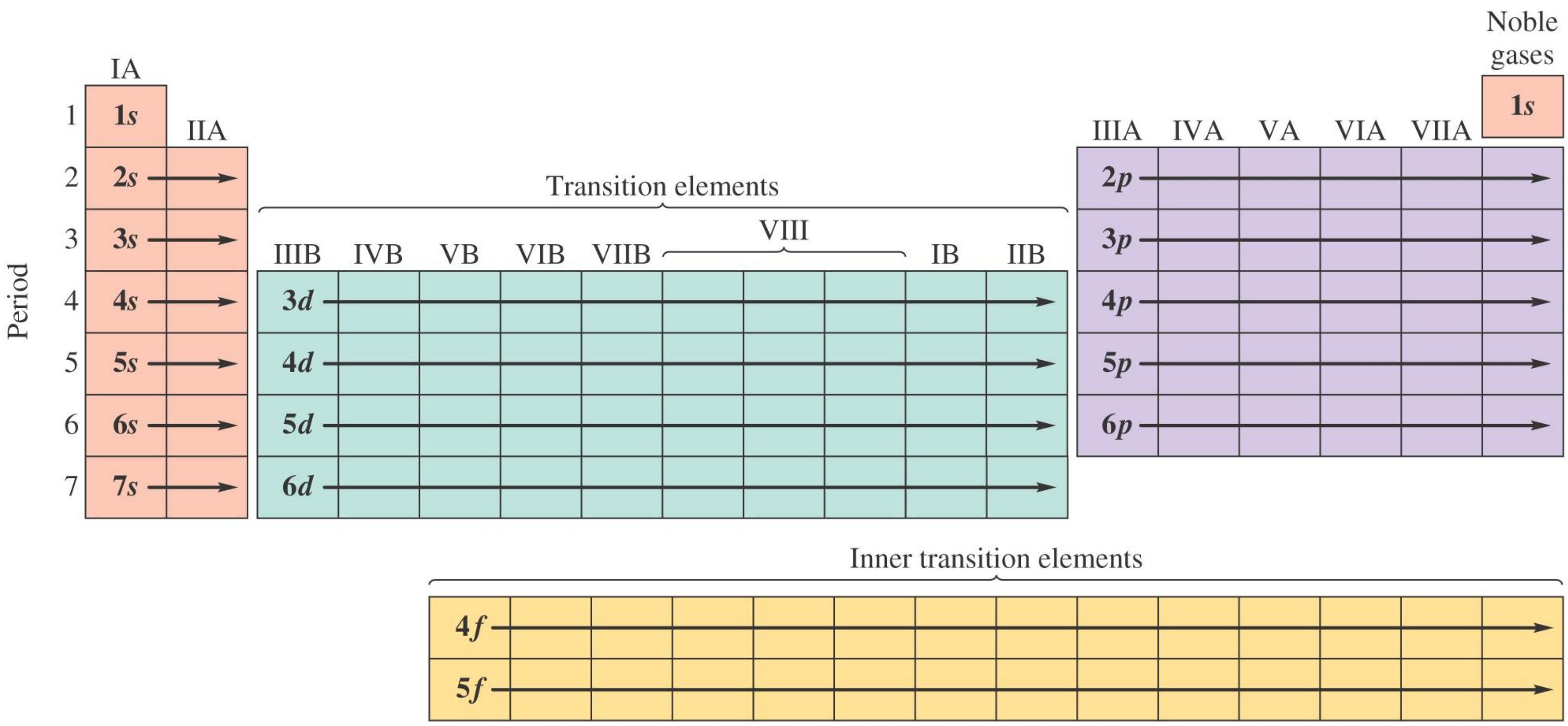
# ELECTRON CONFIGURATION AND THE PERIODIC TABLE

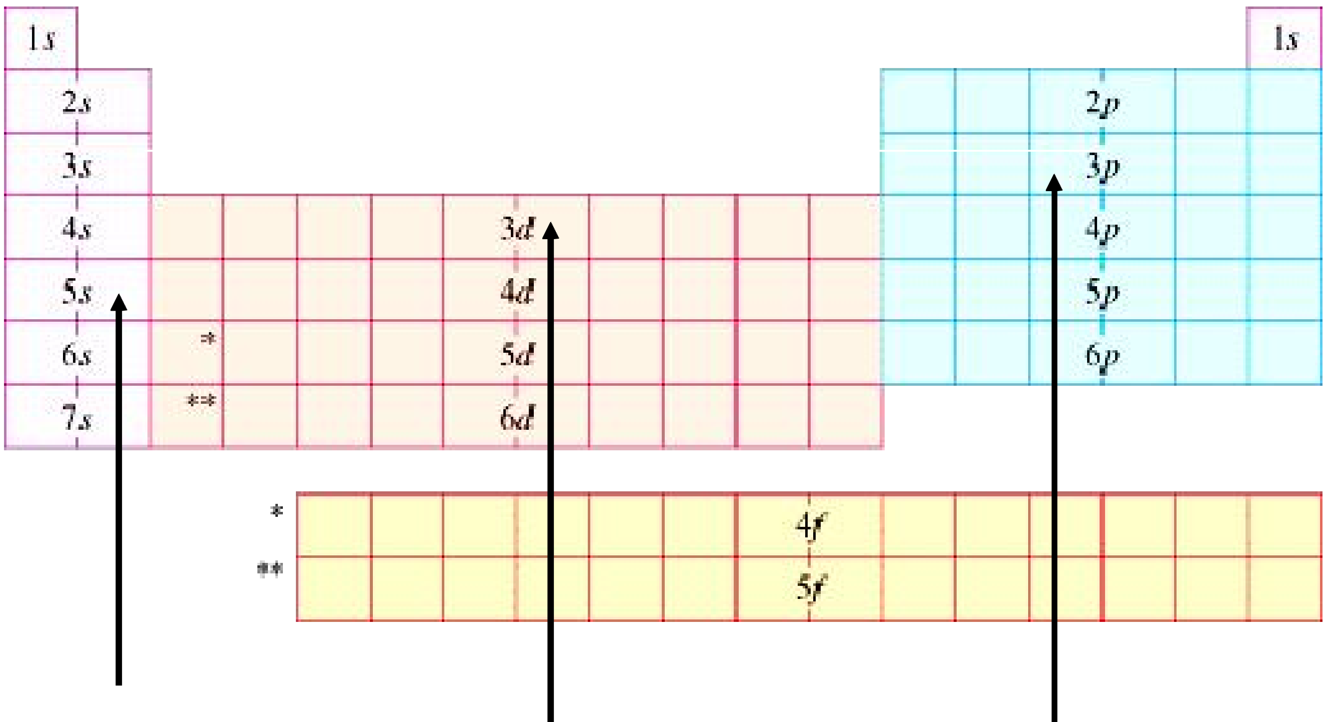
- The electrons in an atom fill from the lowest to the highest orbitals.
- The knowledge of the location of the orbitals on the periodic table can greatly help the writing of electron configurations for large atoms.



