Development of Periodic Table

- Elements in the same group generally have similar chemical properties.

Physical properties are not necessarily similar, however.

Arrange elements to reflect the trends in chemical and physical properties.

•The periodic table arises from the periodic patterns in the **electronic configurations** of the elements.

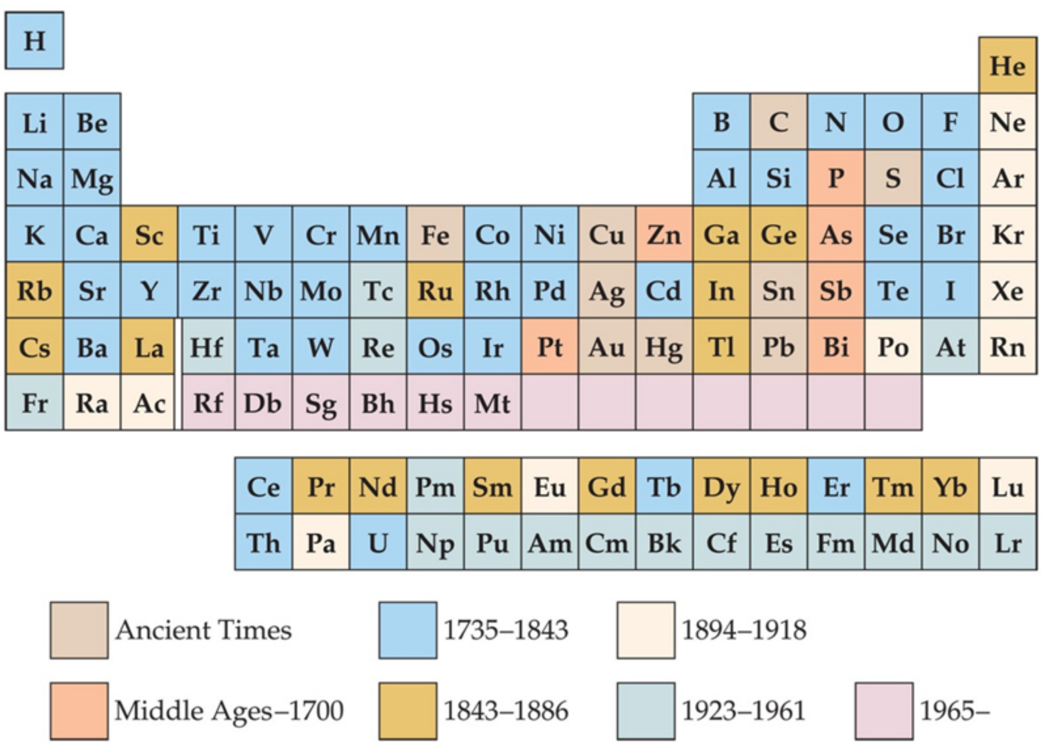
•Elements in the same column contain the same number of **valence electrons**.

•The trends within a row or column form patterns that help us make predictions about chemical properties and reactivity.

★ In the first attempt Mendeleev and Meyer arranged the elements in order of **increasing atomic weight**.

✓ Certain elements were missing from this scheme.

★ In the **modern periodic** table, elements are arranged in order of ***increasing atomic number***



- The International Union of Pure and Applied Chemistry (IUPAC) has just decided on some new and permanent names for the new elements reported in summer 2016
- **Moscovium** (mah-SKOH'-vee-um), symbol Mc, for element 115, and
- **Tennessine** (TEH'-neh-seen), symbol Ts, for element 117.

  ✓ The discovery team is from the Joint Institute for Nuclear Research in Dubna, Russia, the Oak Ridge National Laboratory and Vanderbilt University in Tennessee, and the Lawrence Livermore National Laboratory in California.

- **Nihonium** (nee-HOH'-nee-um), symbol Nh, for element 113.

  ✓ The element was discovered in Japan, and Nihon is one way to say the country's name in Japanese. It's the first element to be discovered in an Asian country.

- **Oganesson** (OH'-gah-NEH'-sun), symbol Og, for element 118.

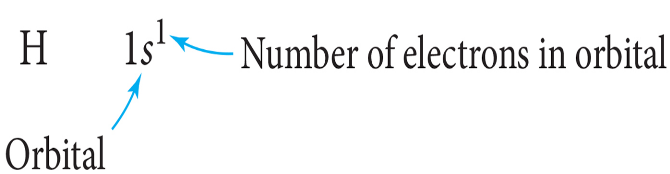
  ✓ The name honors Russian physicist Yuri Oganessian.

