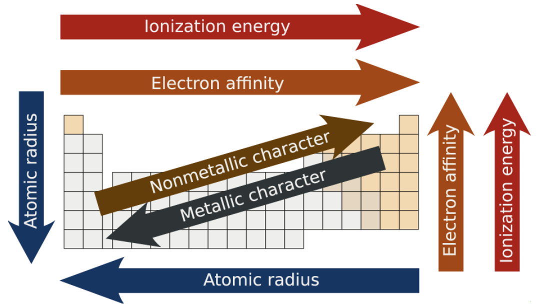
Ch08

Periodic Trends

Exploring electron configurations

and the properties they produce.



version 1.5

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Exploring Trends in Electronic Structure

- Atomic Radius
    - Non-Bonding Radius vs Bonding Radius
  - Trends
    - Across Periodic Table
    - Down Periodic Table
    - Transition Metals
- Magnetism
- Ions
  - Making Cations
    - Main Group vs Transition Metals
    - Electron Configurations
    - Size
    - Ionization Energy
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    - Electron Configurations
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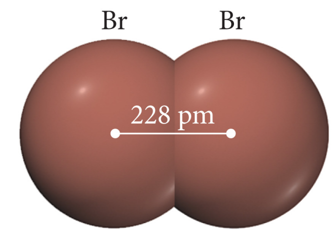
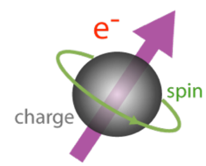
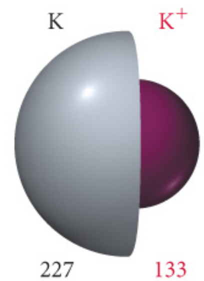
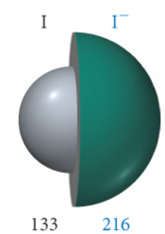
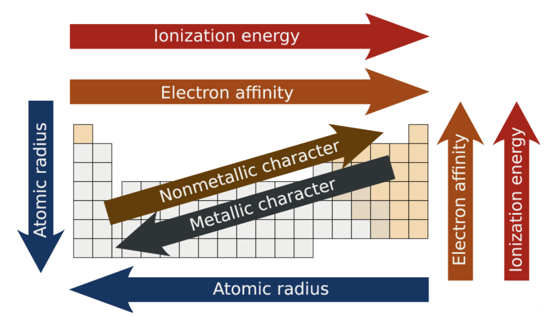
    - Electron Affinity
- Metallic Character
- Patterns in Chemical Reactivity





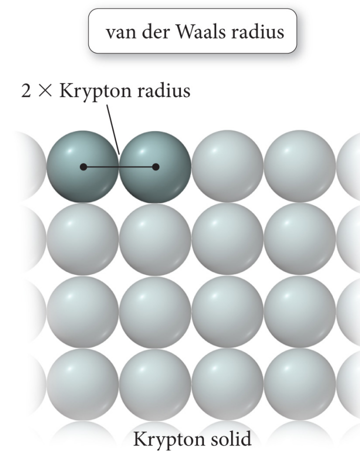


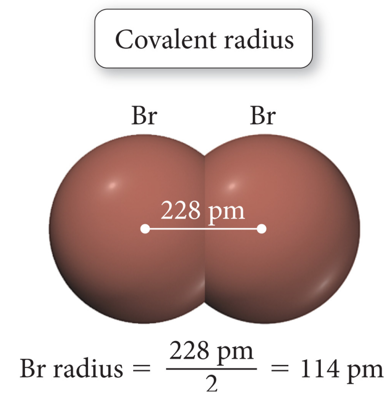


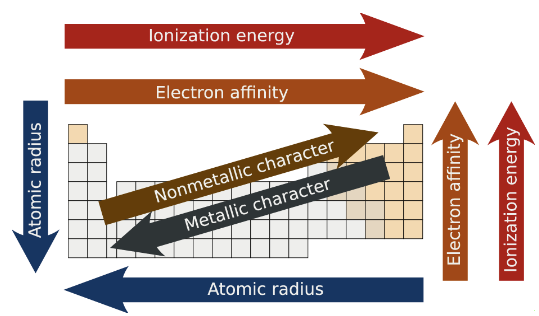


The size of Atoms

- The electron cloud of an atom defines it's size.
- How do you measure the size of a cloud?
  - The edges of a cloud are uncertain.
- We measure the distance between two adjacent atoms.
  - Atomic size is measured by packing atoms close together, finding the distance between adjacent nuclei, and dividing that number by 2.
  - Atoms can be packed densely by capturing them in solid form or capturing them in a another compound that is a solid.
    - These atoms are not bonded, their electron orbitals don't mix.
    - We describe the atomic size we get from this process as the nonbonding atomic radius or van der Waals radius.
  - For metals this involves analyzing metallic crystals (atoms held together with metallic bonds).
  - For non-metals, we look at a large number of compounds that contain the element.
    - We look at the average bond length between atoms.
    - There is overlap between the electron orbitals.
    - We describe the atomic radius found from covalently bonded compounds as the bonding atomic radius or covalent radius.
  - Which atomic radius we use depends on the context.
    - When we say atomic radius, we more often mean bonding atomic radius.
  - We can determine the relative atomic radius of two elements by their position in the periodic table.

