CHAPTER 6

The Periodic Table & Periodic Law

6.1 Developing a Periodic Table

The periodic table was developed to show the properties of an element by simply looking at it's location.

In 1860, chemists agreed on a way to determine atomic mass, so early periodic tables were arranged by atomic mass.

History of the Periodic Table

In 1790, Antoine Lavoisier compiled a list of 23 known elements.

Electricity and spectrometers played a major role in the development of new elements.

In 1870, there were approximately 70 elements.

John Newlands

- In 1864, developed Law of Octaves (8th note)
- Elements were arranged by atomic mass, and noticed the properties of elements repeated every 8th element.
- Law was not generally accepted, but was basically correct because the properties of elements do repeat.

Dmitri Mendeleev

- In 1869, developed the first periodic table
- A Russian chemist that demonstrated the connection between atomic mass and elemental properties.
- Organized the elements by atomic mass, putting elements with similar properties into vertical columns
- Left blank spaces for undiscovered elements and predicted the properties of those undiscovered elements.

Dmitri Mendeleev *(cont.)*

- His periodic table was not completely correct
- Atomic masses were slightly incorrect and new elements were discovered.
- Arranging the elements by mass was placing elements in the wrong groups based on their properties.

Lothar Meyer

- A German chemist that developed a periodic table similar to Mendeleev in the same year.
- Not given as much credit as Mendeleev.
- It was very similar to Mendeleev's, but Mendeleev's periodic table was proven more useful.

Henry Mosely *(discovered proton)*

- Determined nuclear charge (atomic number)
- In 1913, he arranged the periodic table by order of atomic number instead of by atomic mass.
- Resulted in a clear periodic pattern of properties.
- Developed the Periodic Law.

The Periodic Law

When the elements are arranged in order of increasing atomic number, there is a periodic pattern in their physical and chemical properties

Significance of Subatomic Particles

Protons: Define each atom

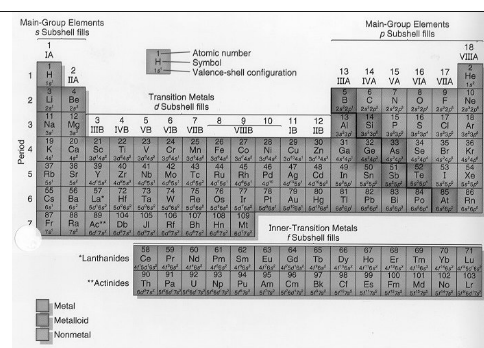
Electrons: Determines the properties of atoms

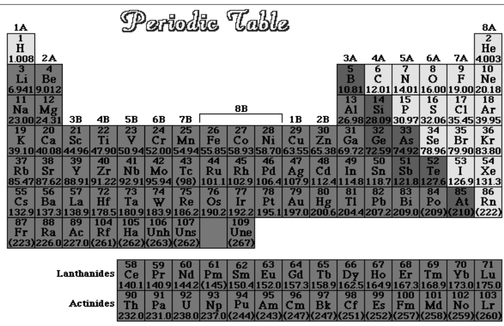
Neutrons: only alters the mass of an atom

The Modern Periodic Table

7 horizontal rows: periods
- Periods correspond to the principal energy level of atoms
- **(Principal = Outer)**

Vertical columns: groups *(aka...families)*
- Corresponds to the number of valence electrons.





Representative Elements

- Groups designated with an A, 1A - 8A.
  - > *Exhibited the **8 properties** of elements.*
- Have a wide range of chemical and physical properties
- Metals and Nonmetals are included.
- Group number determined by the valence electrons.
  - > s and p sublevels in the outermost energy level

- 1A: Alkali metals
- 2A: Alkaline earth metals
- 7A: Halogens
- 8A: Noble Gases

Transition Elements

- Groups designated with an B, 1B - 8B.
- System not as consistent as Representative elements.
- Metals only.
- Properties mainly determined by d sublevel electrons.
- Usually always have 2 valence electrons

3 Classes of Elements

- **Metals: good conductors of heat and electric current, due to mobile electrons**
- **Nonmetals: poor conductors of heat and electric current, due to immobile electrons.**
- **Metalloids: properties of both. It depends on the conditions (B, Si, Ge, As, Sb, Te, At)**

Metals

- About 80% of periodic table
- Most Group A and all Group B elements.
- High sheen and can reflect light.
- Most are ductile (make wires)
- Most are malleable and can be made into thin sheets.

