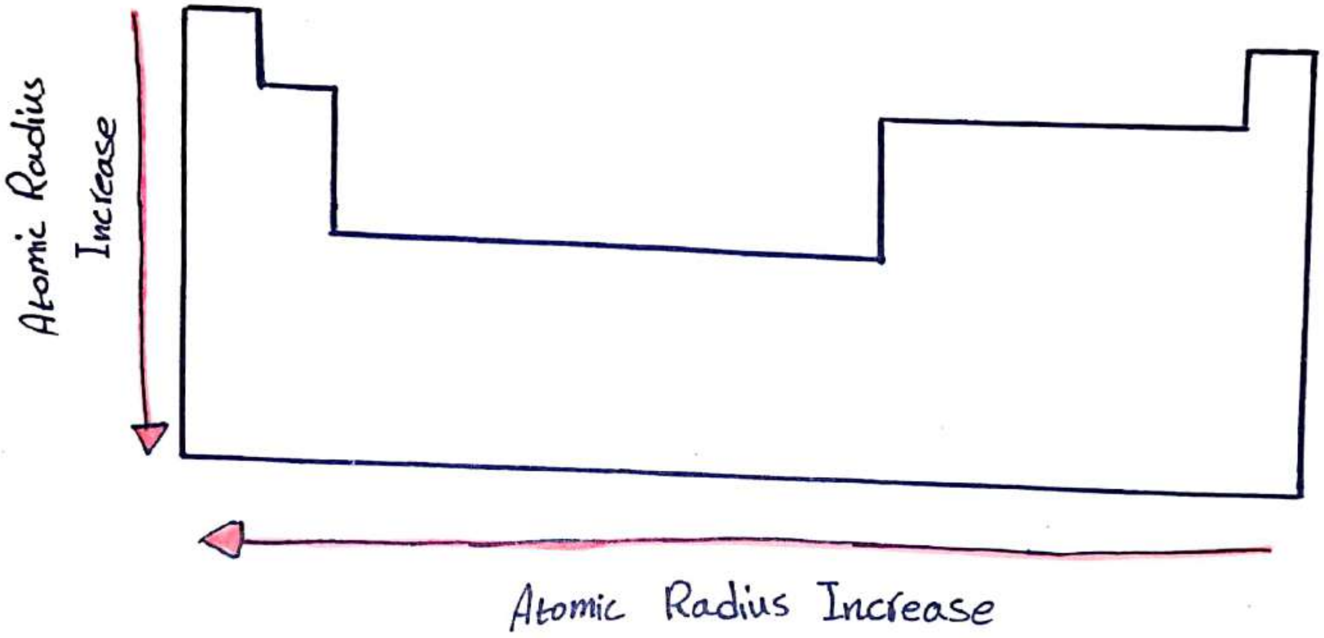
→ the more negative charge the larger. when they have the same no. of electrons.

${}_{7}N^{-3} > {}_{8}O^{-2} > {}_{8}O$ $Au^{+} > Au^{+3}$ $P^{-3} > N^{-3}$ ${}_{10}Ne > {}_{11}Na^{+} > {}_{13}Al^{+3}$

*** Example:** Arrange the following in order of increasing atomic radius:

(1) Al, C, Si   Al > Si > C

(2) Na, Be, Mg   Na > Mg > Be



*effective nuclear charge: the positive charge that an electron experiences from the nucleus, equal to the nuclear charge but reduced by any sheilding or screening from any intervening electron distribution.

* Ionization Energy :

→ first ionization energy: the minimum energy needed to remove the highest-energy (the outermost) electron from the natural atom in the gaseous state.

ionization energy for Lithium:

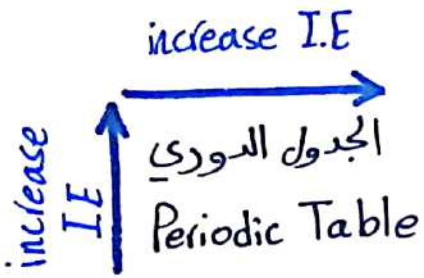
$$Li_{(g)}(1s^2 2s^1) \longrightarrow Li^+_{(g)}(1s^2) + \bar{e}$$

ionization energy = 520 KJ/mol

→ high values of I.E associated for noble gases

→ very low values of I.E associated for 1st group elements

$$\rightarrow 1^{st}\ IE < 2^{nd}\ IE < 3^{rd}\ IE$$



Exceptions

→ 3A element B has smaller I.E than the 2A element Be → Be: $1s^2 2s^2$ ⥮ ⥮ ; B: $1s^2 2s^2 2p^1$ ⥮ ⥮ ↿ _ _ → **Be > B**

→ same as Al and Mg → **Mg > Al**

→ 6A element O has smaller IE than N in 5A → **N > O**

→ same as S and P → **P > S**

→ for 2^{nd} ionization energy: it is better to arrange the electrons to decide.