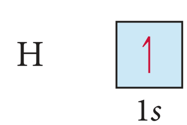
  ★ For more information visit

https://iupac.org/iupac-is-naming-the-four-new-elements-nihonium-moscovium-tennessine-and-oganesson/

Electron Configurations

- Quantum-mechanical theory describes the behavior of electrons in atoms.
- The electrons in atoms exist in orbitals.
- A description of the orbitals occupied by electrons is called an **electron configuration**.





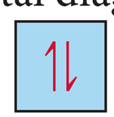
- **Orbital Diagrams**
  - ❖ **Definitions**
  - o <u>**Valence electrons**</u> are electrons in the outmost shell (highest energy level).

- They are the electrons included in the chemical reactions
- **# of valence electrons = Group #** (Main group elements)
- **Noble Gases** have filled shells (**8 valence e-s**)
- **Stable octet in valence shell (except H and He)**

- **We should keep in mind**

★ No two electrons in an atom may have the same set of four quantum numbers **(Pauli exclusion principle)**

★ No orbital may have more than two electrons, and they must have opposite spins.

Electron configuration		Orbital diagram
He	$1s^2$	 $1s$

Allowed Quantum Numbers:

n	*l*	m_l	m_s
1	0	0	$+\frac{1}{2}$
1	0	0	$-\frac{1}{2}$

Electron Configurations:

★ **Aufbau Principle** sometimes called the **building-up principle** :

★ Energy levels and sublevels fill from lowest energy to high:
$s \rightarrow p \rightarrow d \rightarrow f$

★ Orbitals that are in the same sublevel have the same energy.

★ No more than two electrons per orbital, which must have opposite spins (**Pauli exclusion principle**)

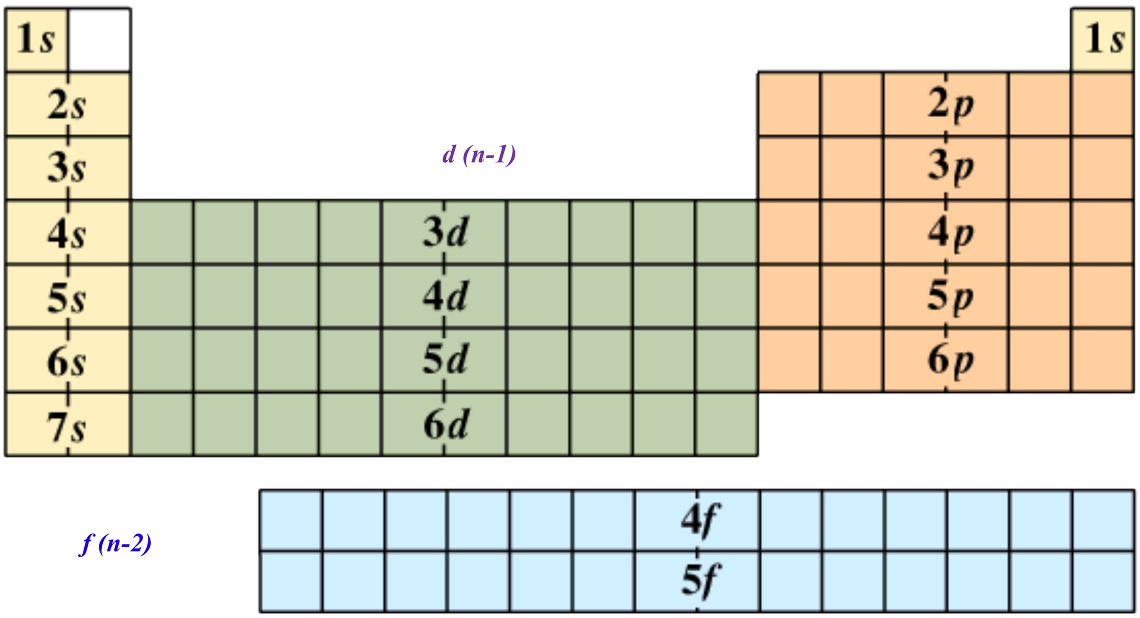
★ When filling orbitals that have the same energy, place one electron in each before completing pairs (**Hund's rule**)

Elelment	# of e	orbital diagram			e configuration
		1s	2s	$2p_x$ $2p_y$ $2p_z$	
H					
Li					
B					
F					
Ne					

Why this order for energy?