(Malleable = Bend)

Alkali Metals

- s^1=1 Valence electron
- Very reactive
- Rarely found alone (always bonding)

Group 1A

Alkaline Earth Metals

- Found in the crust of the earth.

s^2 = 2 valence electrons

Group 2A

Transition Metals

- **Group B Elements**
- **Can bond with almost any non-metal because they are willing to give up many electrons.**
- ***Electron configuration for all very similar*:**
  - **Iron = [Ar] $4s^2\ 3d^6$**
  - **Zirconium = [Kr] $5s^2\ 4d^1$**

Transition Metals *(cont.)*

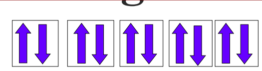
Sc: [Ar] $4s^2\ 3d^1$
Fe: [Ar] $4s^2\ 3d^6$
Zr: [Kr] $5s^2\ 4d^2$

} ***Transition Metals***

- 2 valence electrons
- Their properties are determined by the electrons in the "d" sublevel.

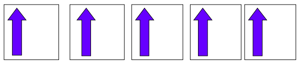
- Diamagnetic: when all of the electrons are paired *(Magnetic effects cancel)*
- Paramagnetic: contain one or more unpaired electrons
- *(Strong attraction to a magnetic field)*
- Ferromagnetism: attraction of iron, cobalt, and nickel for magnetic field
- Very similar to paramagnetic, but stronger

Diamagnetism



- **Each orbital has an electron pair.**
- **The electrons are always pulling in the opposite direction.**
- **These elements are not magnetic!!**

Paramagnetism



- All electrons are pointing in the same direction.
- They are free to move in one direction.
- Strong magnetism

Inner Transition Metals

outermost "s" sublevel and the nearby "f" sublevel generally contain electrons

- Lanthanide Series (58-71)
- Actinide Series (90-103)

NonMetals

- Most are gases at room temperature.
- Bromine is a red liquid
- Solid nonmetals are brittle.
- Have a large variation of properties

Noble Gases

- Group 8A
- Outermost "s" and "p" sublevels are completely filled
- Filled outer levels make atoms stable and inactive.

s^2p^6 = 8 valence electrons

6.2 Classifying the Elements

The electron configuration has a lot to do with an elements location on the periodic table.

Configurations of the Groups

Example: Be $2s^2$

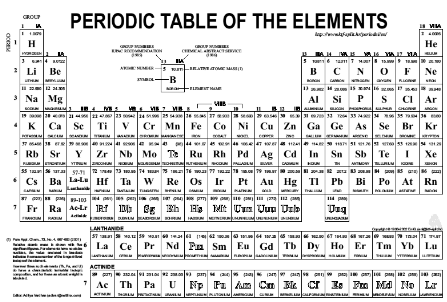
PEL																		
1	H $1s^1$																	He $1s^2$
2	Li $2s^1$	Be $2s^2$											B $2p^1$	C $2p^2$	N $2p^3$	O $2p^4$	F $2p^5$	Ne $2p^6$
3	Na $3s^1$	Mg $3s^2$											Al $3p^1$	Si $3p^2$	P $3p^3$	S $3p^4$	Cl $3p^5$	Ar $3p^6$
4	K $4s^1$	Ca $4s^2$	Sc $4s^2$	Ti $4s^2$	V $4s^2$	Cr $4s^2$	Mn $4s^2$	Fe $4s^2$	Co $4s^2$	Ni $4s^2$	Cu $4s^2$	Zn $4s^2$	Ga $4p^1$	Ge $4p^2$	As $4p^3$	Se $4p^4$	Br $4p^5$	Kr $4p^6$
			$3d^1$	$3d^2$	$3d^3$	$3d^4$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$3d^9$	$3d^{10}$						

Valence Electrons

- Electrons in "s" and "p" orbitals of an atom's outermost energy level only.
- Group number and valence electron number are related.
  > For representative elements, the number of valence electrons is equal to the group number.
- The period shows the energy level at which an atom's valence electrons are located.