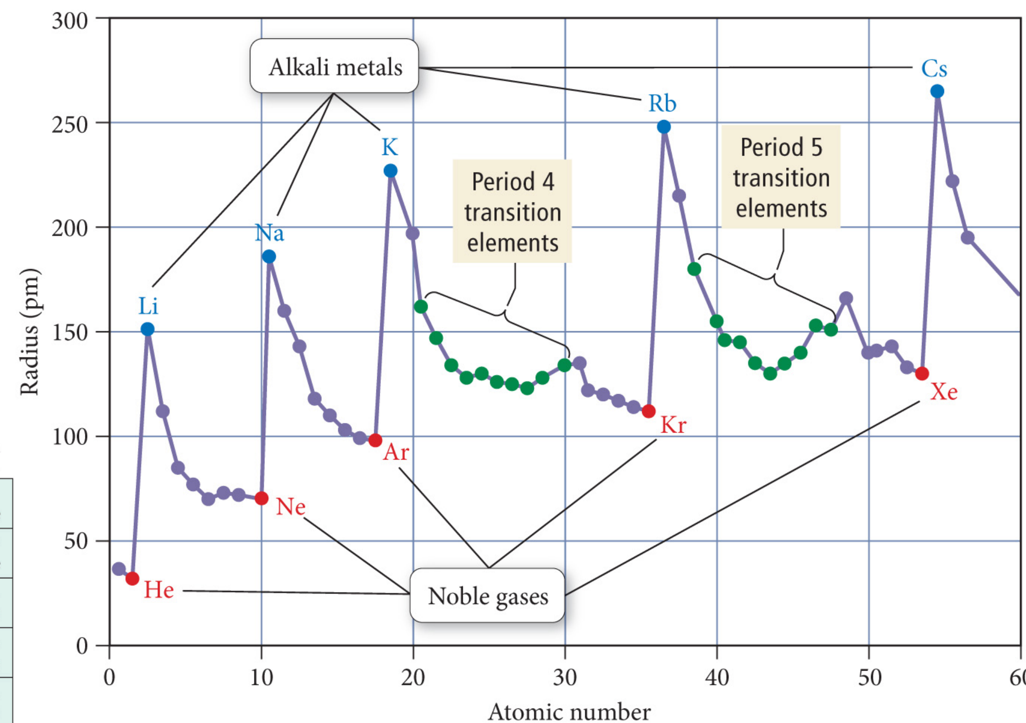


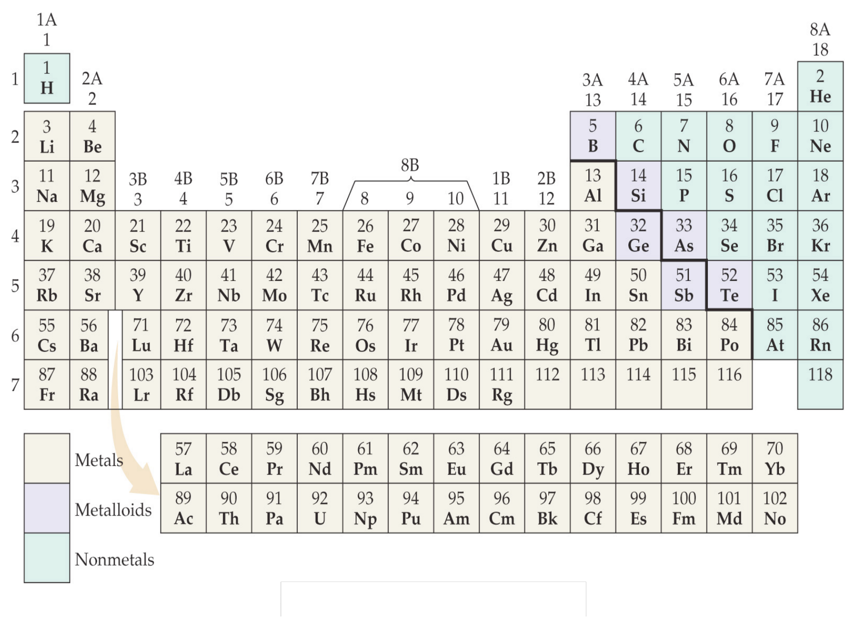


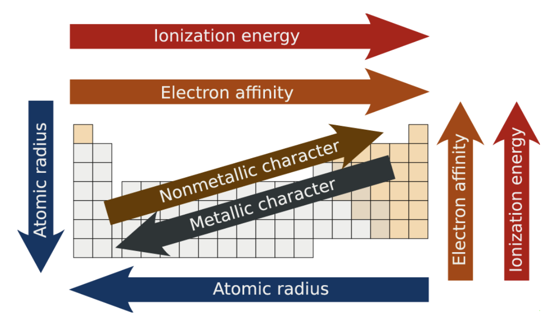


Relative Atomic Radius

- In general, as we move across the periodic table left to right the atomic radius decreases.
  - Transition metals of the same period are *roughly* the same size.
- In general, as we move down the periodic table the atomic radius increases.

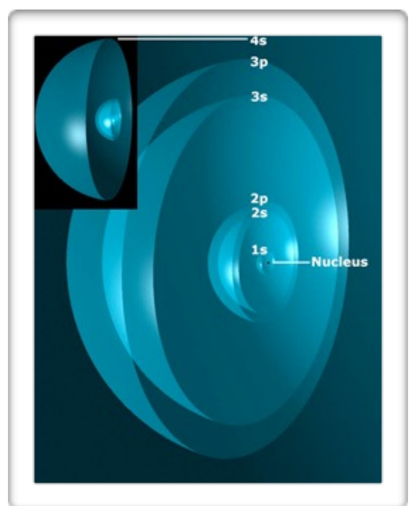


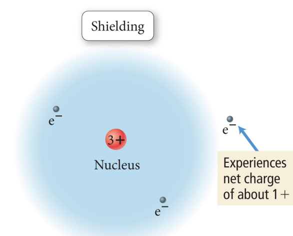


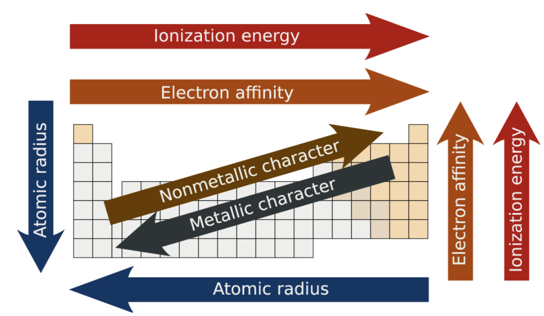


Atomic Size Increases as we go Down

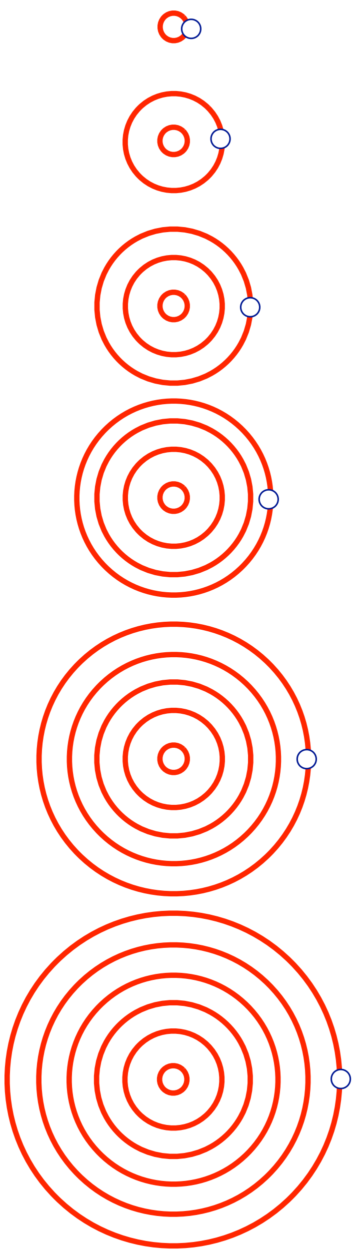
- Atomic Radius increases as we move down the periodic table because each period represents a new electron shell.
  - Each shell shields the previous shell from the nuclear charge reducing the pull of electrons to the nucleus.
  - The electrons in the inner shells repulse the electrons on the outer shells, pushing them farther from the nucleus.
- Each period corresponds to an increase in the principle quantum number n, which describes the size of that shell.
- As we add a larger shell, a new layer to the atom, it get's bigger.



