## Valence electron configuration determines the characteristics of elements in a group

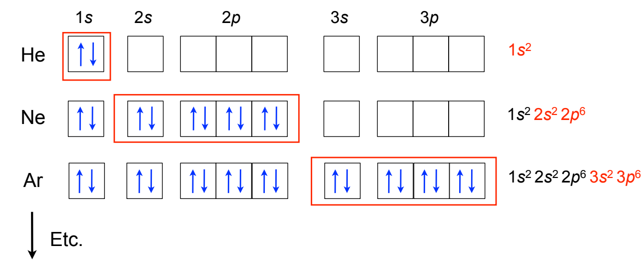
| IA | IIA | IIIA | IVA | VA | VIA | VIIA | Noble gases |
|---|---|---|---|---|---|---|---|
| 1 **H** $1s^1$ | | | | | | | 2 **He** $1s^2$ |
| 3 **Li** $2s^1$ | 4 **Be** $2s^2$ | 5 **B** $2s^22p^1$ | 6 **C** $2s^22p^2$ | 7 **N** $2s^22p^3$ | 8 **O** $2s^22p^4$ | 9 **F** $2s^22p^5$ | 10 **Ne** $2s^22p^6$ |
| 11 **Na** $3s^1$ | 12 **Mg** $3s^2$ | 13 **Al** $3s^23p^1$ | 14 **Si** $3s^23p^2$ | 15 **P** $3s^23p^3$ | 16 **S** $3s^23p^4$ | 17 **Cl** $3s^23p^5$ | 18 **Ar** $3s^23p^6$ |

The chemical behavior and properties of elements in a **group** are associated with the *valence electron configuration* of its elements

## Noble gas configuration

The noble gases (last column in the periodic table) are characterized by completely filled *s* and *p* orbitals

- this is a **very stable** valence electron configuration
- noble gases typically exist as single atoms and do not readily form compounds with other elements




## Atoms of other elements seek to attain a noble gas electron configuration

***Noble gases are highly unreactive and do not readily form compounds***

The rest of the elements in the periodic table tend to combine with other elements to form compounds

When forming compounds, the atoms of these elements ***lose, gain,*** or ***share*** electrons to attain a stable valence electron configuration (*i.e.,* identical to noble gases)

- Atoms that attain a noble gas configuration by losing or gaining electrons (*i.e.,* forming ions) form *ionic bonds*
- Atoms that attain a noble gas configuration by sharing electrons form *covalent bonds*

For most elements, attaining a **noble gas configuration** means having 8 **valence electrons** (two *s* electrons and six *p* electrons)

-- this is called a ***full valence shell*** (also referred to as an ***octet*** )

| | |
|---|---|
| 2 He | $1s^2$ |
| 10 Ne | $1s^2\,2s^2\,2p^6$ |
| 18 Ar | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$ |
| 36 Kr | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6$ |
| 54 Xe | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6$ |
| 86 Rn | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6$ |

For the elements in Period 1 (hydrogen and helium), a full valence shell consists of 2 valence electrons (two *s* electrons)

- this is because the first principal energy level ($n$ = 1) does ***not*** have a *p* sublevel (*i.e.,* no *p* orbitals)

## Alkali metals are highly reactive because they can obtain a stable noble gas configuration by losing one electron

| | | | |
|---|---|---|---|
| 3 Li | $1s^2\,2s^1$ | 2 He | $1s^2$ |
| 11 Na | $1s^2\,2s^2\,2p^6\,3s^1$ | 10 Ne | $1s^2\,2s^2\,2p^6$ |
| 19 K | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1$ | 18 Ar | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$ |
| 37 Rb | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^1$ | 36 Kr | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6$ |
| 55 Cs | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^1$ | 54 Xe | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6$ |
| 87 Fr | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6\,7s^1$ | 86 Rn | $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6$ |

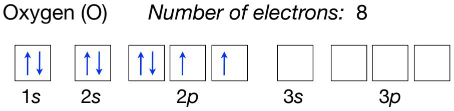
## Electron configuration of ions

**Anions (negative ions):** Formed by *adding* electrons to neutral atoms

- electrons are added to the empty or partially empty orbital with the lowest value of ***n***

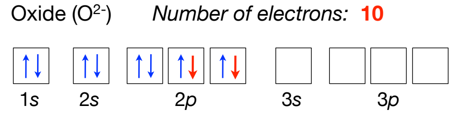
Example:

Oxygen (O) *Number of electrons:* 8




$1s^2\,2s^22p^4$

Oxide ($O^{2-}$) *Number of electrons:* **10**




$1s^2\,2s^2\mathbf{2p^6}$

[Ne]

## Electron configuration of ions

**Cations (positive ions):** Formed by *removing* electrons from neutral atoms

- electrons are removed first from occupied orbitals with the highest principal quantum number $n$
- if there is more than one occupied subshell for a given value of $n$, electrons are first removed from the orbital with the highest value of $l$

For atoms containing occupied *d* subshells:

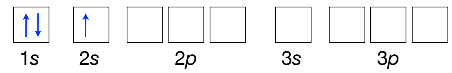
- electrons are removed from the n*s* and n*p* subshells before they are removed from the (n-1)*d* subshell

## Electron configuration of ions

**Cations (positive ions):** Formed by *removing* electrons from neutral atoms

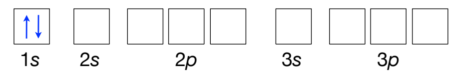
Example:

Lithium (Li) *Number of electrons:* 3



$1s^2\,2s^1$

Lithium ion ($Li^+$) *Number of electrons:* **2**



**$1s^2$**

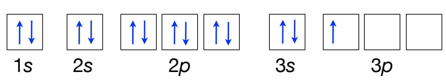
[He]

## Electron configuration of ions

**Cations (positive ions):** Formed by *removing* electrons from neutral atoms

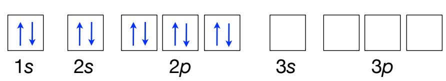
Example:

Aluminum (Al) *Number of electrons:* 13



$1s^2\,2s^2 2p^6 3s^2 3p^1$

Aluminum ion ($Al^{3+}$) *Number of electrons:* **10**



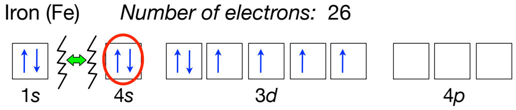
$1s^2\,2s^2\mathbf{2p^6}$

[Ne]

## Electron configuration of ions

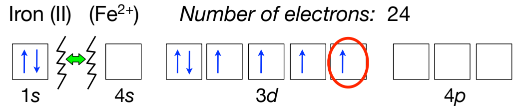
Example:

Iron (Fe) *Number of electrons:* 26



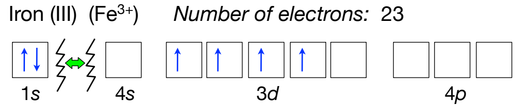
[Ar] $4s^2\,3d^6$

Iron (II) ($Fe^{2+}$) *Number of electrons:* 24



[Ar] $3d^6$

Iron (III) ($Fe^{3+}$) *Number of electrons:* 23

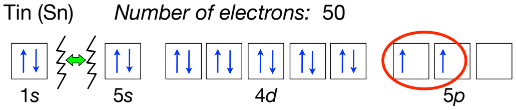


[Ar] $3d^5$

## Electron configuration of ions

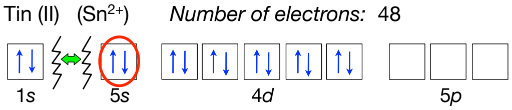
Example:

Tin (Sn) *Number of electrons:* 50



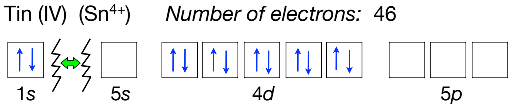
[Kr] $5s^2\,4d^{10}\,5p^2$

Tin (II) ($Sn^{2+}$) *Number of electrons:* 48



[Kr] $5s^2\,4d^{10}$

Tin (IV) ($Sn^{4+}$) *Number of electrons:* 46

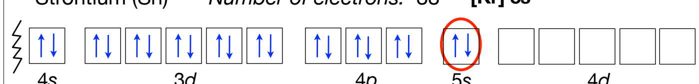


[Kr] $4d^{10}$

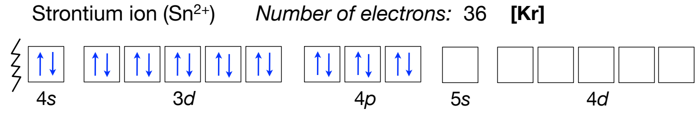
## Electron configuration of ions

Example:

Strontium (Sn) *Number of electrons:* 38 **[Kr] $5s^2$**



Strontium ion ($Sn^{2+}$) *Number of electrons:* 36 **[Kr]**



## Periodic properties of the elements

Certain physical/chemical properties of atoms follow predictable trends in the periodic table as you move:

- ***down*** *a column (group)*
- *from* ***left to right*** *across a row (period)*

*These are called "periodic properties" because they follow a pattern that repeats itself for each row of elements (period) in the periodic table*

**periodicity** -- characterized by a repeating or recurrent cycle ("period")

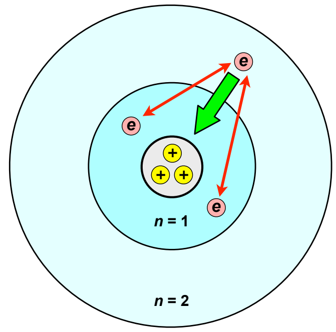
Periodic properties include:

- ionic radius
- ionization energy
- electron affinity

*These properties depend on how strongly their outer electrons are attracted to the nucleus*

## Concept of "effective nuclear charge"

*Many properties of atoms depend on how strongly their outer electrons are attracted to the nucleus*




*Example: Lithium*

Electrons feel an *attractive* force to the positively charged protons in the nucleus (electrostatic force)

From Coulomb's law:

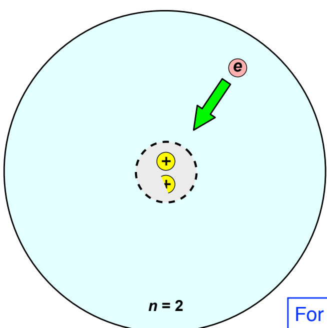
The force is proportional to the magnitude of the charge in the nucleus

***more protons = stronger force***

At the same time, electrons feel *repulsive forces* from the other negatively charged electrons in the atom

## Concept of "effective nuclear charge"

*Many properties of atoms depend on how strongly their outer electrons are attracted to the nucleus*



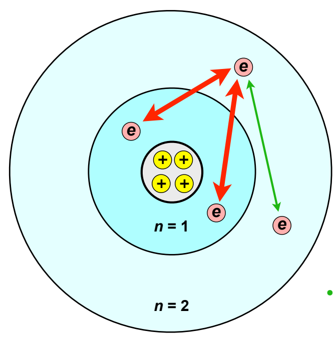

*Example: Lithium*

The net result is that electrons are attracted to the nucleus, but not as strongly as they would be if there were no other electrons present

- the outer electrons are partially shielded from the nuclear charge by the inner electrons
- outer electrons "feel" an an effective positive charge from the nucleus that is less than the actual charge

For an electron in the 2*s* orbital of a lithium atom, the **effective nuclear charge** is **+1.3** (actual nuclear charge is +3)

## Core electrons screen nuclear charge more effectively than electrons in the same shell




*Example: Beryllium*

*Core electrons* are very effective at shielding outer electrons from the charge contained in the nucleus

- *core electrons* are *inner electrons*
  -- *i.e.*, electrons in lower-energy electron shells

Electrons in the *same valence shell* do not shield each other very effectively from nuclear charge

- the *valence shell* is the *outermost shell*
  -- *i.e.*, highest-energy electron shell (highest value of ***n***)

## Calculating effective nuclear charge

The **effective nuclear charge** ($Z_{eff}$) felt by an electron in an atom can be calculated as follows:

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

No. of protons in nucleus (atomic number) — Screening constant

The screening constant ($S$) is a positive number that represents the amount of nuclear charge screened by other electrons in the atom

The value of $S$ can be determined by the following methods:

1) Assume that it is equal to the number of core electrons in the atom
2) Use **Slater's rules**
   -- other electrons in the same shell contribute 0.35 to $S$
   -- electrons in the next lower shell ($n$-1) contribute 0.85 to $S$
   -- electrons in lower shells ($n$-2, $n$-3, etc.) contribute 1.00 to $S$

## Calculating effective nuclear charge

*Example:* What is $Z_{eff}$ felt by a valence electron in a lithium atom?