**For A Groups (Representative Elements):**
**(s and p blocks) n = period number**

**For B Groups (Transition Elements)**
**d blocks: n = period number - 1**
**f blocks: n = period number - 2**

- The electronic configuration of an element can now be given if the position in the Periodic Table is known:




**$1s^2\ 2s^22p^6\ 3s^23p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$**
**or [Kr] $5s^2$**
**(2 outermost shell electrons)**

**$1s^2\ 2s^22p^6\ 3s^23p^3$**
**or [Ne] $3s^23p^3$**
**(5 outermost shell electrons)**

**$1s^2\ 2s^22p^6\ 3s^23p^6\ 4s^2\ 3d^6$**
**or [Ar] $4s^2\ 3d^6$**
**(2 outermost shell electrons)**

- Group VIIIA elements (Noble Gases) have a stable outermost shell electron configuration:

| Element | Orbital Notation | Electron Configuration | # of Outermost Shell Electrons |
|---|---|---|---|
| He | ↑↓ <br> **1s** | $1s^2$ | 2 |
| Ne | ↑↓ ↓↑ ↑↓ ↑↓ <br> **2s 2p** | $2s^2\ 2p^6$ | 8 |
| Ar | ↑↓ ↓↑ ↑↓ ↑↓ <br> **3s 3p** | $3s^2\ 3p^6$ | 8 |
| Kr | ↑↓ ↓↑ ↑↓ ↑↓ <br> **4s 4p** | $4s^2\ 4p^6$ | 8 |
| Xe | ↑↓ ↓↑ ↑↓ ↑↓ <br> **5s 52p** | $5s^2\ 5p^6$ | 8 |
| Rn | ↑↓ ↓↑ ↑↓ ↑↓ <br> **6s 6p** | $6s^2\ 6p^6$ | 8 |

- The electron configuration of **$ns^2\ np^6$** (or $ns^2$) is referred to as a **"noble gas core"**

- Chemical Properties of elements are determined by the outermost shell electrons (electrons in highest energy level); these electrons are involved in bonding.

## VALENCE ELECTRONS:

- Valence electrons are those outside the noble-gas core

- For transition elements, the "s" and "p" electrons of the outermost shell (highest energy level) ("d" electrons are not valence electrons)

**NOTE**:

- Most transition elements have 2 valence electrons
  Reason: They fill "d" or "f" orbitals which are not part of the outer most shell

- Exceptions: Transition elements with irregular electronic configurations: **Cr, Cu, etc.**

# MAGNETIC PROPERTIES OF ATOMS

- Every electron acts like a small magnet
- Magnetic attractions from 2 electrons with opposite spin cancel each other

It follows:
- An atom with paired electrons has no magnetism.
- An atom with unpaired electrons (excess of one spin) exhibits magnetism.
- An element with atoms that exhibit magnetism is attracted to a strong magnet

| DIAMAGNETIC ELEMENTS | PARAMAGNETIC ELEMENTS |
|---|---|
| • Are not attracted by a magnetic field<br>• Sometimes even repelled by a magnetic field<br>• Atoms contain paired electrons only | • Are weakly attracted by a magnetic field<br>• Atoms contain unpaired electrons |
| Hg $\rightarrow$ [Xe] $6s^2\ 4f^{14}\ 5d^{10}$ | Na $\rightarrow$ [Ne] $3s^1$<br>C $\rightarrow$ [He] $2s^2\ 2p^2$ |

**FERROMAGNETIC ELEMENT**
- Are strong permanent magnetism in iron due to the alignment of many unpaired electrons

Fe [Ar] $\underline{\uparrow\downarrow}$ (4s) $\underline{\uparrow\downarrow}$ $\underline{\uparrow}$ $\underline{\uparrow}$ $\underline{\uparrow}$ $\underline{\uparrow}$ (3d)

**Examples:**
1. Write complete and condensed electron configuration for bromine (Z=35) and determine the number of valence electrons.

2. Write complete and condensed electron configuration for palladium (Z=46) and determine the number of valence electrons.

3. Write complete and condensed electron configuration for antimony (Z=51) and determine the number of valence electrons.

# PERIODIC PROPERTIES

- Periodic properties are properties of the elements that are repeated according to a regular "periodic" trend.

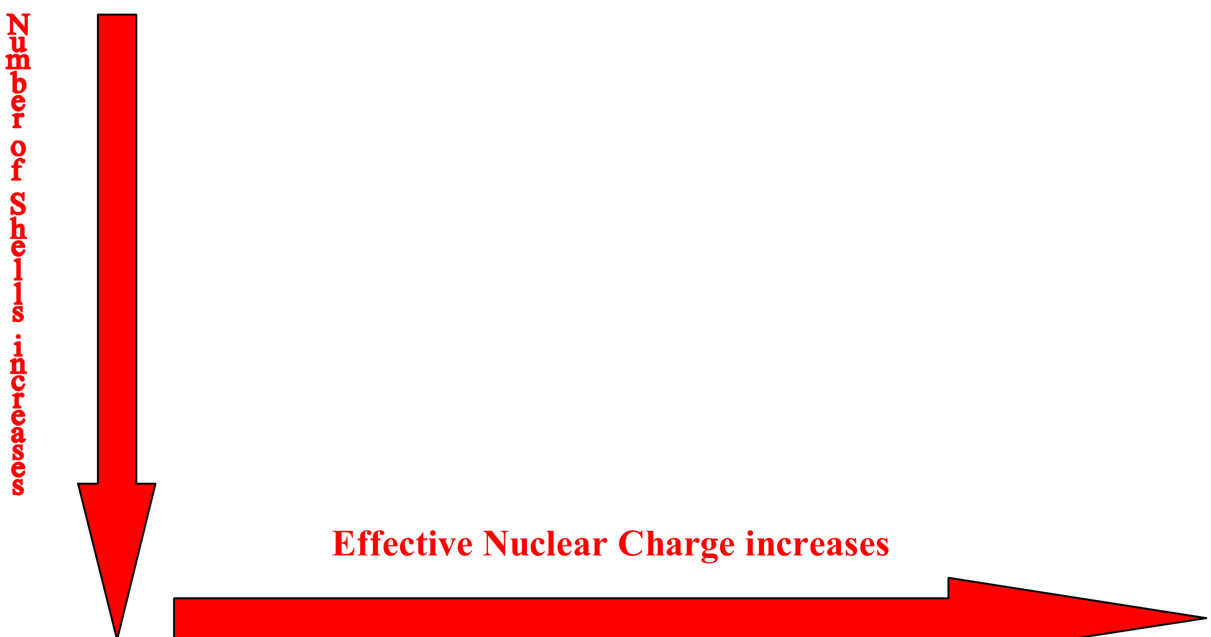
Three important periodic properties will be discussed:

1. **Atomic Radius** (size of the atoms)
2. **Ionization Energy (I.E.)** - energy needed to remove the outermost electron from a neutral atom in the gaseous state to form a positive ion.
3. **Electron Affinity (E.A.)** - energy (absorbed or **released**) during the process of adding an electron to a neutral atom in the gaseous state to form a negative ion

The variation of these 3 Periodic Properties will be discussed :

- within a group (vertical trend)
- along a period (horizontal trend)

Two important factors determine these trends:

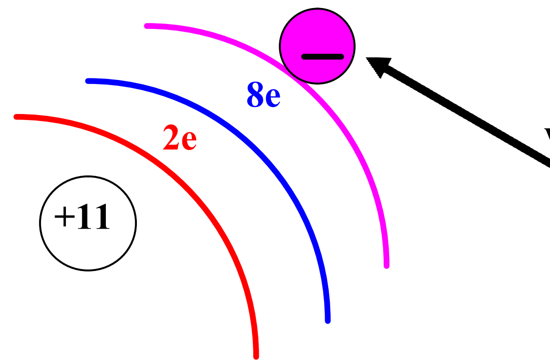


# Effective Nuclear Charge

- Effective nuclear charge ($Z_{eff}$) is the positive charge that an electron experiences from the nucleus.
- Consider the electron configuration of the Na atom ( Z = 11)

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

**3 energy levels:**
- **- 2 electrons on 1st energy level**
- **- 8 electrons on 2nd energy level**
- **- 1 electron on 3rd energy level**




**What is the positive charge that the outermost electron "feels" ?**

1. **It is not +11**, since its electrons of the 1st and 2nd energy level (10 electrons carry a charge equal to –10) = cancel out some of the + 11 charge of the nucleus ('shielding effect")

2. **It is not + 1,** since the 10 core electrons (charge = –10) cannot completely cancel out 10 positive charges of the nucleus.
Reason: The 10 core electrons are in the nucleus, but some distance away