Electron Configuration of Atoms in Their Ground State:

All you need is a periodic table



Periodic Table

Main groups | Transition metals | Main groups

Period	1A[a] 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	1 **H** 1.00794																	2 **He** 4.00260
2	3 **Li** 6.941	4 **Be** 9.01218											5 **B** 10.811	6 **C** 12.0107	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.9984	10 **Ne** 20.1797
3	11 **Na** 22.9898	12 **Mg** 24.3050											13 **Al** 26.9815	14 **Si** 28.0855	15 **P** 30.9738	16 **S** 32.065	17 **Cl** 35.453	18 **Ar** 39.948
4	19 **K** 39.0983	20 **Ca** 40.078	21 **Sc** 44.9559	22 **Ti** 47.867	23 **V** 50.9415	24 **Cr** 51.9961	25 **Mn** 54.9380	26 **Fe** 55.845	27 **Co** 58.9332	28 **Ni** 58.6934	29 **Cu** 63.546	30 **Zn** 65.39	31 **Ga** 69.723	32 **Ge** 72.64	33 **As** 74.9216	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.80
5	37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.9059	40 **Zr** 91.224	41 **Nb** 92.9064	42 **Mo** 95.94	43 **Tc** [98]	44 **Ru** 101.07	45 **Rh** 102.9055	46 **Pd** 106.42	47 **Ag** 107.8682	48 **Cd** 112.411	49 **In** 114.818	50 **Sn** 118.710	51 **Sb** 121.760	52 **Te** 127.60	53 **I** 126.9045	54 **Xe** 131.293
6	55 **Cs** 132.9055	56 **Ba** 137.327	71 **Lu** 174.967	72 **Hf** 178.49	73 **Ta** 180.9479	74 **W** 183.84	75 **Re** 186.207	76 **Os** 190.23	77 **Ir** 192.217	78 **Pt** 195.078	79 **Au** 196.9666	80 **Hg** 200.59	81 **Tl** 204.3833	82 **Pb** 207.2	83 **Bi** 208.9804	84 **Po** [208.98]	85 **At** [209.99]	86 **Rn** [222.02]
7	87 **Fr** [223.02]	88 **Ra** [226.03]	103 **Lr** [262.11]	104 **Rf** [261.11]	105 **Db** [262.11]	106 **Sg** [266.12]	107 **Bh** [264.12]	108 **Hs** [269.13]	109 **Mt** [268.14]	110 **Ds** [271.15]	111 **Rg** [272.15]	112 **Cn** [277]	113 [284]	114 [289]	115 [288]	116 [292]		

Lanthanide series	57 **La** 138.9055	58 **Ce** 140.116	59 **Pr** 140.9077	60 **Nd** 144.24	61 **Pm** [145]	62 **Sm** 150.36	63 **Eu** 151.964	64 **Gd** 157.25	65 **Tb** 158.9253	66 **Dy** 162.50	67 **Ho** 164.9303	68 **Er** 167.259	69 **Tm** 168.9342	70 **Yb** 173.04
Actinide series	89 **Ac** [227.03]	90 **Th** 232.0381	91 **Pa** 231.0359	92 **U** 238.0289	93 **Np** [237.05]	94 **Pu** [244.06]	95 **Am** [243.06]	96 **Cm** [247.07]	97 **Bk** [247.07]	98 **Cf** [251.08]	99 **Es** [252.08]	100 **Fm** [257.10]	101 **Md** [258.10]	102 **No** [259.10]

[a]The labels on top (1A, 2A, etc.) are common American usage. The labels below these (1, 2, etc.) are those recommended

What is the electron configuration of S (Longhand Configuration)

Core Electrons
Valence Electrons

- **What is the electron configuration of S. Use Shorthand Configuration**

Electron configurations for **Transition Metals**

All 4th period elements have the configuration
[argon] ns^x $(n - 1)d^y$ and so are **d-block** elements

Stability related to lower energy this can be achieved by
___ _____ Orbital or _______orbital

Some exceptions across the transition metal region:

- Chromium

EXPECT:
ACTUALLY:

- Copper

EXPECT:
ACTUALLY:

Irregular Electron Configurations

Expected	Found experimentally
Cr = [Ar]$4s^2 3d^4$	Cr = [Ar]$4s^1 3d^5$
Cu = [Ar]$4s^2 3d^9$	Cu = [Ar]$4s^1 3d^{10}$
Mo = [Kr]$5s^2 4d^4$	Mo = [Kr]$5s^1 4d^5$
Ru = [Kr]$5s^2 4d^6$	Ru = [Kr]$5s^1 4d^7$
Pd = [Kr]$5s^2 4d^8$	Pd = [Kr]$5s^0 4d^{10}$

Nb, Rh and Ag also have the same irregularity

Paramagnetic: (unpaired electron) attracted to the magnetic field

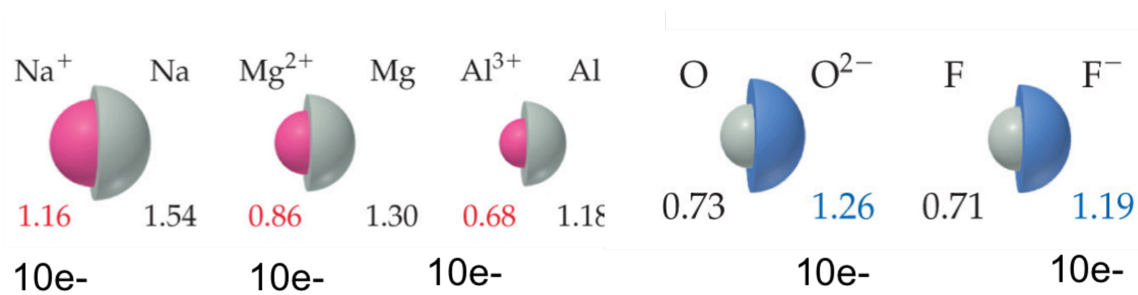
Diamagnetic: (paired electron) not attracted to the magnetic field

- **Ion Electron Configuration**

✓ **Ion Formation**

– Atoms gain or lose electrons to become more stable.

– Isoelectronic with the Noble Gases.



Group	Example
1A	Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$ Na^+: $1s^2\ 2s^2\ 2p^6$ = [Ne]
2A	Mg Mg^{+2}
3A	
5A	
6A	
7A	

Transition Metal Ions

Generally, **only highest energy e- lost** =(Outer s-subshell e-)