**Atomic number: $Z = 3$**
**Electron configuration: $1s^2 2s^1$**

Method 1: $S$ = number of core electrons = 2

$$Z_{eff} = Z - S = 3 - 2 = \mathbf{+1}$$

Method 2: Valence electron is in the $n = 2$ shell

Number of other electrons in the same shell = 0
Number of electrons in the next lower shell ($n = 1$) = 2
Number of electrons lower shells = 0

$$S = (0 \times 0.35) + (2 \times 0.85) + (0 \times 1.00) = 1.7$$

$$Z_{eff} = Z - S = 3.0 - 1.7 = \mathbf{+1.3}$$

## Calculating effective nuclear charge

Example: What is $Z_{eff}$ felt by a valence electron in a chlorine atom?

**Atomic number: $Z = 17$**
**Electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^5$**

Method 1: $S$ = number of core electrons = 10

$$Z_{eff} = Z - S = 17 - 10 = \mathbf{+7}$$

Method 2: Valence electron is in the $n = 3$ shell

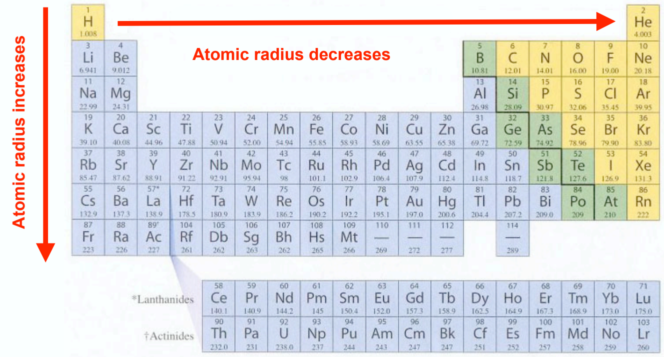
Number of other electrons in the same shell = 6
Number of electrons in the next lower shell ($n = 2$) = 8
Number of electrons lower shells ($n = 1$) = 2

$$S = (6 \times 0.35) + (8 \times 0.85) + (2 \times 1.00) = 10.9$$

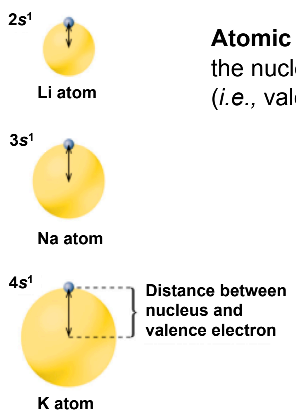
$$Z_{eff} = Z - S = 17 - 10.9 = \mathbf{+6.1}$$

## Periodicity: Atomic radius

**Atomic radius** -- the distance between the nucleus of an atom and its outermost (*i.e.,* valence) electrons




## Going down a group, atomic radius *increases*

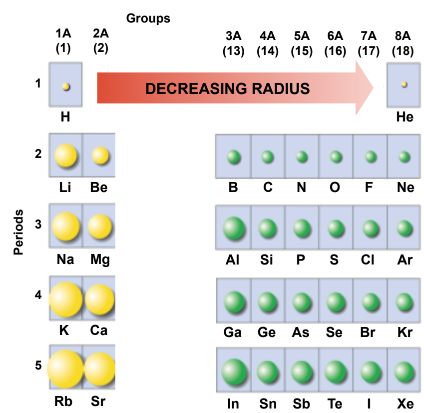



**Atomic radius** -- the distance between the nucleus of an atom and its outermost (*i.e.,* valence) electrons

Atomic radius ***increases*** going down a group in the periodic table

- as the principal energy level ($n$) increases, electrons are found on average farther away from the nucleus

## Going across a period, atomic radius *decreases*



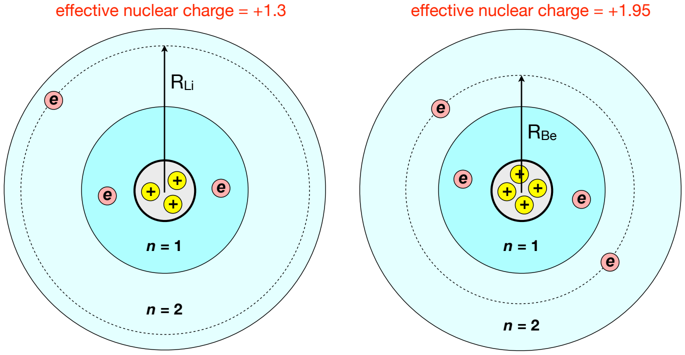

How can the size of the atom decrease as more electrons are added?