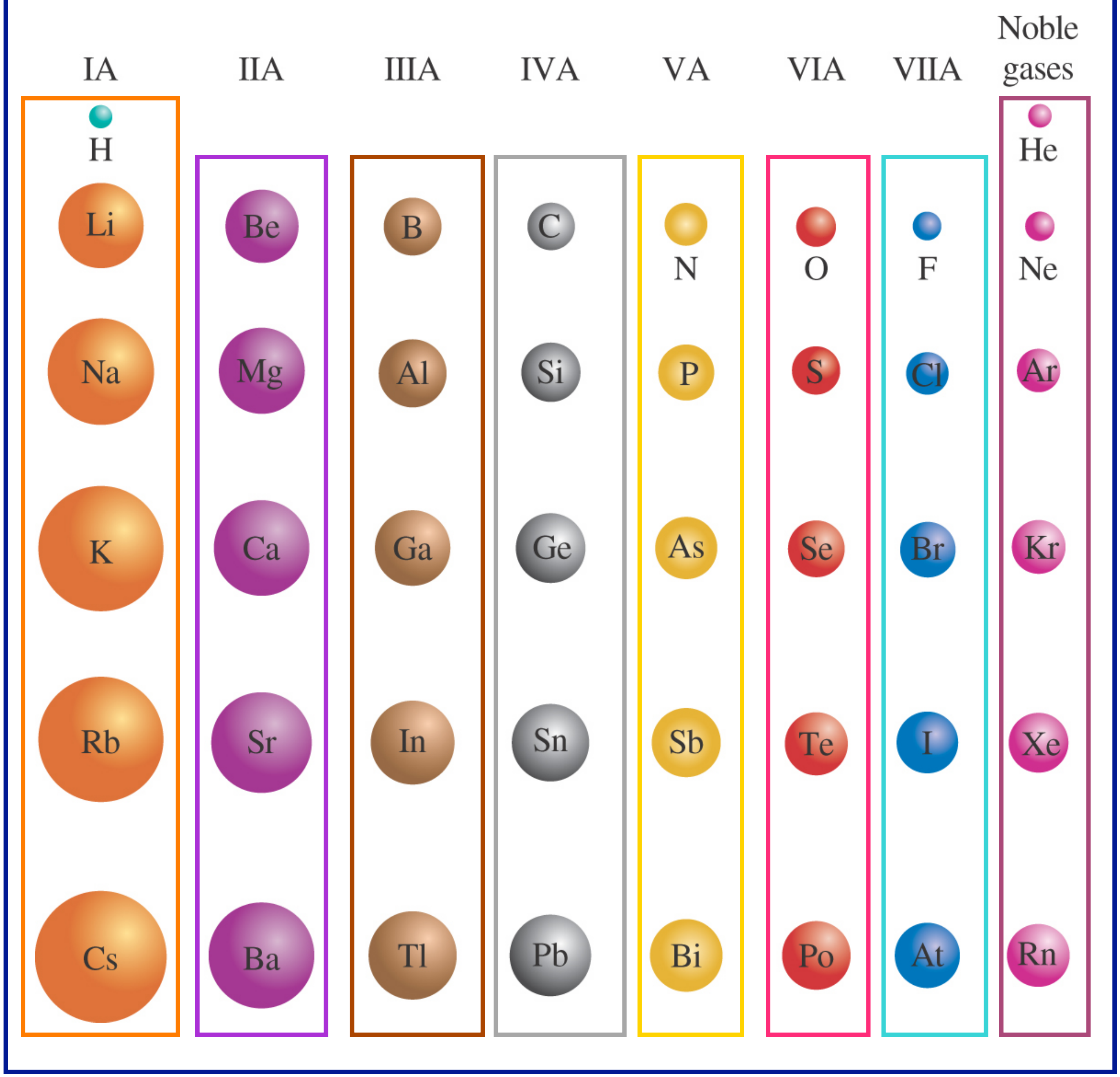




Atomic radii increase down a group.

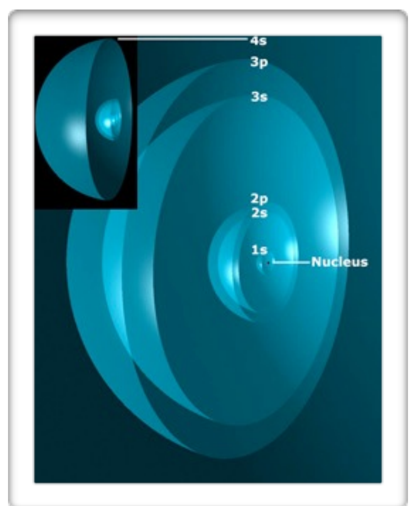


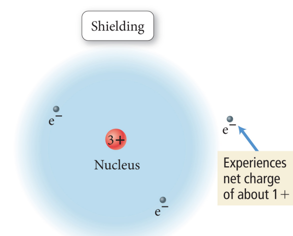
For each step down a group, electrons enter the next higher energy level.

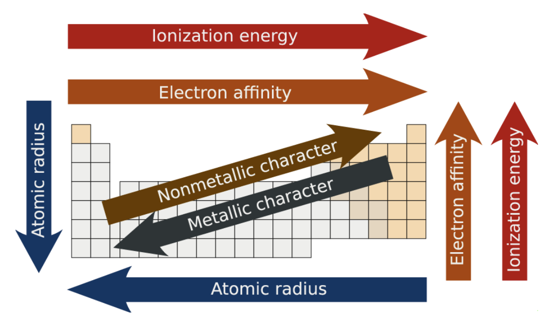


Atom Size decreases as we move Across

- Atomic Radius decreases as we move across the periodic table because the nuclear charge increases with each new row.
  - Shielding within a period is minimal.
  - Electrons have a minimal repulsion within a shell because they are in separate orbitals.
  - They fit together well.
- As the effective nuclear charge increases it nuclear attraction all electrons in that shell feel.
- It tightens the atom.
- The increased effective nuclear charge pulls all the electrons in the outer shell closer.
- It makes the atom smaller.





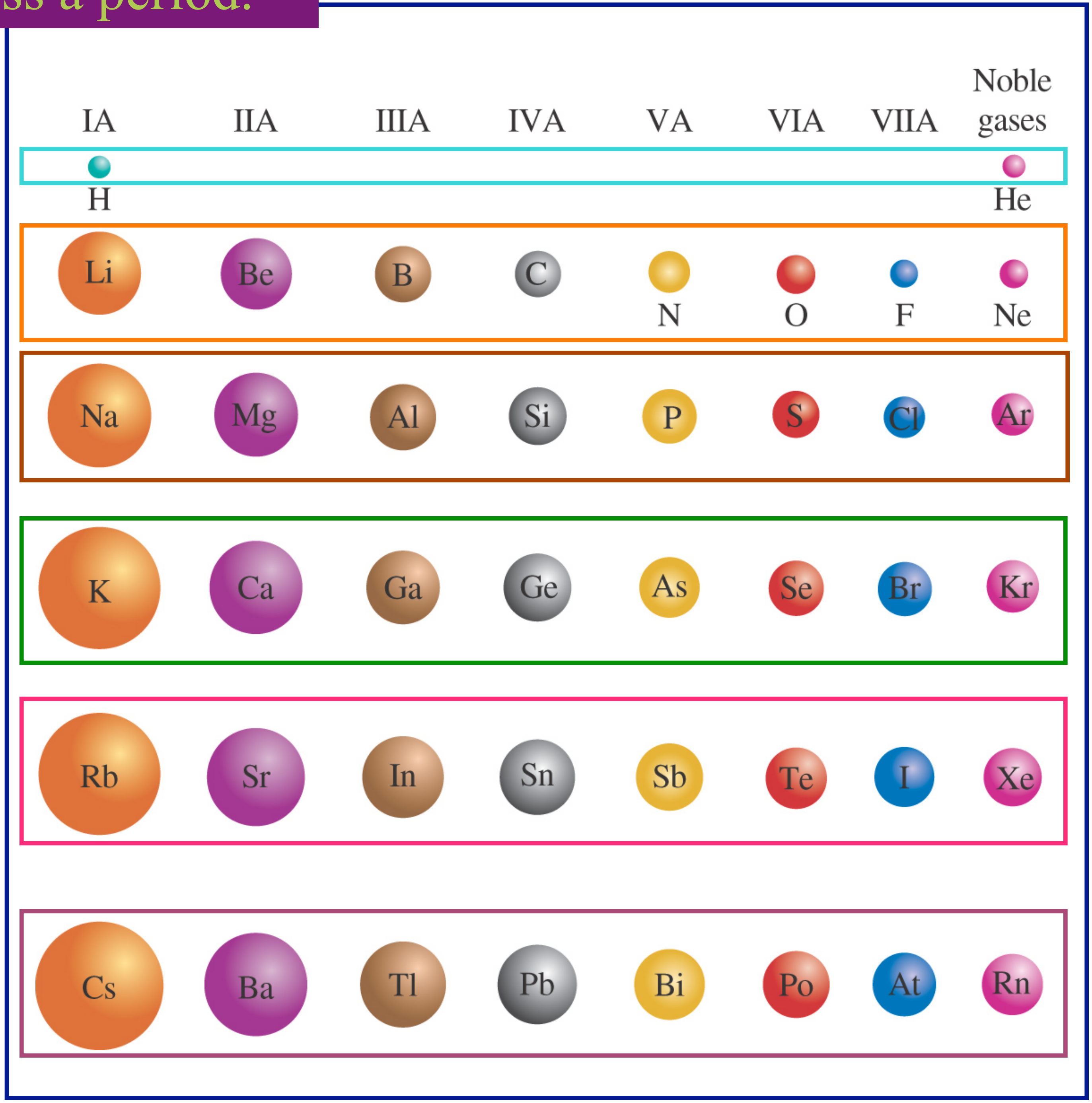


Radii of atoms tend to decrease from left to right across a period.