- Many transition metals form +2 cations

✓ **lose both s-subshell e-**

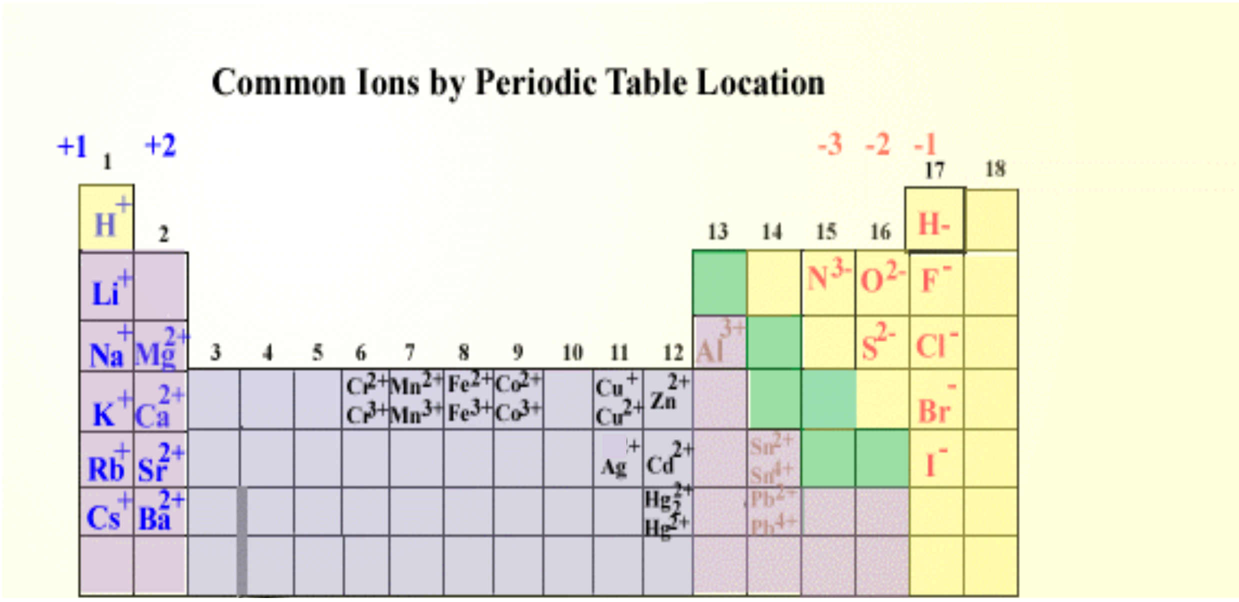
- For ions of higher charge

d-subshell e- are lost also

Fe [Ar] $4s^2\ 3d^6$

loses 2 electrons forms Fe^{2+} [Ar] $4s^0\ 3d^6$

Overall, common charges for ions are based on electron configuration

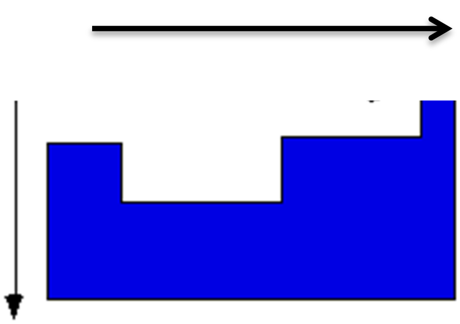


Need more Help on Electron Configuration Go to:
http://www.kentchemistry.com/links/AtomicStructure/Sublevels.htm

More help on understanding most of the topics covered in Chapters 6and 7
http://wps.prenhall.com/wps/media/objects/602/617465/AtomicRadii.html

Periodic Trends (see Examples 8.5-8.9 pages 356-368)

1. **Effective Nuclear Charge**



Shielding and Effective Nuclear Charge (Z_{eff}):

✓ Each electron in a multi-electron atom experiences both the attraction to the nucleus and the repulsion by other electrons in the atom.

- These repulsions cause the electron to have a net reduced attraction to the nucleus; it is **shielded** from the nucleus.
- The total amount of attraction that an electron feels for the nucleus is called the **effective nuclear charge (Z_{eff})** of the electron.

★ Z_{eff} is lower than the actual charge

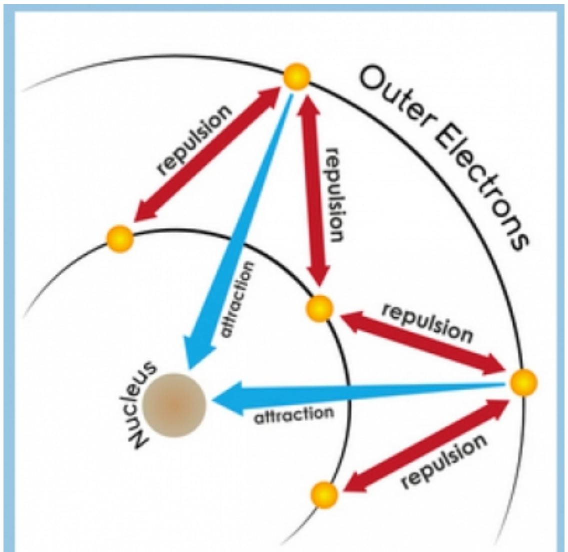


Figure 4-10. Effective Nuclear Charge
© Science Media Group.

$Z_{eff} = Z - S$

Z= Actual nuclear charge

S= Charge screened by other e-s

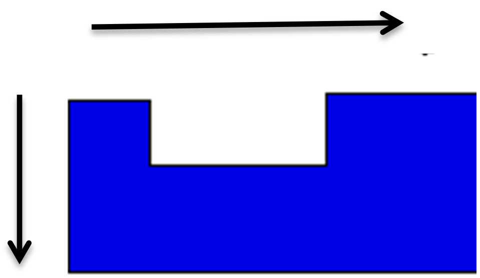
(charge of the core e-s)

Coulomb's Law: the potential energy (E) of two charged particles depends on their charges (q_1 and q_2) and on their separation (r)

$$E = \frac{1}{4\pi\varepsilon_0}\frac{q_1 q_2}{r}$$

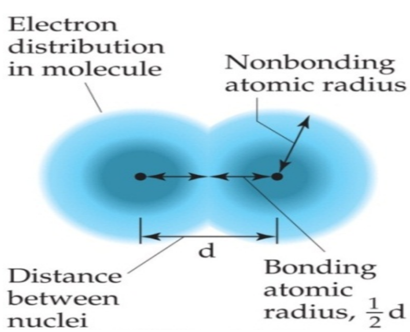
$\varepsilon_0 = 8.85 \times 10^{-12}\ C^2/J$