Remember that for all of the elements in a given period, the ***valence electrons*** occupy the ***same*** principal energy level

- effective nuclear charge increases from left to right across a period

**Lithium**
1 valence electron in principal energy level $n = 2$
3 protons in nucleus
effective nuclear charge = +1.3

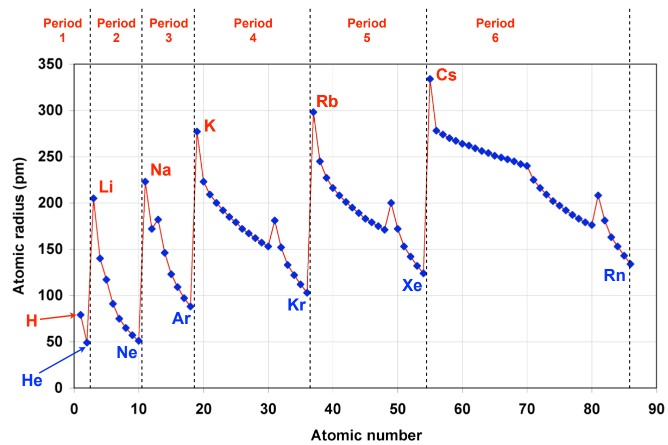
**Beryllium**
2 valence electrons in principal energy level $n = 2$
4 protons in nucleus
effective nuclear charge = +1.95




The **effective nuclear charge** felt by *valence electrons* in Be is greater than the **effective nuclear charge** felt by *valence electrons* in Li

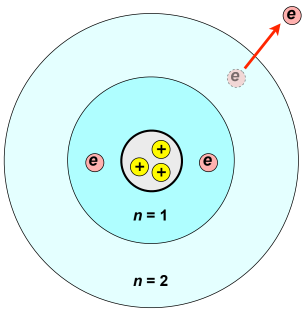
- **the valence electrons of Be are pulled in closer ($R_{Be} < R_{Li}$)**

## Periodicity: Atomic radius



## Ionization energy

**Ionization energy** -- the energy required to remove the *least* tightly bound electron (*i.e.,* a valence electron) from the ground state of an atom in the gaseous state




***Example:*** *Lithium*

$Li(g)$ + **520 kJ / mol** ⟶ $Li^{+}(g)$ + $e^{-}$

Ionization energy

Ionization energy measures the ease with which an atom ***loses*** an electron

*The lower the value, the more readily the atom loses an electron*

## Ionization energy trends

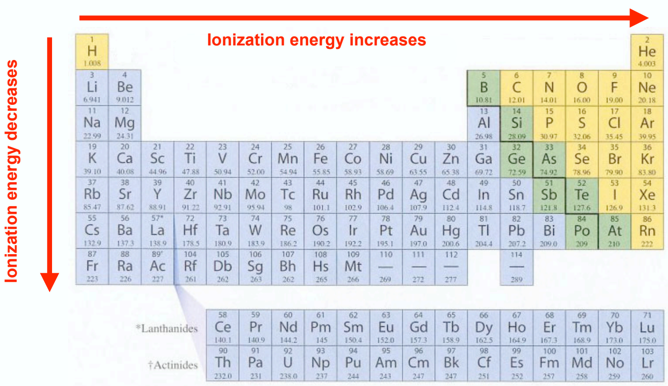
Ionization energy generally <u>increases</u> moving left to right across a row of elements (period) in the periodic table

- *effective nuclear charge increases from left to right across a period*
- *valence electrons are more strongly bound to the atom, requiring higher energy to pull them away*

Ionization energy <u>decreases</u> moving down a column (group) in the periodic table

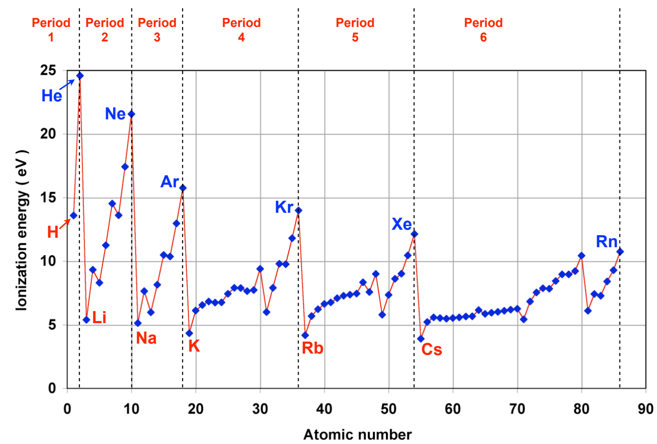
- *valence electrons are situated farther away from the nucleus (this overcomes the slight increase in effective nuclear charge moving down a column)*
- *the force binding valence electrons to the atom is weaker, requiring less energy to pull them away*