Each time an electron is added, a proton is also added to the nucleus.

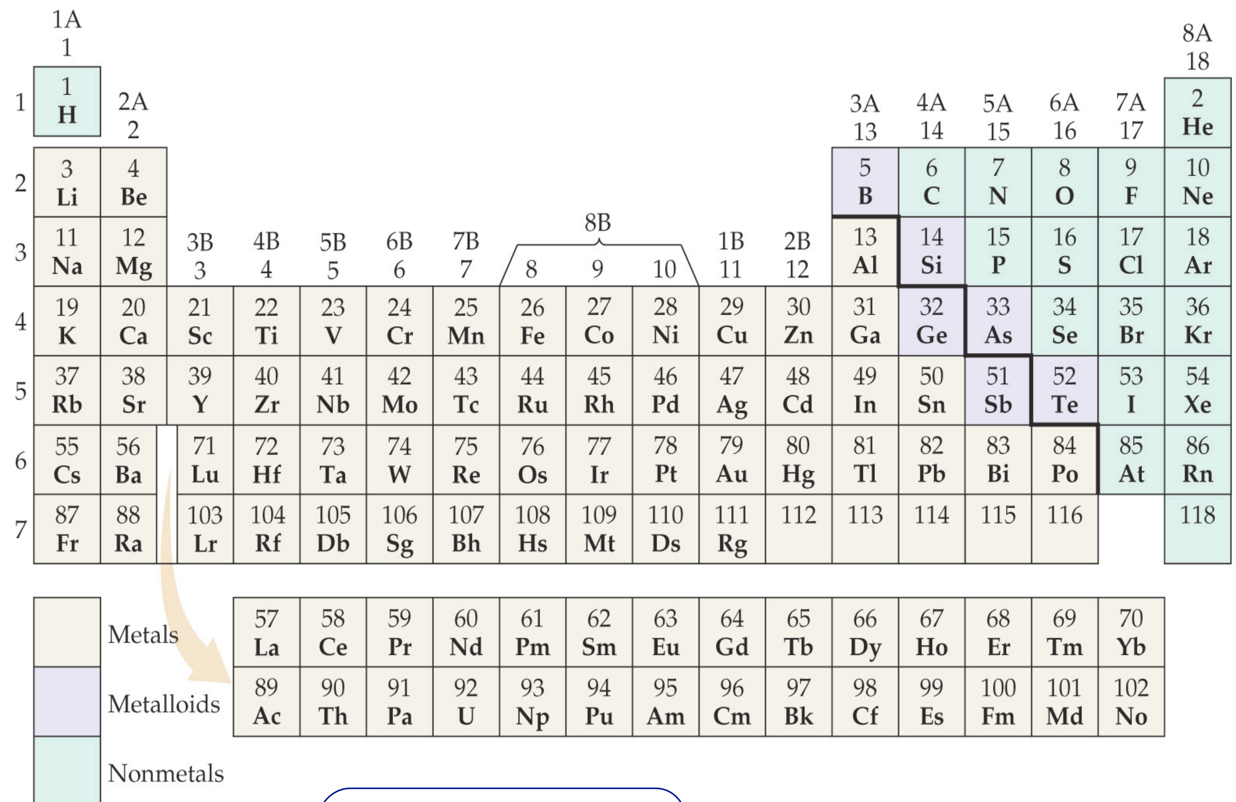
This increase in positive nuclear charge pulls all electrons closer to the nucleus.

For representative elements within the same period, the energy level remains constant as electrons are added.

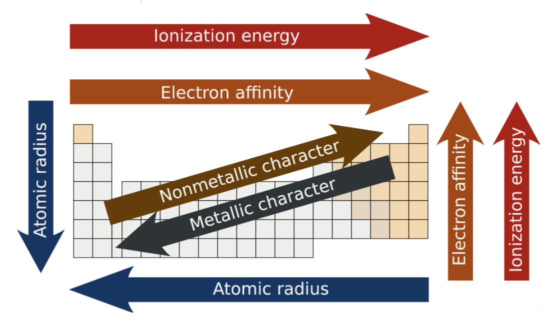


Which Atom is Larger?

- Which atom is larger?
  - F or Li — Li
  - Cl or Br — Br
  - Ge or Se — Ge
  - P or O — P
  - Al or Ge — Unclear
  - Al or Po — Po



Within the P block or within the S block, movie down the periodic table has a bigger effect than moving across it. But trust this only if the distance is two or more rows or columns.



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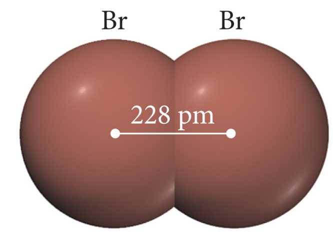
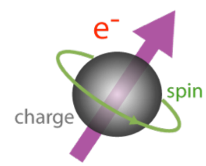
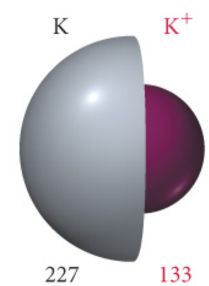
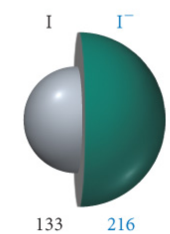
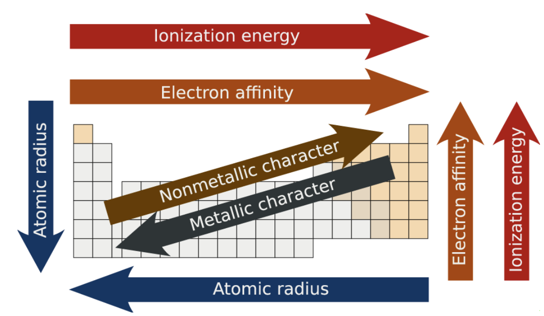




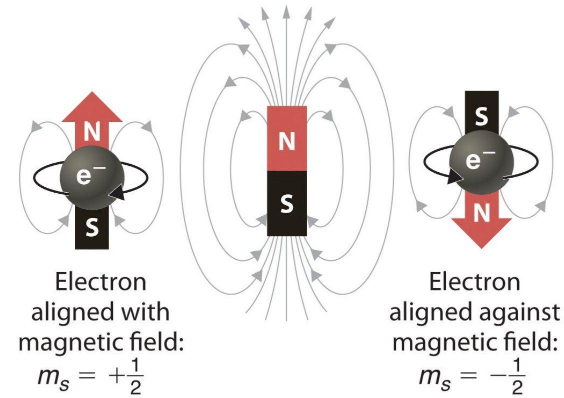






Magnetism

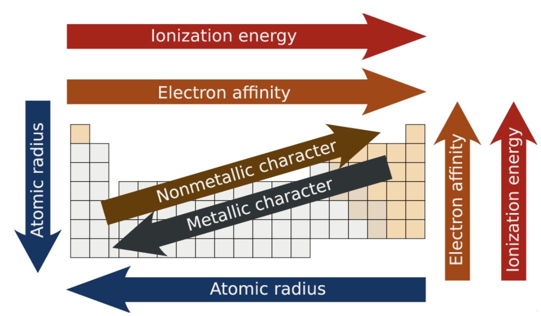
- Electron spin interacts with magnetic fields.
- In elements that have electron configurations with paired electrons the spins provide equal and opposite interactions – they cancel each other out.
  - These materials are described as diamagnetic, they can't be pulled by a magnet. (in fact they are slightly repulsed by it)



Ca $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

- Elements that have electron configurations with unpaired electrons have a net spin. And therefore a net magnetic field.
  - These materials are described as paramagnetic, they can be pulled by a magnet.