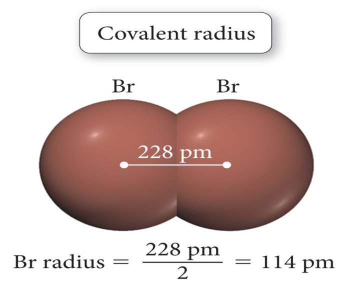
2. Trend in Atomic Radius: Main Group (see Example 8.5 page 356)



- Atomic radius decreases across period (left to right)
  - ✓ Adding electrons to same valence shell
  - ✓ Effective nuclear charge increases
  - ✓ Valence shell held closer

★ We actually cannot measure the size of the atom (no exact boundary for atoms or orbitals). Instead we use the bond length to estimate.





- ***Bonding Atomic Radius***= ½ bond length

3. Sizes of Ions

1) Cations

always smaller than parent atom

- ➢ The outermost electron is removed and repulsions between electrons are reduced.
- ➢ The effective nuclear charge has increased.

Cs^+ smaller than Cs

- ✓ Size decreases with increasing ionic charge

$Fe3^+ < Fe2^+$; $Cu^{2+} < Cu^+$

2) Anions

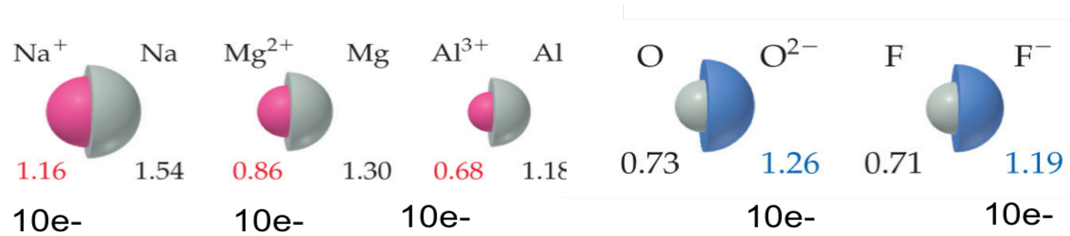
always larger than parent atom

- Electrons are added and repulsions between electrons are increased
- Trends in ion sizes are the same as atom sizes

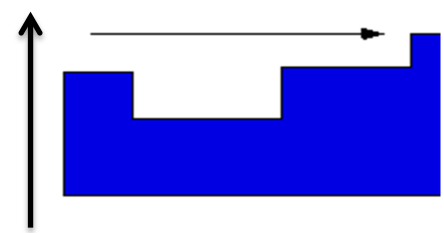
$Br^- > Br$ $S^{2-} > S$

- In an **isoelectronic series**, ions have the same number of electrons.
  - Ionic size decreases with an increasing nuclear charge.

    $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$



4. **Ionization Energy**: minimum energy required to remove an electron from **gaseous atom or ion**



- First ionization energy = energy to remove electron from gaseous neutral atom
  - Mg (g) ⟶ Mg^+ (g) +e- Ei_1 = 738 kJ/mol

- second ionization energy = energy to remove from gaseous ion, etc.
  - Mg^+ (g) ⟶ Mg^{2+} (g) +e- Ei_2 = 1450 kJ/mol

★ Energy cost depends on location in the periodic table.

Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	495	4562			(inner-shell electrons)		
Mg	738	1451	7733				
Al	578	1817	2745	11,577			
Si	786	1577	3232	4356	16,091		
P	1012	1907	2914	4964	6274	21,267	
S	1000	2252	3357	4556	7004	8496	27,107
Cl	1251	2298	3822	5159	6542	9362	11,018
Ar	1521	2666	3931	5771	7238	8781	11,995

There are some exceptions:

N: [Ne} $2s^22p^3$ ⟶ N^+: [Ne} $2s^22p^2$ Ei=1402KJ/mol

O: [Ne} $2s^22p^4$ ⟶ O^+: [Ne} $2s^22p^3$ Ei=1314KJ/mol

Be: [He} $2s^2$ ⟶ Be^+: [He} $2s^1$ Ei=899KJ/mol

B: [He} $2s^22p^1$ ⟶ B^+: [He} $2s^2$ Ei=801KJ/mol

❖ What about adding an electron?
Electron Affinity, EA: Energy associated with the gain of an electron by a **gaseous atom or ion.**

- Electron affinity is defined as exothermic (−) but may actually be endothermic (+).
  - Some alkali earth metals and all noble gases are endothermic. Why?
- The more energy that is released, the larger the electron affinity.
  - The more negative the number, the larger the EA.