## Ionization energy




**In general:** Ionization energy is ***low*** for metals and ***high*** for non-metals

## Periodicity: Ionization energy




## Successive ionization energies

**First ionization energy ($I_1$):** The energy required to remove the first electron from a neutral atom

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

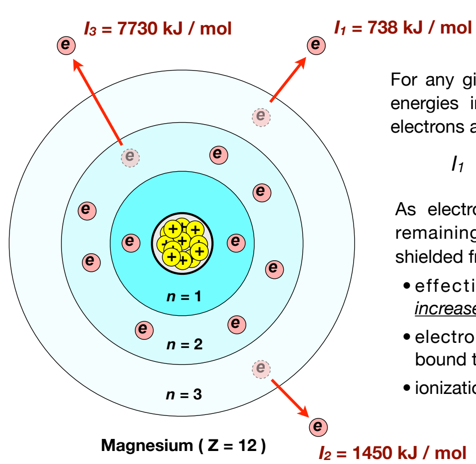
**Second ionization energy ($I_2$):** The energy required to remove the second electron from a neutral atom

$Mg^{+}(g) \longrightarrow Mg^{2+}(g) + e^{-}$   $\boldsymbol{I_2 = 1450\ kJ / mol}$

**Third ionization energy ($I_3$):** The energy required to remove the third electron from a neutral atom

$Mg^{2+}(g) \longrightarrow Mg^{3+}(g) + e^{-}$   $\boldsymbol{I_3 = 7730\ kJ / mol}$

## Successive ionization energies




**Magnesium ( Z = 12 )**

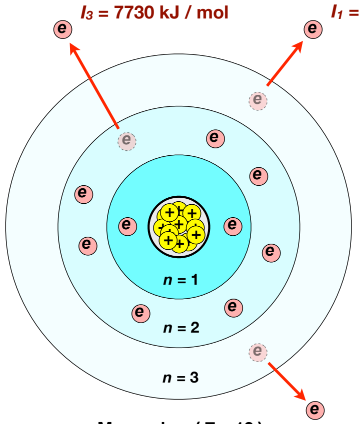
For any given element, ionization energies increase as successive electrons are removed:

$I_1 < I_2 < I_3 \ldots$

As electrons are removed, the remaining electrons are *<u>less</u>* shielded from the nucleus

- effective nuclear charge *<u>increases</u>*
- electrons are more tightly bound to nucleus
- ionization energy *<u>increases</u>*

## Successive ionization energies




Magnesium ( Z = 12 )

**But why is $I_3$ so much larger than $I_1$ and $I_2$ for the Mg atom?**

The first two electrons are removed from the outer (valence) shell

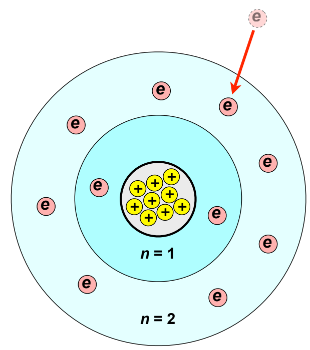
But the third electron is removed from an inner (core) shell

- core electrons are closer to the nucleus
- core electrons are more tightly bound to nucleus
- <u>*much*</u> more energy required to remove core electrons

***That is why $Mg^{3+}$ ion is not observed***

## Electron affinity

**Electron affinity** -- the energy change associated with adding an electron to an atom in the gaseous state




***<u>Example:</u>*** *<u>Fluorine</u>*

$F(g) + e^- \longrightarrow F^-(g) + $ **328 kJ / mol**

**Electron affinity = ΔE = –328 kJ / mol**

*sign is negative because energy is released (exothermic process)*

Electron affinity measures the ease with which an atom ***gains*** an electron

*The more negative the value, the more readily the atom gains an electron*

## Electron affinities (kJ/mol) for the *s*- and *p*-block elements in the first five periods

| | | | | | | | |
|---|---|---|---|---|---|---|---|
| **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 |

Periodic trends for electron affinity are not as clear as they are for atomic radius and ionization energy.

<u>In general</u>: Non-metals have <u>*high*</u> electron affinity, metals have <u>*low*</u> electron affinity