- Note: The closer to the nucleus the core electrons are, the more effective they are in canceling out some of the positive charge of the nucleus (shielding effect)

shielding effect of a 1st shell electron > shielding effect of a 2nd shell electron

3. The outermost electron is attracted to the nucleus by a positive charge which is less than the actual nuclear charge (Z = +11) because of the shielding effect of the core electrons, but more than +1

<u>**Effective Nuclear Charge = Zeff**</u> = + 2.8 (in this particular case)

- **The Nuclear Charge an outermost electron experiences is reduced by the shielding effect of other electrons (Zeff = Z – Shielding Effect)**

**Zeff = Z – Shielding effect**

**Zeff**: Effective Nuclear Charge

**Z**: Actual Nuclear Charge

**Along a period, moving from left to right:**

Protons are added to the nucleus → Z increases considerably

Electrons are added to the outermost shell → Minimal increase in shielding effect

**Zeff = Z – Shielding effect**

Zeff: Effective Nuclear Charge

Z: Actual Nuclear Charge → Increases considerably

Shielding effect → Minimal increase

| Moving across a period **Zeff increases** (Z increases a lot, Shielding Effect increases little) |
|---|

**Along a group, moving from up to down:**

Protons are added to the nucleus → Z increases considerably

Electrons are added to the new shell → Considerable increases shielding effect

**Zeff = Z – Shielding effect**

Zeff: Effective Nuclear Charge

Z: Actual Nuclear Charge → Increases considerably

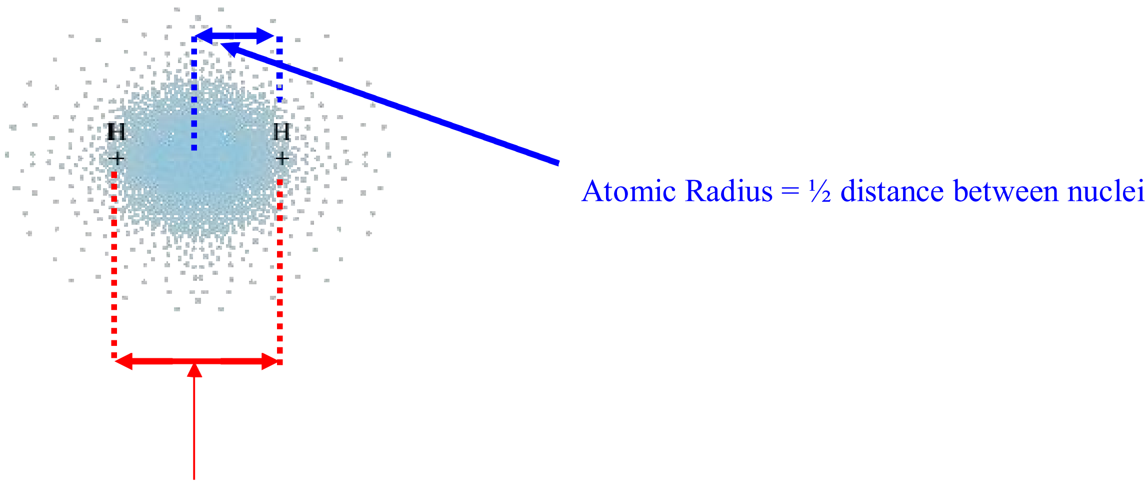
Shielding effect → Increases considerably

| Moving down a group **Zeff does NOT increase** (Z increases a lot, Shielding Effect increases a lot) |
|---|

# ATOMIC RADIUS

- Values of Atomic radii are obtained from measurements of distances between the nuclei of atoms in the chemical bonds of molecular substances.

  Example: experimental Determination of the radius of H atom:




distance between nuclei is determined experimentally (by X-ray crystallography)

- Atomic Radius determined in this manner are referred to as Covalent Radii.

- Atoms are very small; consequently, atomic radii have very small values

- Values of Atomic Radii are listed on the Periodic Table (back side) and are commonly expressed in:

  Angstroms (1 A° = $10^{-10}$ m)

  or

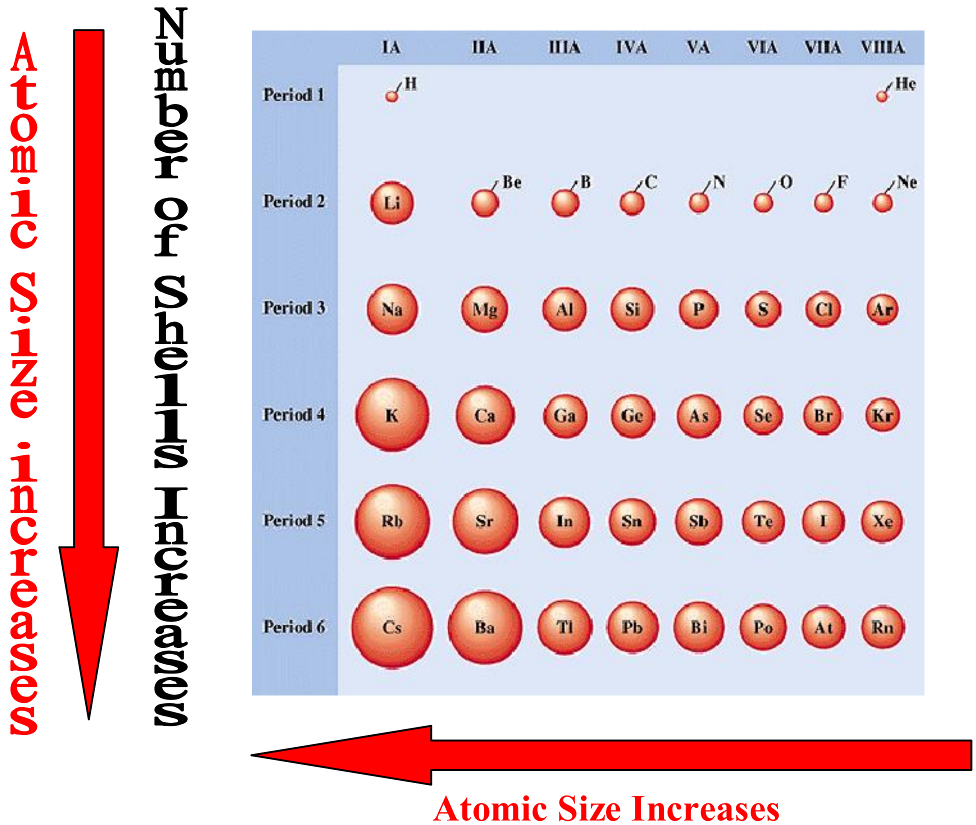
  Nanometers (1 nm = $10^{-9}$ m)

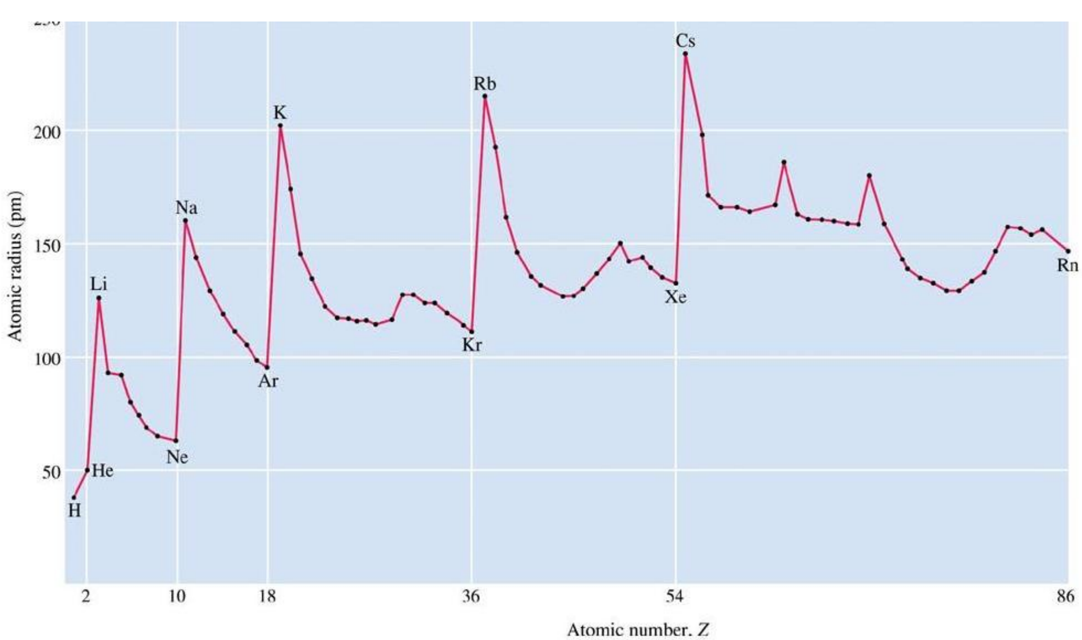
  or

  Picometers (1 pm = $10^{-12}$ m)