Representative Elements

1A: s^1 = 1 valence electrons
2A: s^2 = 2 valence electrons
3A: s^2p^1 = 3 valence electrons
4A: s^2p^2 = 4 valence electrons
5A: s^2p^3 = 5 valence electrons
6A: s^2p^4 = 6 valence electrons
7A: s^2p^5 = 7 valence electrons
8A: s^2p^6 = 8 valence electrons



Valence Electrons *(cont.)*

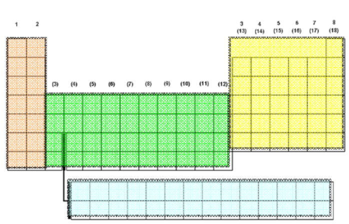
- How many valence electrons do transition metals have?
- They all have 2.
- See the electron configuration.

---

s, p, d, f blocks

The location of elements within these blocks corresponds to the four different energy sublevels.

Identifies the location of the highest energy electrons.



IONS

An ion is an atom or group of atoms that has a positive or negative charge

CATION: Positive Ions (Mg^{2+}) P=12 E=10

- Lost one or more electrons

ANION: Negative Ions (S^{2-}) P=16 E=18

-Gained one or more electrons

Why do atoms form ions?

Remember, as long as the proton numbers stays the same, the atom does not change. The electron number can change to attain the electron configuration of a noble gas.

Movement of Electrons

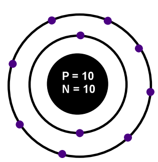
1A: 1 lost easily
2A: 2 lost easily
3A: 3 lost easily
5A: 3 gained
6A: 2 gained
7A: 1 gained

Max Loss: 4

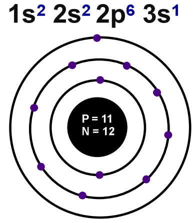
Max Gain: 3

Neon vs. Sodium

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

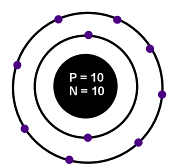


$1s^2\ 2s^2\ 2p^6\ 3s^1$

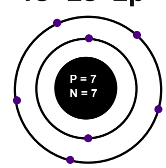


Neon vs. Nitrogen

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



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



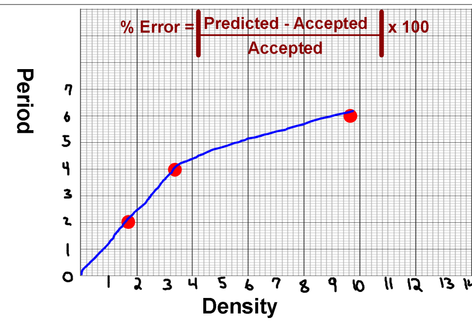
Charges of Atoms

1A: 1^+
2A: 2^+
3A: 3^+
5A: 3^-
6A: 2^-
7A: 1^-

Ionic Charges

Noble Gas	Element	Charge
He: $1s^2$	**Ca:** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$	
Ne: $1s^2\ 2s^2\ 2p^6$	**N:** $1s^2\ 2s^2\ 2p^3$	
Ar: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	**Al:** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	
Kr: [Ar] $4s^2\ 4p^6$	**S:** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$	

Element	Abbreviated Configuration	Valence Electrons	Element	Abbreviated Configuration	Valence Electrons
Mg			Br		
Ne			Ge		
Te			I		



Group 7A Densities

F: 1.7
Br: 3.12
At: 9.87

Accepted Value

Cl = 3.2

Soluble: will dissolve

Insoluble: will not dissolve

Group 2A Elements	SO_4^{2-}	CO_3^{2-}	$C_2O_4^{2-}$	CrO_4^{2-}
Mg^{2+}	*S*	*I*	*I*	*S*
Ca^{2+}	*I*	*I*	*I*	*S*
Sr^{2+}	*I*	*I*	*I*	*I*
Ba^{2+}	*I*	*I*	*I*	*I*

6.3 PERIODIC TRENDS

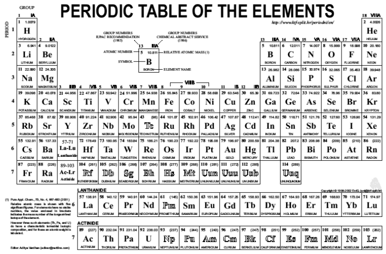
- The properties of all atoms follow trends.
- These trends can be easily understood by using the periodic table.
- Trends go according to groups and periods.
- You must know all 7 trends!!!