V $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$



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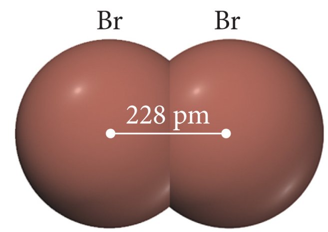
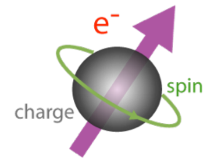
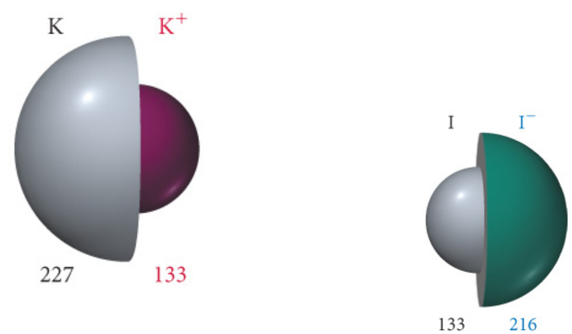
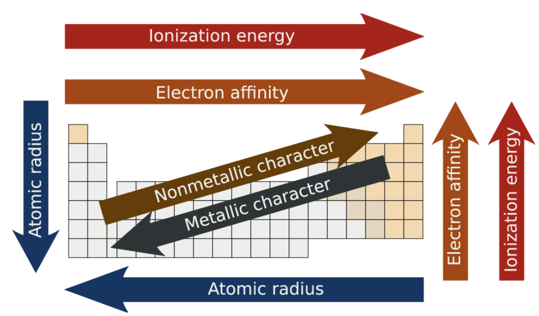




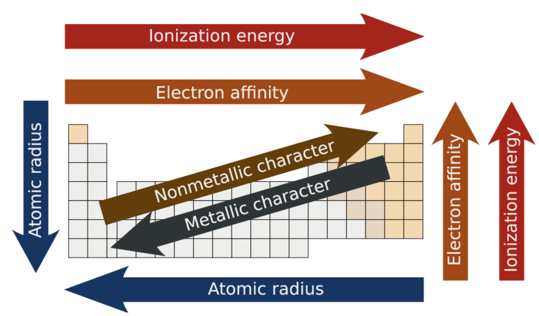




Electron Configuration of Cations

- We build up the electron configuration of a neutral atom by considering three principles:
  - Aufbau Principle "work up from the bottom"
  - Hund's Rule "don't double book unless you have to" & "align single electrons"
  - Pauli Exclusion Principle "pair spins when you double book"
- To form a cation, we start with the neutral atom, and then remove electrons.
  - For main group cations we just reverse the above process.
  - For transition metal cations, we remove electrons from the highest n-value orbitals first *even if this does not reverse the order in which they were filled*.

K	19 electrons	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	K^+	19 electrons - 1 e	$1s^2 2s^2 2p^6 3s^2 3p^6$
Al	13 electrons	$1s^2 2s^2 2p^6 3s^2 3p^1$	Al^{3+}	13 electrons - 3 e	$1s^2 2s^2 2p^6$
Mg	12 electrons	$1s^2 2s^2 2p^6 3s^2$	Mg^{2+}	12 electrons - 2 e	$1s^2 2s^2 2p^6$



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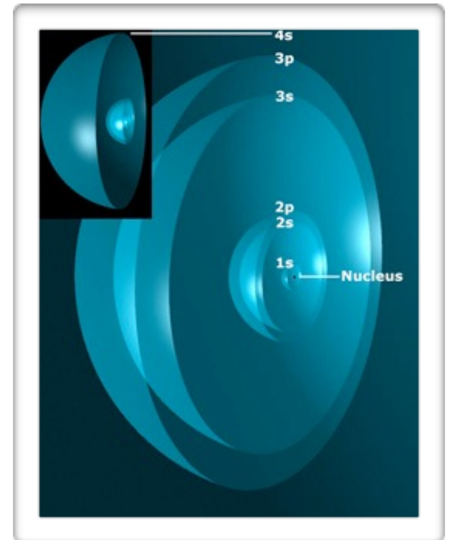
- To form a cation, we start with the neutral atom, and then remove electrons.
  - We remove electrons from the highest n state first.
  - For main group cations this just reverses the process we used to add electrons.
  - For transition metal cations this means removing electrons from s orbitals before d orbitals *even though this does not reverse the order in which those orbitals were filled.*



Ti^{2+} 22 electrons - 2 e

4s 3d

~~$1s^22s^22p^63s^23p^64s^2$~~ ↑↓ ~~diamagnetic~~

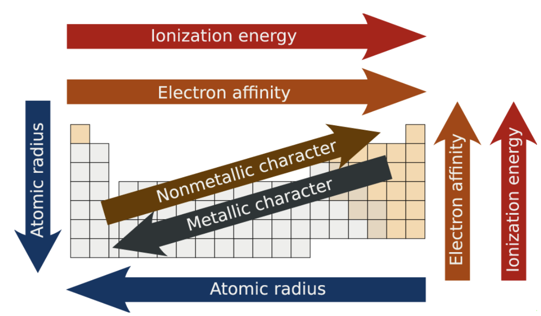
4s 3d

$1s^22s^22p^63s^23p^63d^2$ ↑ ↑ paramagnetic

Ti 22 electrons

4s 3d

$1s^22s^22p^63s^23p^64s^23d^2$ ↑↓ ↑ ↑ paramagnetic



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  - For main group cations we just reverse the above process.
  - For transition metal cations, we remove electrons from the highest n-value orbitals first *even if this does not reverse the order in which they were filled*.