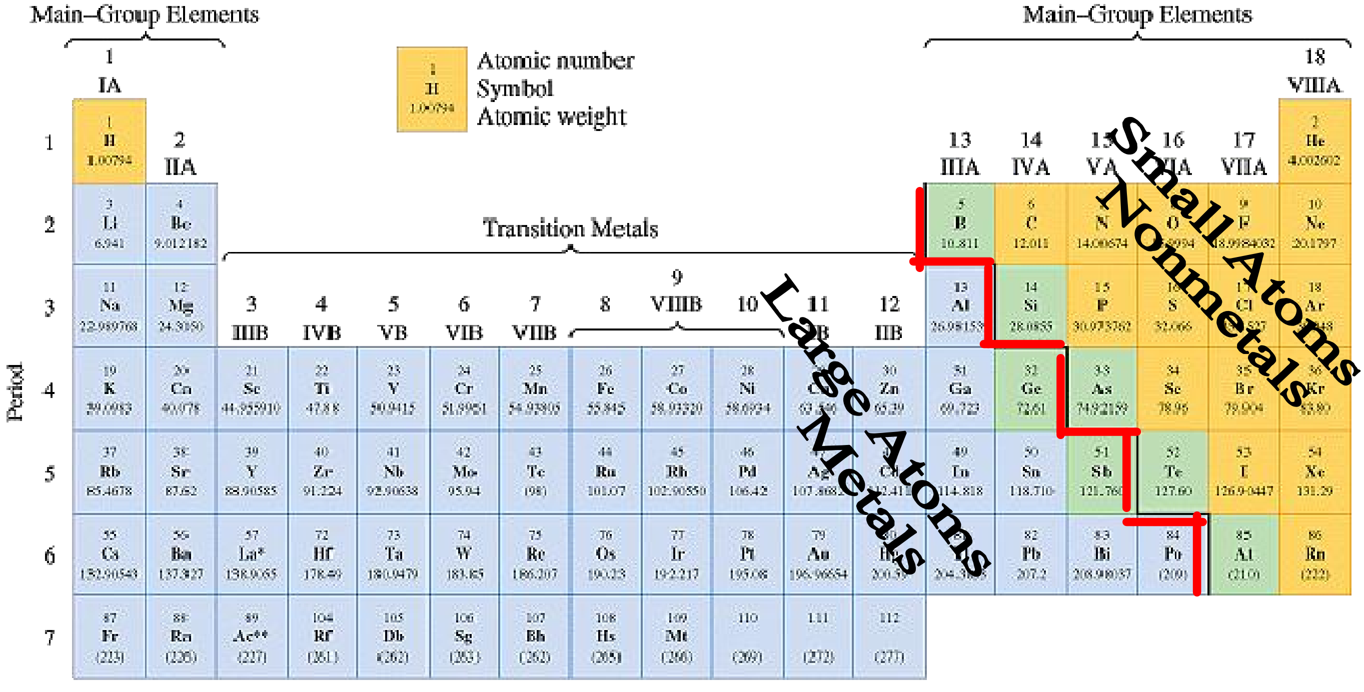
- H is the smallest atom: Covalent radius of H atom = 0.32 A

# ATOMIC RADII



# ATOMIC RADII




Note: There is little variation in atomic Size throughout a row of Transition Elements:

Reason: Consider the elements completing the 3d subshell:

From Sc ……………………..to …………………Zn

[Ar] $4s^2 3d^1$ ………. …...to ……………… [Ar] $4s^2 3d^{10}$

Moving along period 4 from left to right:

- **1 proton is added when moving from one transition element to the next**
  **ACTUAL NUCLEAR CHARGE (Z) INCREASES CONSIDERABLY**

Sc ……………………………………………………………………Zn

- 1 electron is added to the 3d subshell of the 3rd shell (an inner shell)
  SHIELDING EFFECT CAUSED BY ADDED ELECTRON INCREASES CONSIDERABLY

**Recall:** **Zeff** = **Z** – **Shielding Effect**

**increases considerably** **increases considerably**

**Result:** **Zeff does not change**

- **Insignificant variation in Atomic Size for 3d block elements**
  This trend holds true for all transition elements (d block elements)

# IONIZATION ENERGY

- Ionization Energy (IE) is the energy needed to remove the outermost electron from a neutral atom in the gaseous state to form a positive ion.
- Ionization Energy :
    - is commonly measured in kJ/mol
    - is also referred to as Ionization Potential (electrical energy) and can also be measured in Volts
    - is listed on the Periodic Table (back side)

Consider the removal of the outermost shell electron from sodium:

$$Na(g) \quad + \quad IE_1 \longrightarrow Na^{+}(g) \quad + \quad 1\ e^{-}$$

| Na(g) | IE$_1$ | Na$^+$(g) |
|---|---|---|
| $1s^2\,2s^2 2p^6\,3s^1$ | First Ionization Energy | $1s^2\,2s^2 2p^6$ |
| [Ne] $3s^1$ | | [Ne] |
| (11p, 11e) | | (11p, 10e) |

Consider the removal of a second electron:

$$Na^{+}(g) \quad + \quad I\,E_2 \longrightarrow Na^{2+}(g\,) \quad + \quad 1\ e^{-}$$

| Na$^+$(g) | I E$_2$ | Na$^{2+}$(g) |
|---|---|---|
| $1s^2\,2s^2 2p^6$ | Second Ionization Energy | $1s^2\,2s^2 2p^5$ |
| [Ne] | | [He] $2s^2 2p^5$ |
| (11p, 10e) | | (11p, 9e) |