1. Nuclear Charge

- Based on Proton Number
- Increases as atomic number increases.
- As proton number increases, the overall positive charge of the nucleus increases.
- Periodic Trend: increases from left to right
- Group Trend: increases down the group

2. Shielding Electrons

- Electrons between the nucleus and the valence electrons.
- Shield the valence electrons from the pull of the nucleus.
- You can use the electron configuration.
- Total electrons - valence electrons
- Periodic Trend: increases from left to right
- Group Trend: increases down the group *(more energy levels)*



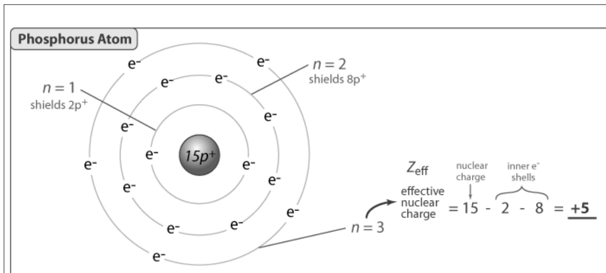
Shielding Effect

- The further an energy level is from the nucleus, the more it is shielded from the attraction to the nucleus.
- The outer energy level is shielded by the inner energy levels.
- This is one reason why atoms with more energy levels are larger and more easily lose electrons.

3. Effective Nuclear Charge (ENC)

Nuclear Charge (Protons) - Shielding Electrons

- The nuclear charge that the valence electrons actually feel.
- As the effective nuclear charge increases, the size of the atom decreases because the nucleus has more pulling power on the outer energy level of the atom.



Shielding Electrons

Total Electrons – Valence Electrons

- Sodium:
  - Nuclear Charge = 11
  - Valence Electrons = 1
  - Shielding Electrons = 10

ENC = Nuclear Charge - Shielding

ENC = __________

- Sulfur:
  - Nuclear Charge = 16
  - Valence Electrons = 6
  - Shielding Electrons = 10

ENC = Nuclear Charge - Shielding

ENC = __________

Shielding Electrons

Total Electrons – Valence Electrons

- Selenium:
  - Nuclear Charge = 34
  - Valence Electrons = 6
  - Shielding Electrons = 28

ENC = Nuclear Charge - Shielding

ENC = __________

- Vanadium:
  - Nuclear Charge = 23
  - Valence Electrons = 2
  - Shielding Electrons = 21

ENC = Nuclear Charge - Shielding

ENC = __________

Effective Nuclear Charge

Fluorine Atom:

- 9 protons & 9 electrons
- Valence electrons: 7
- Shielding electrons: 2
- Effective Nuclear Charge = +7

  Nuclear Charge (9) - Shielding Electrons (2)

Effective Nuclear Charge

Carbon Atom:

- 6 protons & 6 electrons
- Valence electrons: 4
- Shielding electrons: 2
- Effective Nuclear Charge = +4

  Nuclear Charge (6) - Shielding Electrons (2)

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$
Potassium	**Arsenic**
Total Electrons =	Total Electrons =
Valence Electrons =	Valence Eelctrons =
Shielding Electrons=	Shielding Electrons =
ENC =	ENC =

Nitrogen Atom

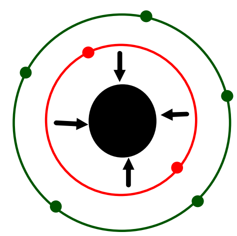
Nuclear Charge:______

Total Electrons:______

Valence Electrons: _____

Shielding Electrons:_____

ENC:______

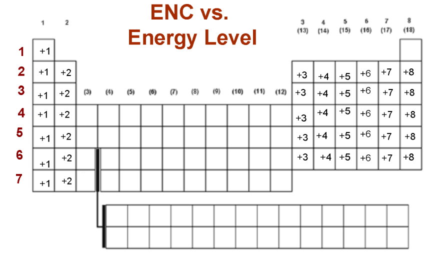


Trends in Effective Nuclear Charge

- ENC = # Valence Electrons
- The greater the ENC, the more the nucleus pulls on the outer energy level.
- **Periodic Trend**: increases from left to right
- **Group Trend**: decreases down the group because as the number of energy levels increase, the pull decreases

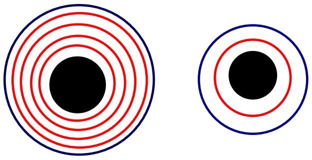
  Although ENC # stays the same!