Sc 21 electrons

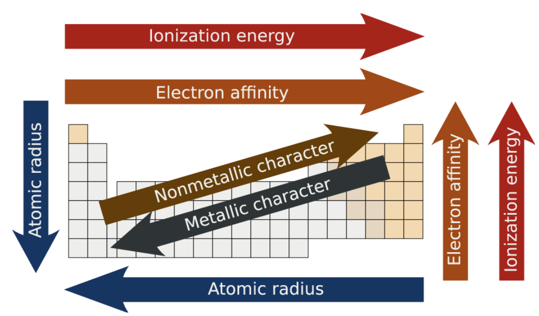
Sc^{+1}

Ni 28 electrons

Ni^{+2}

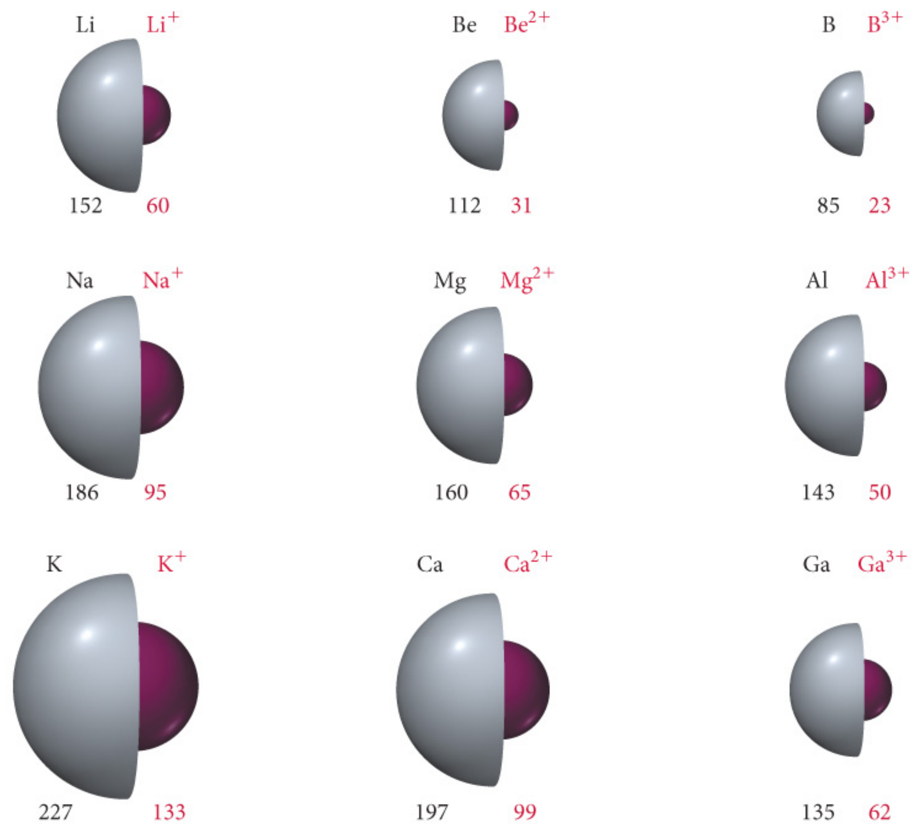
Ge 32 electrons

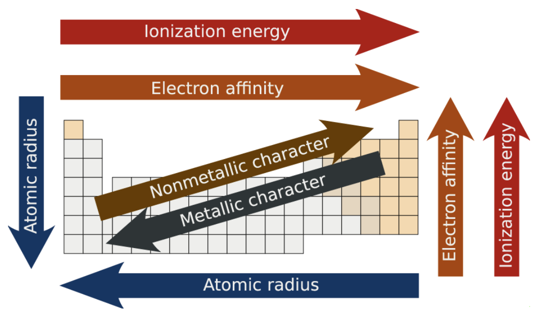
Ge^{+3}



Size of Cations

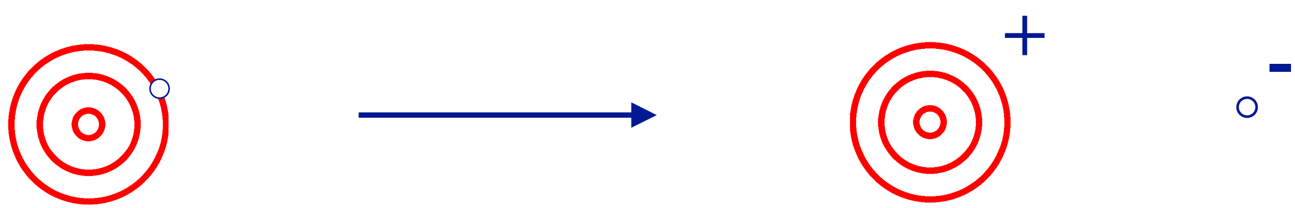
- Cations have a smaller atomic radius than the corresponding neutral atom.
- Cations are formed by striping electrons from the outermost electron shell.
- For main group elements the entire valence shell is usually removed.
- The cation therefore has an electron configuration equal to the nobel gas in the previous period.
- It’s equal to the electron configuration of an element higher in the periodic table and across the periodic table — both trends that reduce atomic radius.

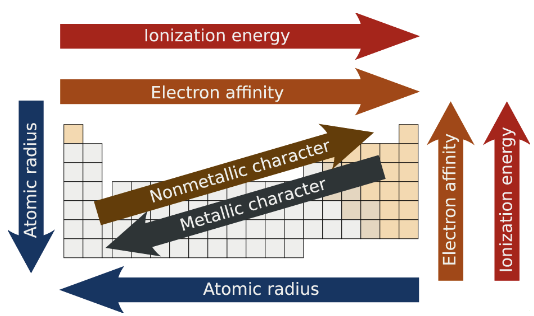




The **ionization energy** of an atom is the energy required to remove an electron from an atom.

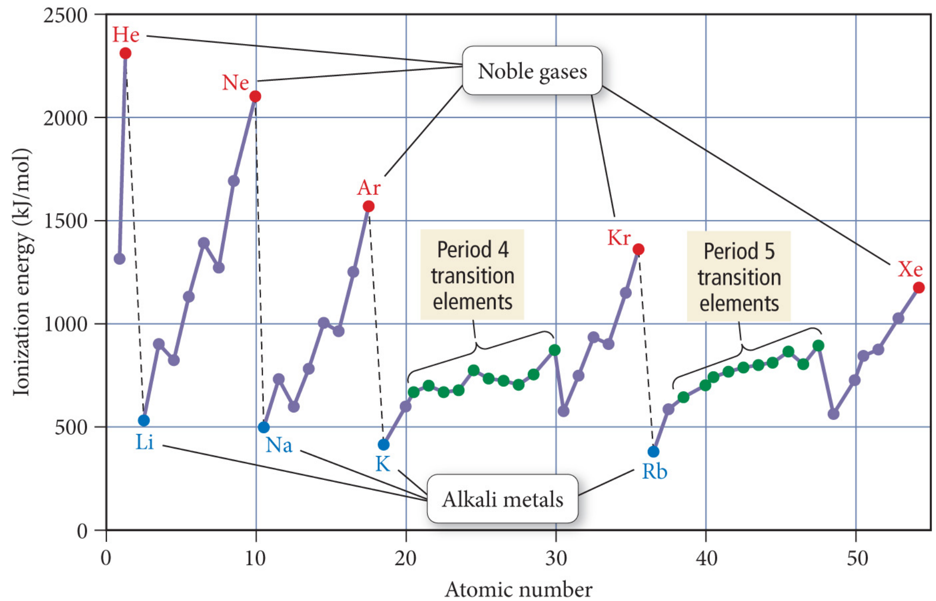
Na + ionization energy $\rightarrow Na^{+} + e^{-}$