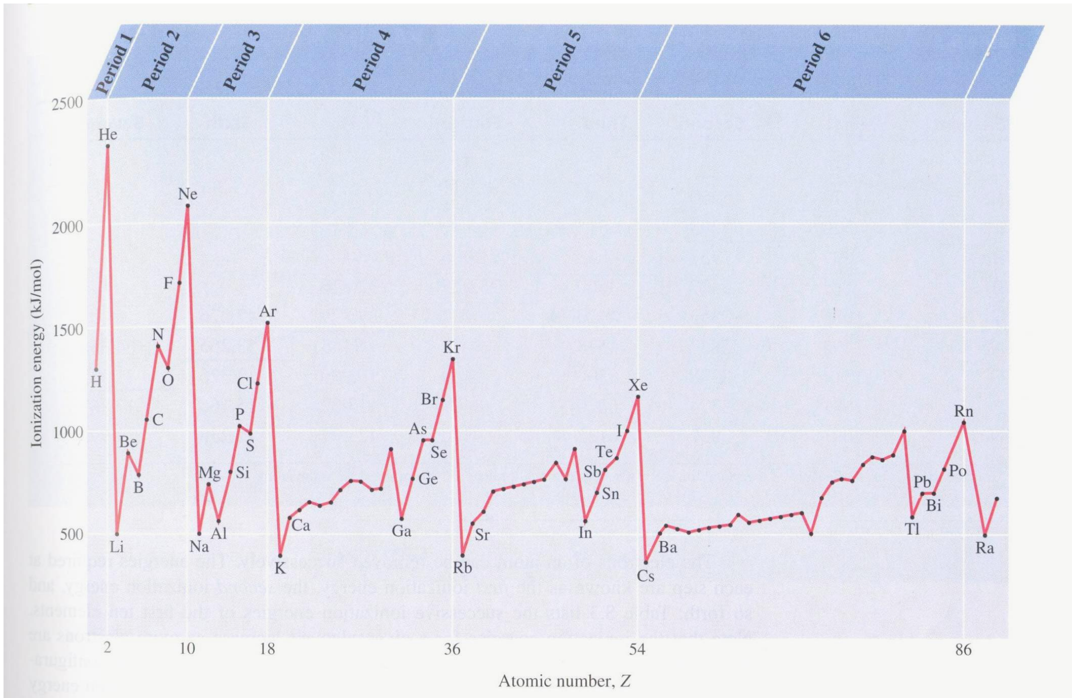
**NOTES:**

1. Removal of an electron is an endothermic process (it requires energy to remove an electron)
   Atoms do not "lose" electrons; Energy is required to remove the electrons

2. All the electrons can be removed successively form an atom.
   The energies required at each step are known as:
   First Ionization Energy (IE$_1$), Second Ionization Energy (IE$_2$), Third Ionization Energy (IE$_3$) ….

3. In general:
   The Ionization Energies for a given element increase as more electrons are removed:
   $I\,E_1 < I\,E_2 < I\,E_3 < I\,E_4 < I\,E_5 < I\,E_6 < I\,E_7 < I\,E_8$ and so on.
   Reason: The electron is being removed from ions with increasingly larger positive charges
   (The larger the positive charge of the ion, the stronger the electron is attracted)

# IONIZATION ENERGY




- Down a group IE decreases, due to increased atomic size (the larger the atom, the less strongly it holds its electrons in the outer shell)

- Across a period, IE generally increases, due to increased effective nuclear charge (the greater the nuclear charge, the stronger the atom hold its electrons in the outer shell)

- Some anomalies in IE trend occur in each period (e.g. in period 2, B is lower than Be and O is lower than N; similarly in period 3, Al is lower than Mg and S is lower than P).

- These anomalies are due to the increased stabilities of the electron shell arrangements for each atom.

- The increase in the values of successive Ionization Energies **is not gradual**:

Element Successive Ionization Energies (kJ/mol)

| | $IE_1$ | $IE_2$ | $IE_3$ | $IE_4$ | $IE_5$ | $IE_6$ | $IE_7$ |
|---|---|---|---|---|---|---|---|
| **H** | 1312 | | | | | | |
| **He** | 2372 | 5250 | | | | | |
| **Li** | 520 | 7298 | 11,815 | | | | |
| **Be** | 899 | 1757 | 14,848 | 21,006 | | | |
| **B** | 801 | 2427 | 3660 | 25,025 | 32,826 | | |
| **C** | 1086 | 2353 | 4620 | 6222 | 37,829 | 47,276 | |
| **N** | 1402 | 2857 | 4578 | 7475 | 9445 | 53,265 | |
| **O** | 1314 | 3388 | 5300 | 7469 | 10,989 | 13,326 | 71,333 |
| **F** | 1681 | 3374 | 6020 | 8407 | 11,022 | 15,164 | 17,867 |

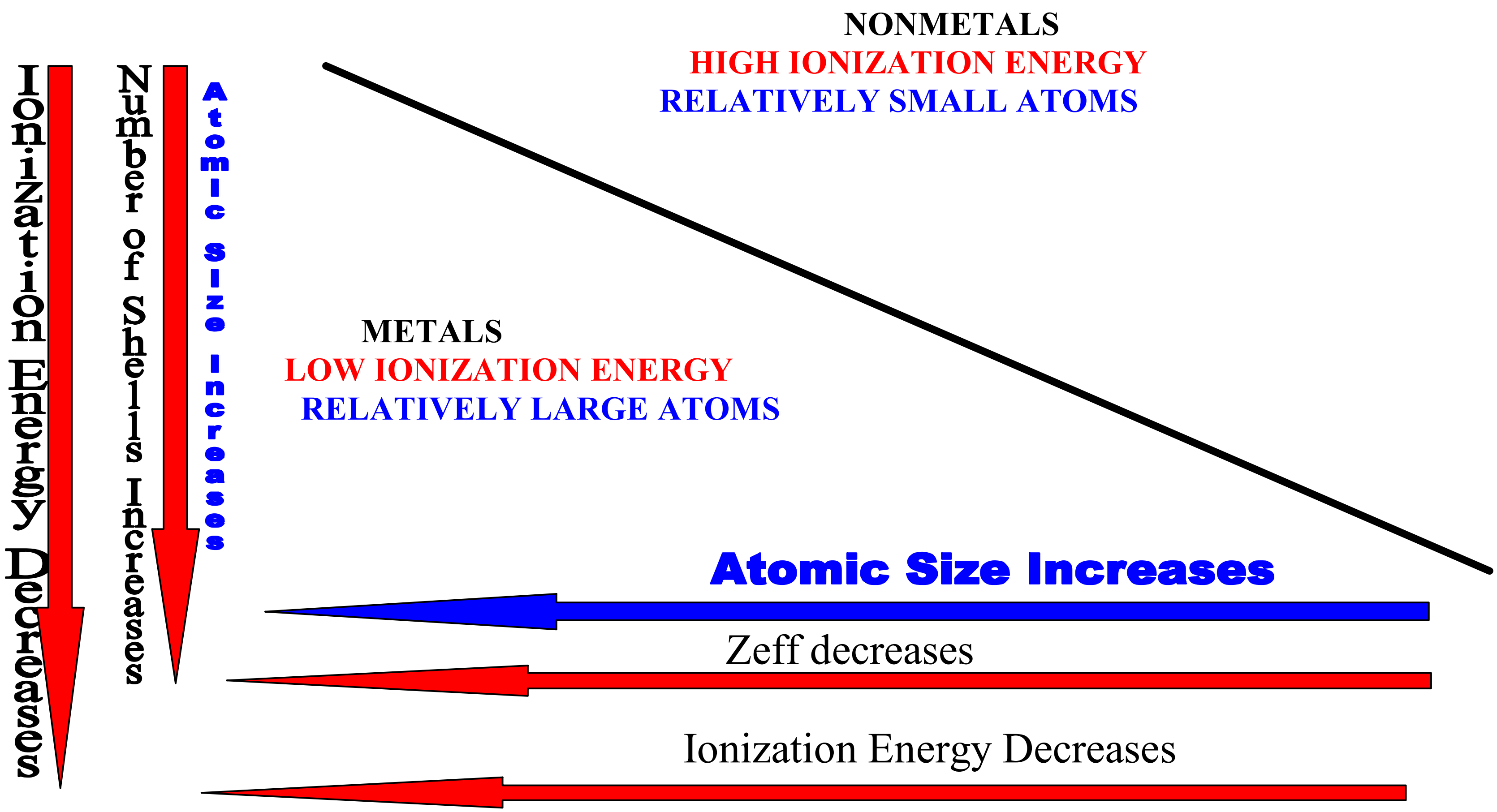
- NOTE: A very large jump in the I E value occurs after all the valence electrons have been removed

Element Successive Ionization Energies (kJ/mol)

| | $IE_1$ | $IE_2$ | $IE_3$ | $IE_4$ | $IE_5$ | $IE_6$ | $IE_7$ |
|---|---|---|---|---|---|---|---|
| **H** | 1312 | | | | | | |
| **He** | 2372 | 5250 | | | | | |
| **Li** | 520 | **7298** | 11,815 | | | | |
| **Be** | 899 | 1757 | **14,848** | 21,006 | | | |
| **B** | 801 | 2427 | 3660 | **25,025** | 32,826 | | |
| **C** | 1086 | 2353 | 4620 | 6222 | **37,829** | 47,276 | |
| **N** | 1402 | 2857 | 4578 | 7475 | 9445 | **53,265** | |
| **O** | 1314 | 3388 | 5300 | 7469 | 10,989 | 13,326 | **71,333** |
| **F** | 1681 | 3374 | 6020 | 8407 | 11,022 | 15,164 | 17,867 |

- Once the valence electrons are removed, a stable noble-gas configuration is obtained. Further removal of electrons will involve electrons from a complete inner shell.
- For purposes of comparing the chemical behavior of different elements, usually only the First Ionization Energy ($IE_1$) is compared.