Electron Affinities (kJ/mol)

1A	2A	3A	4A	5A	6A	7A	8A
H −73							**He** >0
Li −60	**Be** >0	**B** −27	**C** −122	**N** >0	**O** −141	**F** −328	**Ne** >0
Na −53	**Mg** >0	**Al** −43	**Si** −134	**P** −72	**S** −200	**Cl** −349	**Ar** >0
K −48	**Ca** −2	**Ga** −30	**Ge** −119	**As** −78	**Se** −195	**Br** −325	**Kr** >0
Rb −47	**Sr** −5	**In** −30	**Sn** −107	**Sb** −103	**Te** −190	**I** −295	**Xe** >0

- First **Electron Affinity**

$Cl(g) + e^- \longrightarrow Cl^-(g)$ EA = -349 kJ/mol

- Second **Electron Affinity**

$O^-(g) + e^- \longrightarrow O^{2-}(g)$ EA = +710 kJ/mol

- 2nd e- must be forced onto a negative charged ion which requires energy

Exceptions in Electron Affinity

Group 1A Vs Group 2A

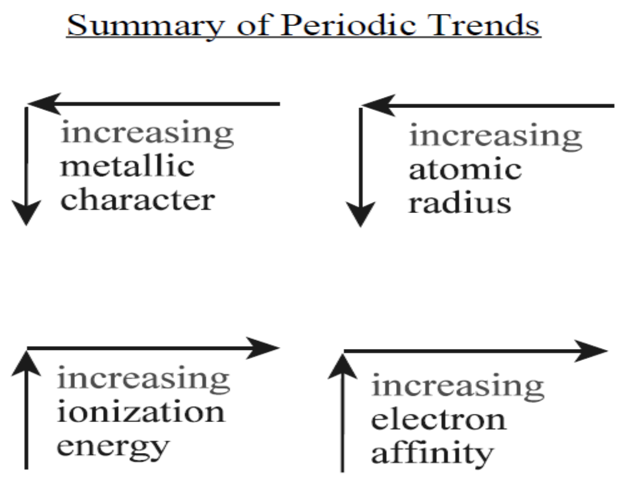
Na Mg

Na^- Mg^-

Group 4A Vs Group 5A

C N

C^- N^-



Periodic Chemical behavior

❖ **Group 1A**

- They are found only in compounds in nature, not in their elemental forms.
- Alkali metals are NEVER found pure in nature; they are too reactive.
- They have low densities and melting points.
- They also have low ionization energies.
- Their chemistry is dominated by the loss of their single *s* electron:
  $M \rightarrow M^+ + e^-$
- Reactivity increases as we move down the group.

❖ **Chemical Reactions of Group 1A elements:**

a) With Halogens:

$M + X_2 \rightarrow 2MX$

b) With Water

$M + H_2O \rightarrow MOH + H_2$

c) With Oxygen

$Li + O_2 \rightarrow Li_2O$

$Na + O_2 \rightarrow Na_2O_2$

$K + O_2 \rightarrow KO_2$

d) With Hydrogen

$Na + H_2 \rightarrow NaH$

❖ **Group 7A**

✓ Exist as diatomic

- Very high electron affinities
  - Good oxidizing agents; easy to reduce
  - Very reactive; not found uncombined in nature
  - React with metals to form salts
  - Compounds generally soluble in water ∴ found in seawater

★ They are very reactive , they react with almost every thing (except for noble gases)

Fluorine is one of the most reactive substances known:

$2F_2(g) + 2H_2O(l) \rightarrow 4HF(aq) + O_2(g)$ $\Delta H = -758.9$ kJ

❖ **Group 8A**

❖ Very unreactive

- Only found uncombined in nature
- Used as "inert" atmosphere when reactions with other gases would be undesirable.

- (In 1962) Xe forms three compounds:
  - XeF_2
  - XeF_4
  - XeF_6
- Kr forms only one stable compound:
  - KrF_2
- The unstable HArF was synthesized in 2000.