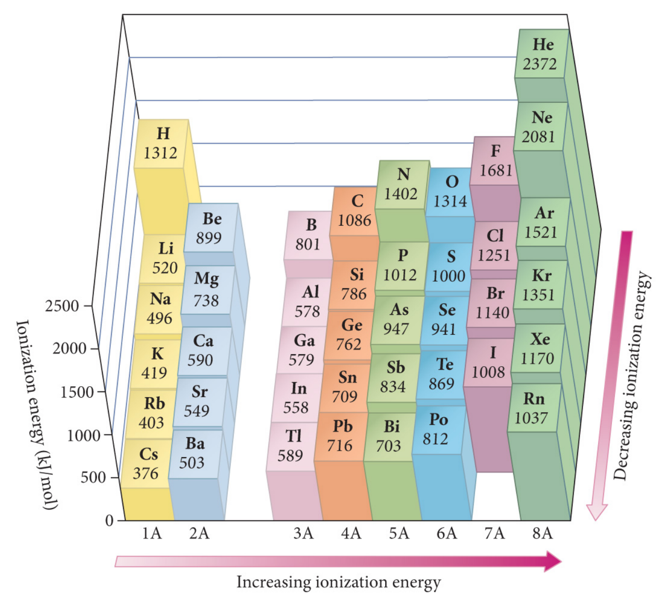




Ionization Energy

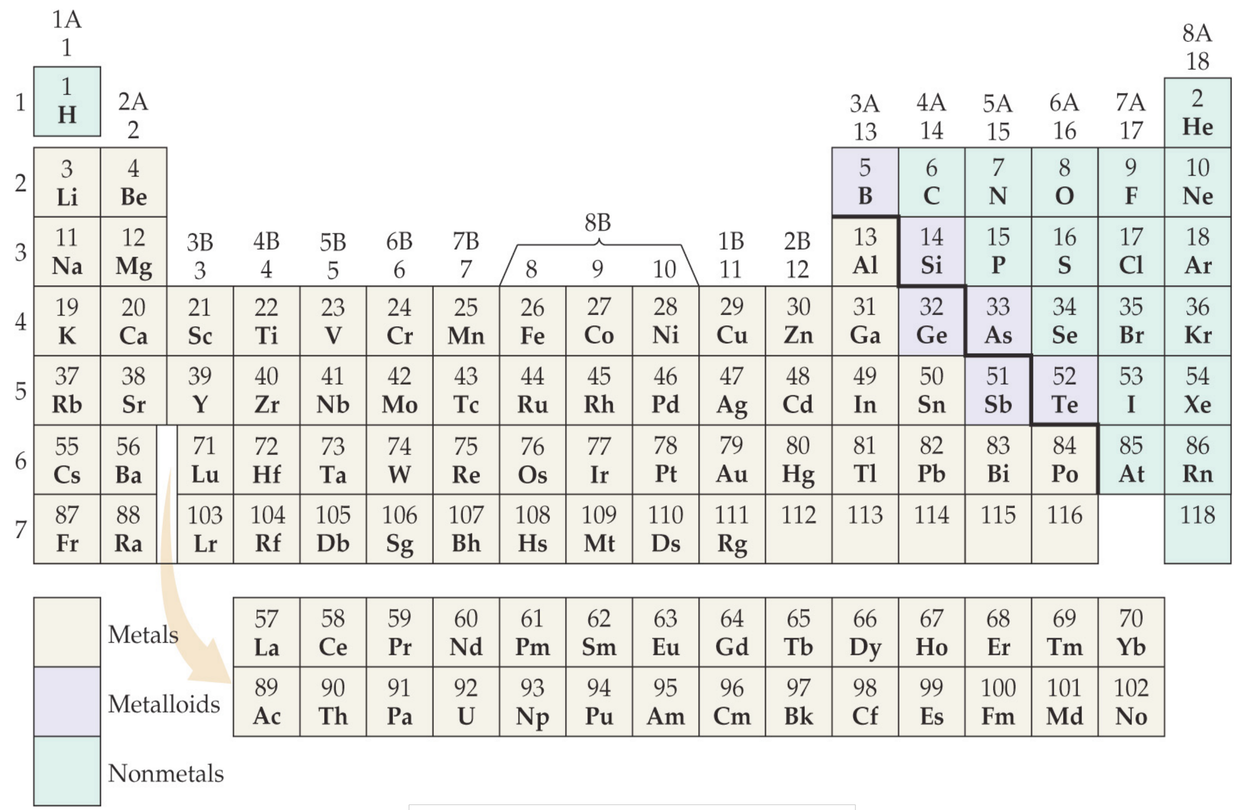
- Ionization energy is the energy required to remove an electron from an atom or ion.
- Ionization energy get's larger as you move across the periodic table from left to right.
  - As you move across the periodic table, the effective nuclear charge increases.
  - The pull on each electron in the outermost shell increases.
  - So it's harder to remove those electrons.
- Ionization energy get's smaller as you move down the periodic table.
  - As you move down the periodic table the radius of the valence shell increases.
  - While nuclear charge increases, shielding reduces the effect of that increased nuclear charge.
  - The outer electrons are held more loosely.
  - It's easier to remove electrons from these larger shells.
- Nobel gases are almost impossible to ionize.
- Of the remaining elements, Fluorine is the king, as you get farther from Fluorine it becomes easier to steel electrons.
- Hydrogen is an exception to the pattern.

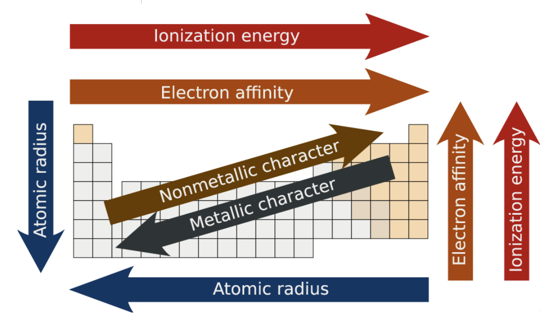




Which Element has a higher IE?

- Which has a higher IE?
  - F or Li — F
  - Cl or Br — Cl
  - Ge or Se — Se
  - P or O — O
  - Al or Ge — Unclear





Second Ionization Energy

- Second Ionization energy is the energy required to remove a second electron from a neutral atom.
- There can be dramatic differences between first, second, third and further ionization energies.
- The factors that control ionization energy change dramatically when the entire outermost shell is removed.
  - Removing electrons from an inner shell means that shell means loosing two shells of electron shielding.
  - The effective nuclear charge increases dramatically.
- The position in the periodic table can be used to predict where an element will run into this barrier.

TABLE 8.1 Successive Values of Ionization Energies for the Elements Sodium through Argon (kJ/mol)

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	496	4560					
Mg	738	1450	7730				
Al	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,100		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
Cl	1251	2300	3820	5160	6540	9460	11,000
Ar	1521	2670	3930	5770	7240	8780	12,000

Core electrons

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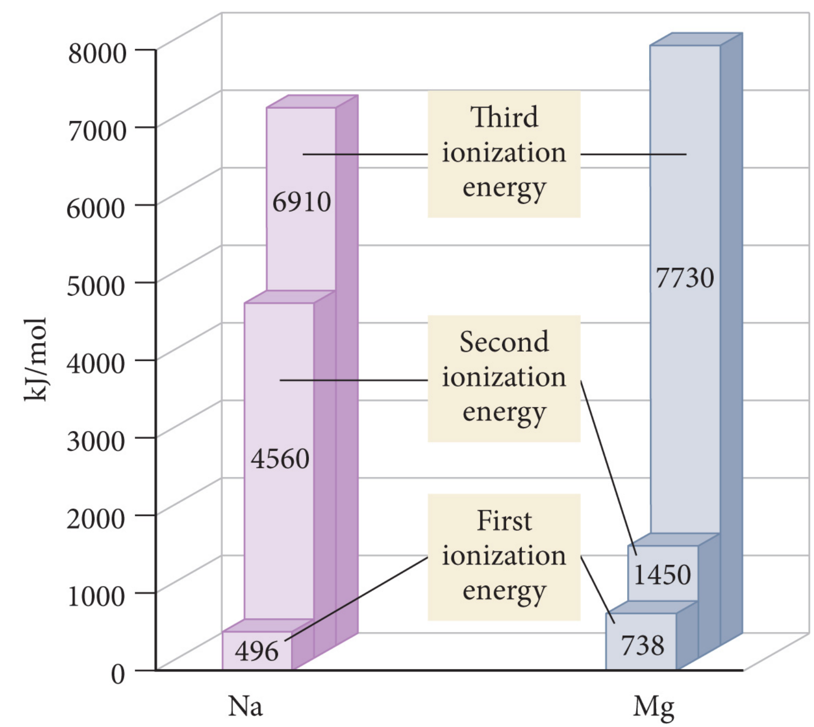

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  - The effective nuclear charge increases dramatically.
- The position in the periodic table can be used to predict where an element will run into this barrier.
- This is why we can predict the charge of cations from an elements position in the periodic table.