# SUMMARY OF ATOMIC RADIUS & IONIZATION ENERGY



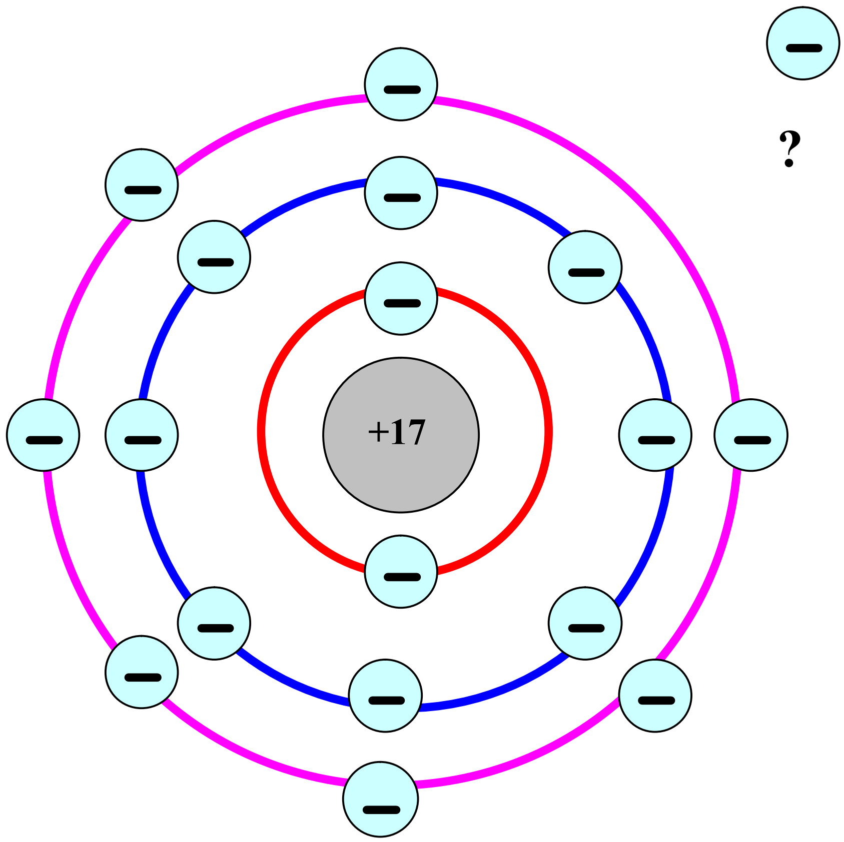
# ELECTRON AFFINITY (EA)

- Consider the electron configuration of the Cl atom (Z = 17)

**$1s^2$ $2s^2 2p^6$ $3s^2 3p^5$**

3 energy levels:

**- 2 electrons on the 1st energy level**
**- 8 electron on the 2nd energy level**
**- 7 electrons on the 3rd energy level**

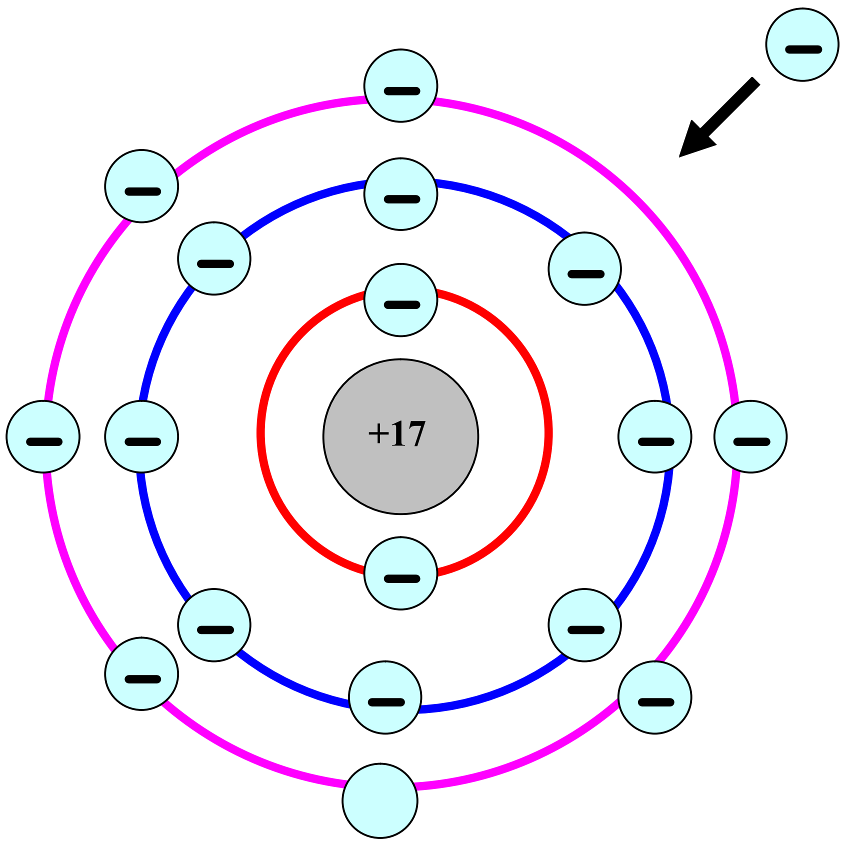



- What is the effect (if any) of the Cl atom on an outside electron ?

Hypothetically, there are 3 possibilities:

1. The atom is neutral (+ 17 nuclear charge is canceled out by the 17 negative electrons) As such, **the outside electron should be unaffected by the neighboring Cl atom**
2. The 17 electrons of the Cl atom (charge = -17) **should repel the outside electron**
3. The 17 electrons of the Cl atom cancel out the positive nuclear charge of +17 only partially (shielding effect). The positively charged nucleus creates an electric field which **should attract the outside electron**

- The outside electron is attracted to the Cl atom, and is accepted by the Cl atom.
- The addition of the electron to the Cl atom does not require energy (not endothermic)
- The electron is "naturally" attracted and releases energy (exothermic process)




$$\mathbf{Cl} + \mathbf{e^-} \longrightarrow \mathbf{Cl^-} + \mathbf{349\ kJ/mole}$$

**$1s^2\ 2s^2 2p^6\ 3s^2 3p^5$**
**[Ne] $3s^2 3p^5$**

**$1s^2\ 2s^2 2p^6\ 3s^2 3p^6$**
**[Ne] $3s^2 3p^6$**
**[Ar]**

Recall that the heat term is always written on the reactant side (Thermochemistry)

$$Cl + e^{-} \longrightarrow -349 \text{ kJ/mole} + Cl^{-}$$

$$Cl + e^{-} \longrightarrow Cl^{-} \qquad \Delta H = -349 \text{ kJ/mole}$$

**(exothermic)**

**ELECTRON AFFINITY**

- **<u>ELECTRON AFFINITY</u> (EA) is the energy change (commonly released) for the process of adding an electron to a neutral atom in the gaseous state to form a negative ion.**
  **(it is a measure of how easily an atom accepts an electron)**