→ He has the largest 1^{st} IE, Li has the largest 2^{nd} IE, Be has the largest 3^{rd} IE, ...

→ to decide the element in which group by his IE's 1^{st}, 2^{nd}, 3^{rd}, 4^{th}, ...
we subtract them by each other the largest difference decides that the element is in the smallest no. of IE between the two

1^{st} IE → A	B-A	which is bigger?
2^{nd} IE → B	C-B	
3^{rd} IE → C	D-C	
4^{th} IE → D		

suppose D-C is the larger, then it is in 3^{rd} group because C is th 3^{rd} IE.

***Example:** arrange the following in order of increasing of I.E: Ar, Se, S

Ar > S > Se

*** Example:** the 1^{st} IE for Cl is 1251 KJ/mole, which of the following values would be the more likely IE for I atom:

(a) 1000 KJ/mol   (b) 1400 KJ/mol

*** Example:** which of the following represent the 1^{st} IE of Al:

a] $Al^{-}_{(g)} \rightarrow Al_{(g)} + \bar{e}$ X

b] $Al_{(g)} \rightarrow Al^{-}_{(g)} + \bar{e}$ X

c] $Al_{(g)} \rightarrow Al^{+}_{(g)} + \bar{e}$

d] $Al^{+}_{(g)} + \bar{e} \rightarrow Al_{(g)}$

*** Example:** which of the following atom has the highest ionization energy?

a] $[Ne]3s^2 3p^2$ x

b] $[Ne]3s^2 3p^3$ y

c] $[Ar]3d^{10} 4s^2 4p^3$ z

d] $[Xe]4f^{14} 5d^{10} 6s^2 6p^3$ r

*** Example:** when trying to remove electrons from Be which of the following sets of IE makes the most sense going from first to third I.E ?

a] 1^{st} IE = 900, 2^{nd} IE = 1750, 3^{rd} IE = 15000 ✓

b] 1^{st} IE = 1750, 2^{nd} IE = 900, 3^{rd} IE = 15000 X

c] 1^{st} IE = 900, 2^{nd} IE = 15000, 3^{rd} IE = 22000 X

d] 1^{st} IE = 900, 2^{nd} IE = 1750, 3^{rd} IE = 1850 X

* electron Affinity: the energy required to remove an electron from the atom's negative ion (in its ground state).

الجدول الدوري
Periodic Table

$$Cl([Ne]3s^2 3p^5) + \bar{e} \longrightarrow Cl^-([Ne]3s^2 3p^6) \; ; \Delta E = -349 \text{ KJ/mol}$$

$$Cl^-([Ne]3s^2 3p^6) \longrightarrow Cl([Ne]3s^2 3p^5) + \bar{e} \; ; \Delta E = 349 \text{ KJ/mol}$$

***Exceptions:**

→ group 2A elements has lower EA than group 1A → 1A > 2A

→ group 5A elements has lower EA than group 4A → 4A > 5A

* group 8A (noble gases) have zero or smaller negative values.

* group 6 and 7 have the largest EA.

* Cl has the largest EA

* Exceptions:

→ in general we can say that the EA increases with the decreasing the no. of period, but it is not allways true.

الجدول الدوري

Periodic Table

example: based on that rule the largest EA is F but the truth is, it is Cl.

*Cl have the greatest EA.

Period	↑ 1A	* ↓ 2A	↑ 3A	↑ 4A	* 5A ↓	6A ↑	7A ↑	8A
1	H 73							He ≤0
2	Li 60	Be ≤0	B 27	C 122	N ≤0	O 141	F 328	Ne ≤0
3	Na 53	Mg ≤0	Al 44	Si 134	P 72	S 200	Cl 349	Ar ≤0
4	K 48	Ca 2	Ga 41	Ge 119	As 78	Se 195	Br 325	Kr ≤0
5	Rb 47	Sr 5	In 37	Sn 107	Sb 101	Te 190	I 295	Xe ≤0
6	Cs 46	Ba 14	Tl 36	Pb 35	Bi 91	Po 180	At 270	Rn ≤0