ENC vs. Energy Level



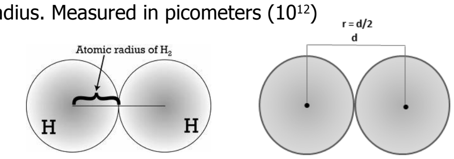
What if 2 atoms have the same ENC?

Large atoms have more shielding electrons and are further away from nucleus. So, there is less pull on the valence electrons of a large atom than a small atom.



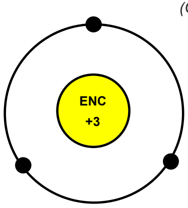
4. Atomic Size/Radius

- Radius of an atom cannot be measured directly because you cannot see the outer electrons to measure the radius.
- Half the distance between the nuclei of <u>two like atoms</u> is the atomic radius. Measured in picometers (10^{12})

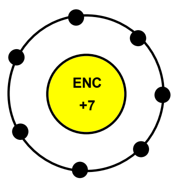


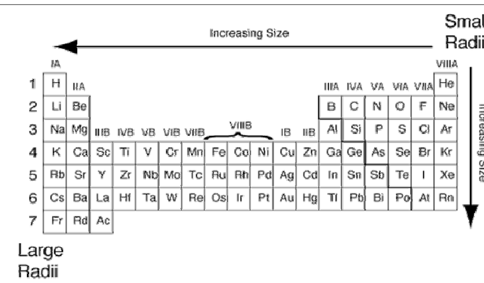
Aluminum vs. Chlorine

(Outer Energy Level Only)



Both atoms have 3 energy levels, but chlorine's outer level shrinks due to the high ENC.





Trends in Atomic Size/Radius

- **<u>Periodic Trend</u>**: decreases from left to right due to ENC increase. Large ENC makes a small atom because there is more pull on outer energy level.
- **<u>Group Trend</u>**: increases down the group because more energy levels increases the size of the atom and there are more shielding electrons.

Atomic Size/Radius *(cont.)*

Comparing 2 atoms in the same group

- Atoms with more energy levels are larger, regardless of ENC.

Iodine	Fluorine
ENC = +7 5 energy levels	ENC = +7 2 energy levels

Atomic Size/Radius *(cont.)*

Comparing 2 atoms in the same period

- High ENC atoms shrink because the outer level is pulled very close to the nucleus.
- Low ENC atoms are large because the nucleus has little pulling power on the outer energy level.

2 Factors determine most trends

1. ENC (Periodic Trends)

- Size
- Electronegativity
- Ionization Energy

+1 Weak Pull

+7 Strong Pull

2. Principal Energy Level (Group Trends)

- Size
- Shielding Electrons
- Attraction of valence electrons to the nucleus

5. Ionic Size/Radius

- When atoms lose electrons, they get smaller
  - > Less electrons, more pull from nucleus
- When atoms gain electrons, they get bigger.
  - > More electrons, less pull from nucleus

Magnesium	Sulfur
$1s^2\ 2s^2\ 2p^6\ 3s^2$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
Will lose 2 electrons.	Will gain 2 electrons.
Mg^{2+}	S^{2-}
Atom shrinks	Atom expands
Protons vs. Electrons	Protons vs. Electrons
12 vs. 10	16 vs. 18

CATIONS

- Smaller than neutral version.
- Positively Charged
- Lose Electrons
- Metals
- The more they lose, the smaller they get.
- They lose an energy level which decreases shielding.
- Electrons lose the "tug of war" with the protons.

TRENDS IN IONIC SIZE

- **Periodic Trend**: cations decrease from left to right because more electrons are lost.
- **Group Trend**: cations increase down the group because atoms within a group lose the same number of electrons, but increase the number of energy levels.

ANIONS

- Larger than neutral version.
- Negatively Charged
- Gain Electrons
- Nonmetals
- The more electrons gained, the bigger the ion gets.
- Electrons win the tug of war with the protons.