**<u>NOTE:</u>**

1. The majority of atoms have a high tendency to gain electrons (ex: Cl)

**The atoms have HIGH EA's**
**HIGH EA**
**means**

**LARGE NEGATIVE EA**

**Large Amount of Energy given off when an electron is gained "willingly" by the neutral atom (stable ion formed) (addition of electron is exothermic)**

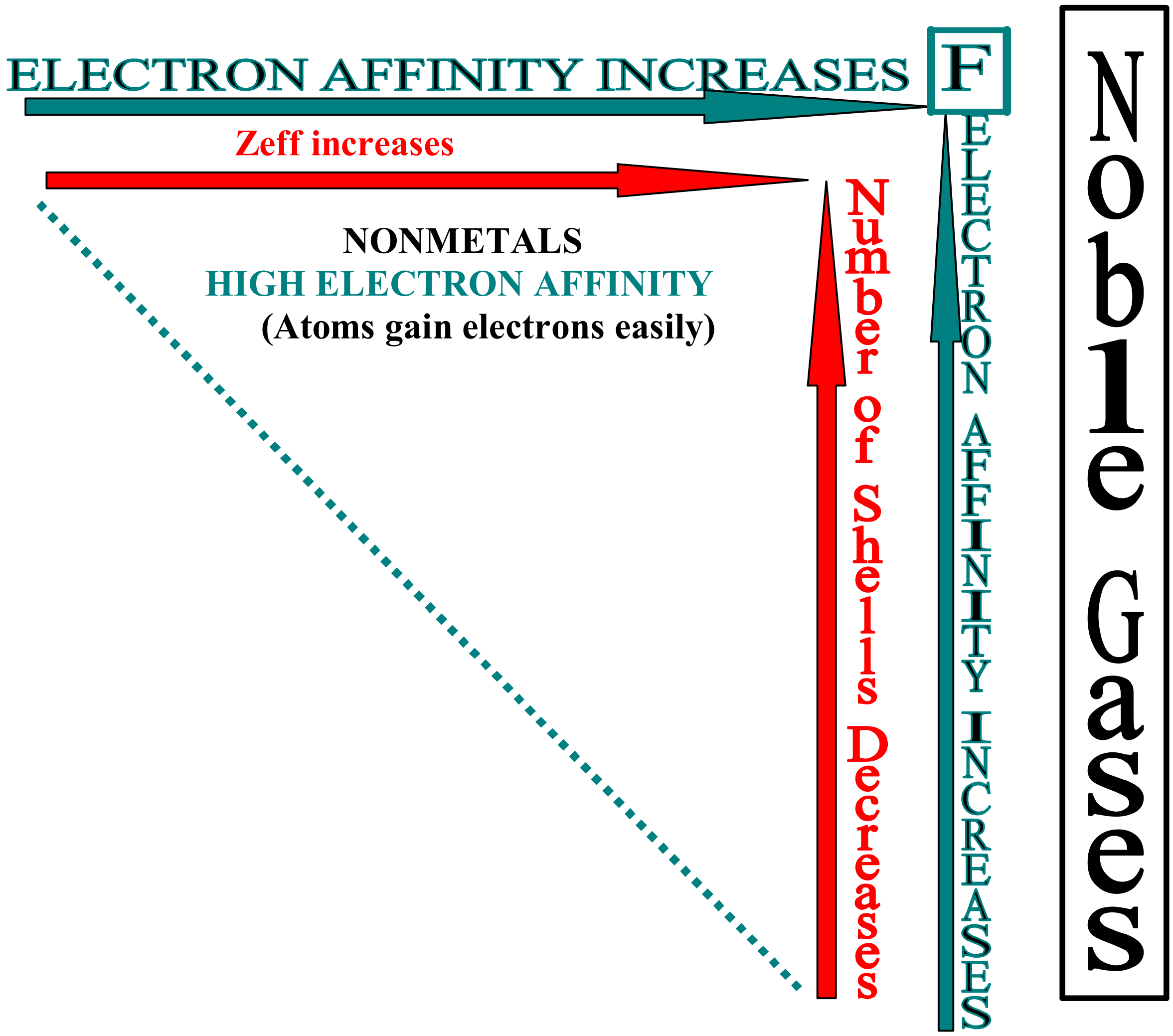
2. Relatively few atoms resist the addition of an electron(ex:Noble Gases and some metals)

**They have LOW EA's**
**LOW EA**
**means**

**Zero or Positive EA**

**Energy is required when an electron is forced in the neutral atom (unstable ion formed) (addition of the electron is endothermic)**

## ELECTRON AFFINITY TRENDS



# VARIATIONS IN ELECTRON AFFINITIES

**I. Effect caused by Interelectron repulsions between outer shell electrons**

| | IVA | VA | VIA | VIIA |
|---|---|---|---|---|
| Period 2 | **C** | **N** | **O** | **F** |
| Period 3 | **Si** | **P** | **S** | **Cl** |

**Expected Trend of EA**

**EA Should increase**

| | IVA | VA | VIA | VIIA |
|---|---|---|---|---|
| Period 2 | **C** | **N** | **O** | **F** |
| EA (kJ/mol) | –122 | 0 | –141 | –328 |
| Period 3 | **Si** | **P** | **S** | **Cl** |
| EA (kJ/mol) | –134 | –72 | –200 | –349 |

**Actual Trend of EA**

**EA increases**

**Reason:**

- Atoms of Period 2 elements are much, much smaller than the atoms of Period 3 elements (large energy gap between the 2$^{nd}$ and the 3$^{rd}$ energy levels; this energy gap diminishes as one moves to 4$^{th}$, 5$^{th}$, 6$^{th}$, 7$^{th}$ energy levels)

- In small atoms there is very strong inter-electron electron repulsion (electrons are crowded). The mutual repulsions between 2$^{nd}$ shell electrons act against adding an additional electron to the atom.

**Period 2 Elements** → **Smaller Atoms** → **Crowded electrons in outer shell (2$^{nd}$ shell)** → **Strong inter-electron Repulsions between 2$^{nd}$ shell electrons**

**Period 3 Elements** → **Larger Atoms** → **Less crowding of electrons in outer shell (3$^{nd}$ shell)** → **Weaker inter-electron Repulsions between 3$^{nd}$ shell electrons**

**II. <u>Effect caused by the stability of half-filled subshells</u>**

| | IIIA | IVA | VA | VIA | VIIA |
|---|---|---|---|---|---|
| Period 2 | **B** | **C** | **N** | **O** | **F** |
| Period 3 | **Al** | **Si** | **P** | **S** | **Cl** |

**EA should increase**

**Expected trend of EA**

| | IIIA | IVA | VA | VIA | VIIA |
|---|---|---|---|---|---|
| Period 2<br>**E.A. (kJ/mol)** | **B**<br>**–27** | **C**<br>**–122** | **N**<br>**0** | **O**<br>**–141** | **F**<br>**–328** |
| Period 3<br>**E.A. (kJ/mol)** | **Al**<br>**–44** | **Si**<br>**–134** | **P**<br>**–72** | **S**<br>**–200** | **Cl**<br>**–349** |

**E.A. increases** **E.A. increases**

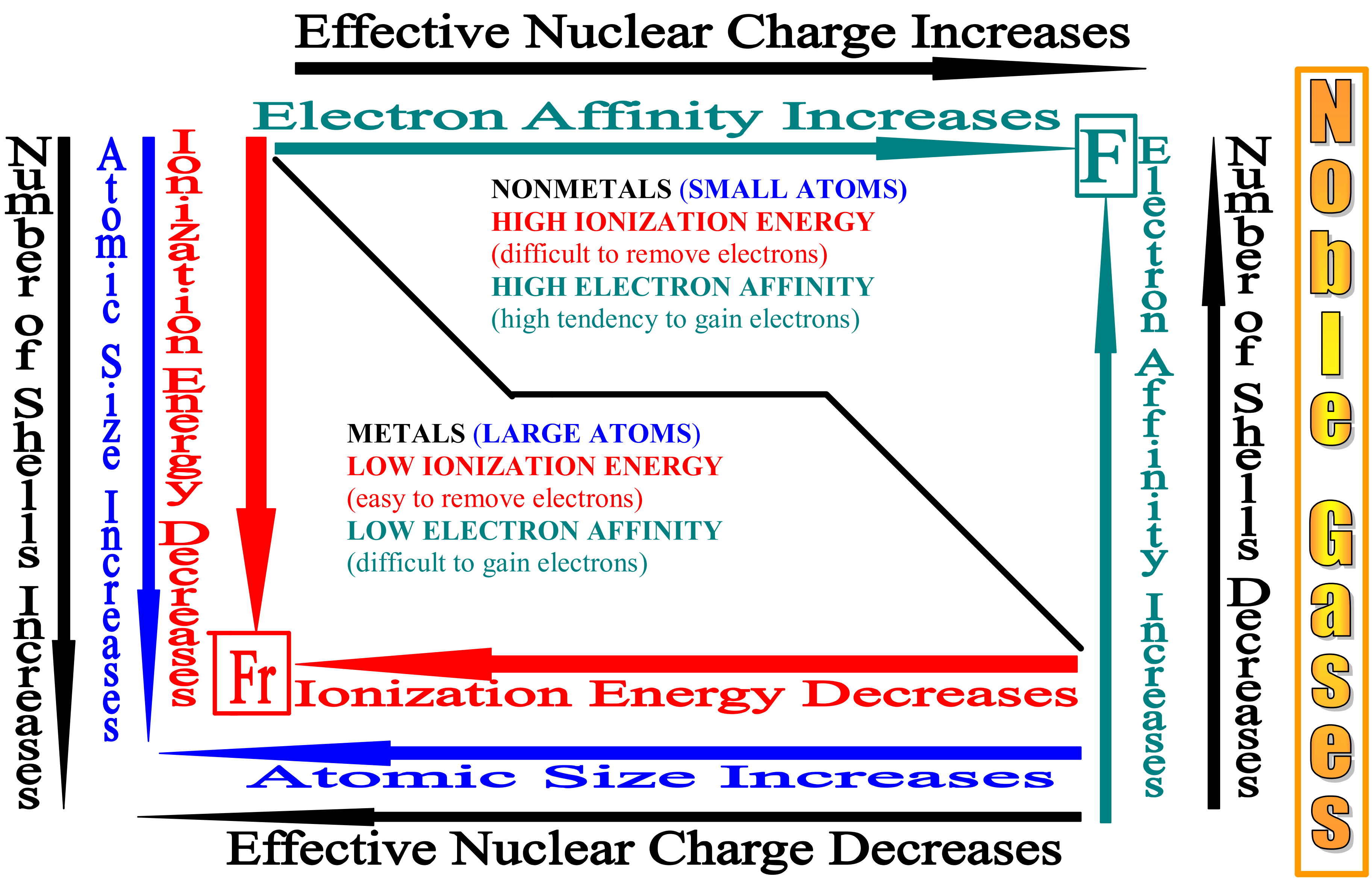
**Break in trend !**
**E.A is less than expected !**

**<u>Reason:</u>**

- The more stable half-filled shell of group VA elements rejects electrons, leading to a lower EA.

| | IVA | VA |
|---|---|---|
| **Period 2** | **C**<br>**[He]** ↑↓ (2s) ↑ ↑ __ (2p) | **N**<br>**[He]** ↑↓ (2s) ↑ ↑ ↑ (2p) |
| **Period 3** | **Si**<br>**[He]** ↑↓ (3s) ↑ ↑ __ (3p) | **P**<br>**[He]** ↑↓ (3s) ↑ ↑ __ (3p) |

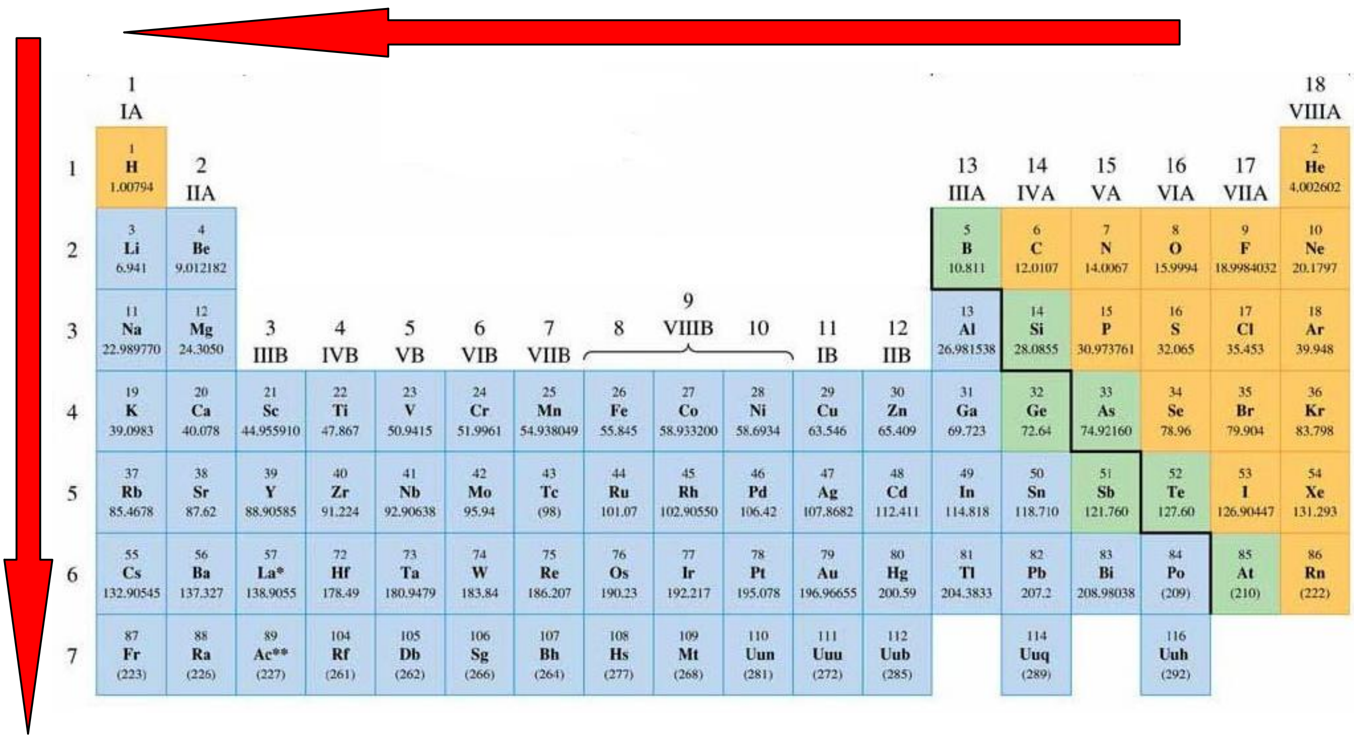
## SUMMARY OF TRENDS IN PERIODIC PROPERTIES



# METALLIC CHARACTER

- The physical and chemical properties of the main-group elements clearly display periodic character.
- The classification of elements in the periodic table into metals, non-metals and metalloids is based on fundamental periodic trends.
- The more an element exhibits the physical and chemical properties of metals, the greater its **metallic character**.

**Metallic character increases**




- The basic-acidic behavior of the oxides of elements is a good indicator of their metallic character.
- Oxides are classified as basic or acidic based on their reactions with acids and bases.
- A basic oxide is an oxide that reacts with acids. Most metal oxides are basic.
- An acidic oxide is an oxide that reacts with bases. Most non-metal oxides are acidic oxides.
- An amphoteric oxide is an oxide that has both basic and acidic properties.