TRENDS IN IONIC SIZE

- **Periodic Trend**: Anions decrease from left to right because fewer electrons are gained.

- **Group Trend**: anions increase down the group because atoms within a group gain the same number of electrons, but increase the number of energy levels.

Which is a smaller ion?

Boron: $1s^2\ 2s^2\ 2p^1$

Aluminum: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

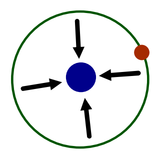
Which is larger?

Oxygen: $1s^2\ 2s^2\ 2p^4$

Oxygen Ion: $1s^2\ 2s^2\ 2p^6$

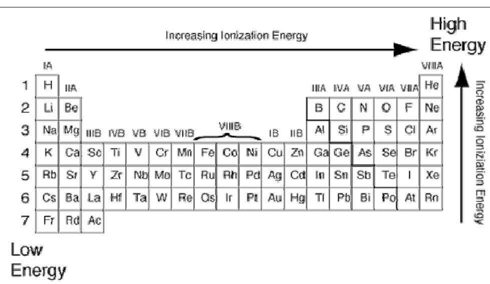
6. TRENDS IN IONIZATION ENERGY

- The energy that is required to overcome the attraction of the nuclear charge and remove an electron from an atom.



Weak Pull = Low ENC = **Low IE** Easy to Remove

Strong Pull = High ENC = **High IE** Hard to Remove



TRENDS IN IONIZATION ENERGY

3 Things determine whether electrons can be removed...

1. ENC (effective nuclear charge)
2. The number of energy levels
3. The Tug of War after electrons have been removed

Trends in Ionization Energy

- Increases right and up.
- **Group Trend**: Decreases as you go down because it is harder to remove electrons from atoms with outer energy level close to nucleus and easier to remove electrons from outer energy levels further from the nucleus.
- **Periodic Trend**: increase from left to right because it is harder to remove electrons when the ENC is large and easier to remove electrons when the ENC is small.

1st IONIZATION ENERGY

- The energy required to remove the first electron from an atom.
- Depends on how many electrons exist in the outer energy level.

1st, 2nd, & 3rd Ionization Energy

Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$	*Remove 1*	*Remove 2*	*Remove 3*
Mg: 1s2 2s2 2p6 $3s^2$	*Remove 1*	*Remove 2*	*Remove 3*
Al: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	*Remove 1*	*Remove 2*	*Remove 3*

2nd IONIZATION ENERGY

Energy required to remove the second electron from an atom.

Depends on how many electrons exist in the outer energy level.

It always requires more energy to remove the second electron than it does to remove the first.

Test Question:
Which has highest 2nd IE? Sodium, Bromine, or Oxygen
Answer: Sodium, because it becomes a noble gas after losing 1 electron.

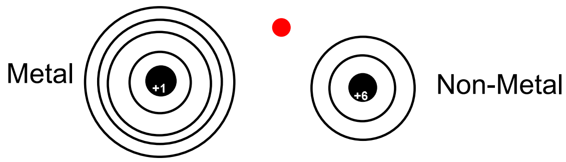
7. ELECTRONEGATIVITY

- The tendency of atoms to attract electrons when they are chemically combined with another element.
- Electrons within a bond are more attracted to atoms with a high electronegativity
- Each element is assigned an electronegativity value

Trends in Electronegativity

- **Periodic Trend**: increases across the periodic table because ENC increases
- **Group Trend**: decreases as you move down a group because atoms get larger and ENC decreases

Which atom will attract the free electron?



Which has the higher electron affinity?

Know these 7 Trends

1. **ENC**
2. **Nuclear Charge**
3. **Shielding Electrons**
4. **Atomic Size *(Atomic Radii)***
5. **Ionic Size**
6. **ionization Energy**
7. **Electronegativity**

What makes an atom large? 1. 2.	What makes an atom small? 1. 2.
What gives an atom a strong ENC? 1. 2.	**What gives an atom a weak ENC?** 1. 2.
What give an atom High Electronegativity? 1. 2.	**What gives an atom Low Electronegativity?** 1. 2.
Which atoms have High Ionization Energy? 1. 2.	**Which atoms have Low Ionization Energy?** 1. 2.