TABLE 8.1 Successive Values of Ionization Energies for the Elements Sodium through Argon (kJ/mol)

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	496	4560					
Mg	738	1450	7730			Core electrons	
Al	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,100		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
Cl	1251	2300	3820	5160	6540	9460	11,000
Ar	1521	2670	3930	5770	7240	8780	12,000

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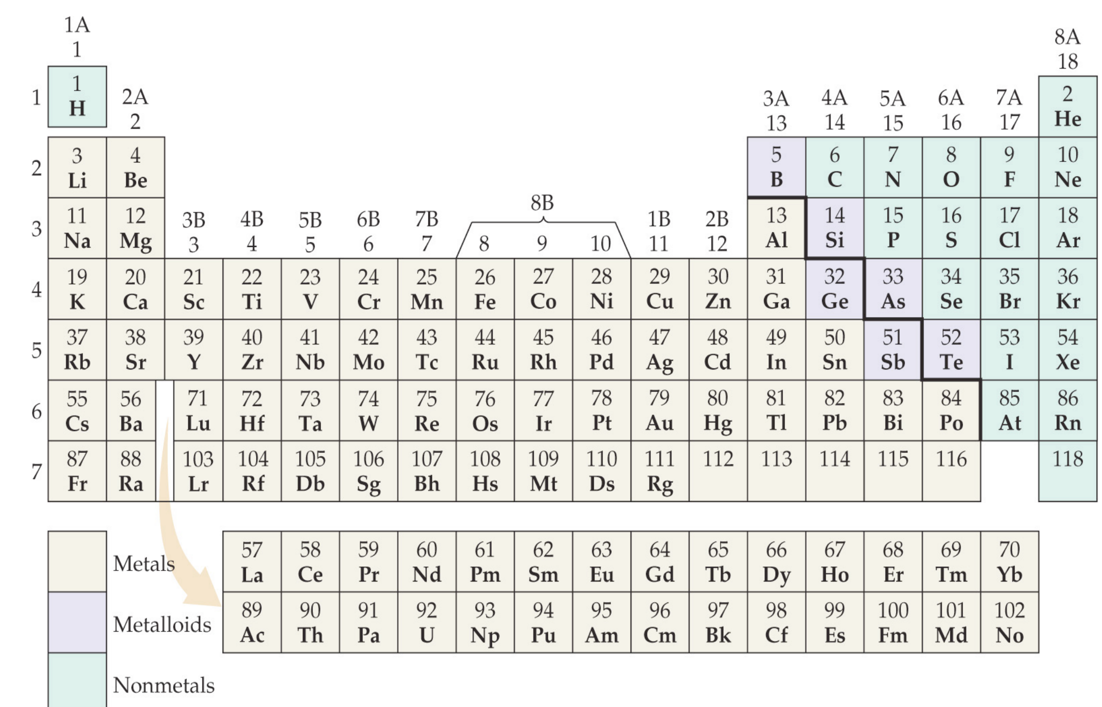


Forming Cations

- Does it take more energy to form K^+ or Na^+ from the neutral element?

  Ionization energy goes down as you move down the table.

  $\boxed{\text{Na}^+ \text{ takes more.}}$

- Does it take more energy to form Al^{3+} or Mg^{3+} from the neutral element?

  n=2 electrons take more energy.

  $\boxed{\text{Mg}^{3+} \text{ takes more}}$

- What's the electronic configuration of Ca^{2+}? Of Ca^{3+}?

  Ca: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

  Ca^{2+}: $1s^2 2s^2 2p^6 3s^2 3p^6$

  Ca^{3+}: $1s^2 2s^2 2p^6 3s^2 3p^5$

- Which has unpaired electrons, a diamagnetic or paramagnetic material?

  In diamagnetic materials all electrons are paired!

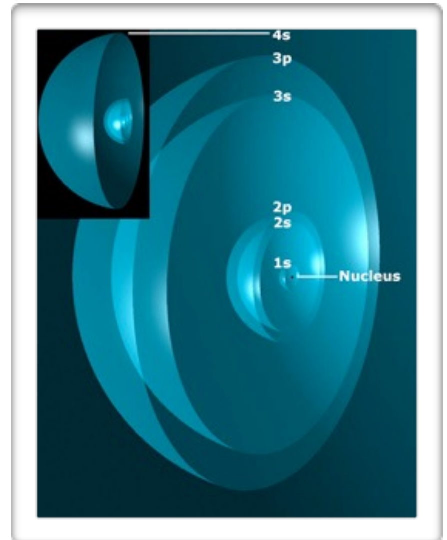
  $\boxed{\text{paramagnetic materials have unpaired electrons}}$

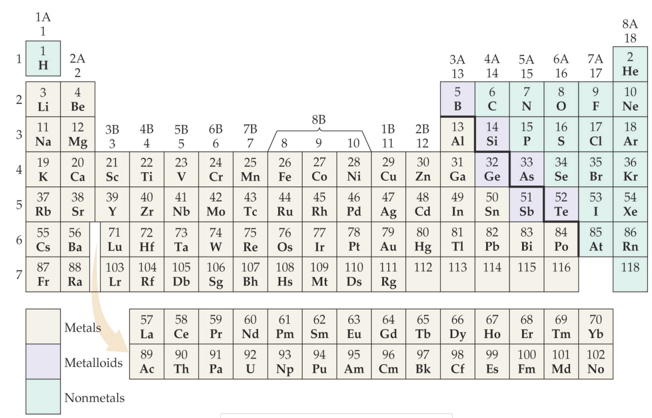
- Is V^{3+} diamagnetic or paramagnetic?

  paramagnetic

V $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

V^{3+} $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2$





Exploring Trends in Electronic Structure

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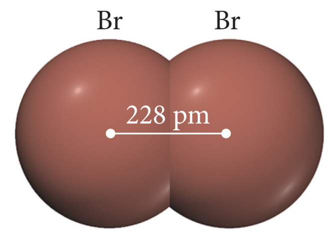
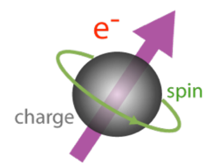
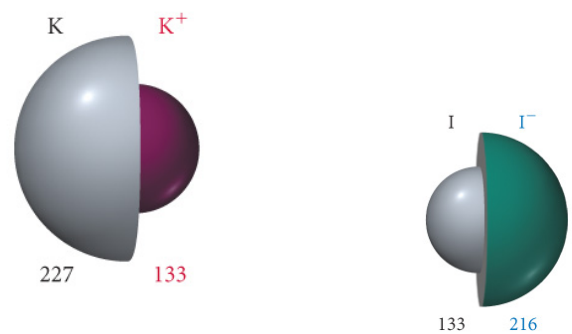
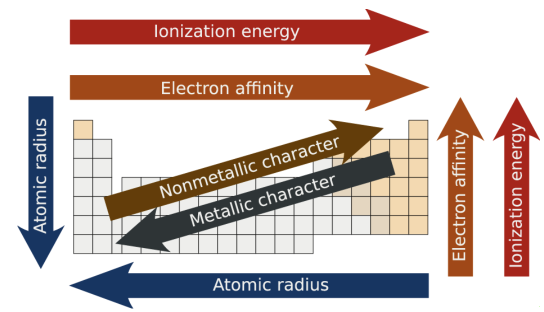
        - Electron Configurations
        - Size
        - Electron Affinity
- Metallic Character
- Patterns in Chemical Reactivity





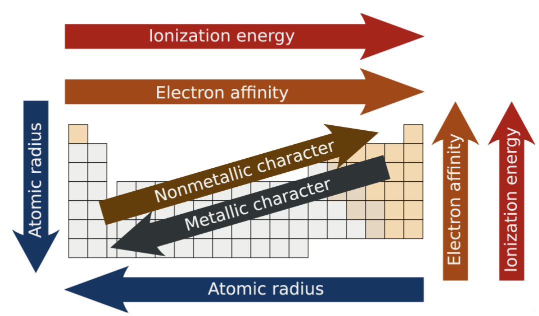




Electron Configuration of Anions

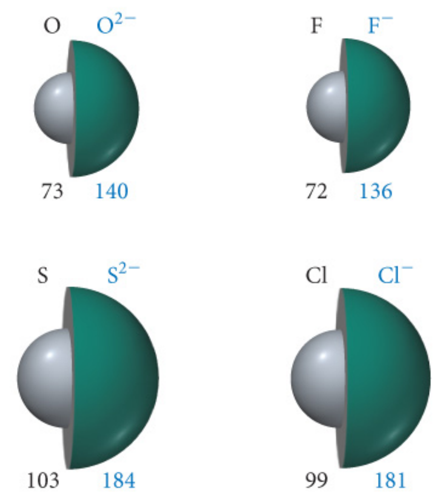
- We build up the electron configuration of a neutral atom by considering three principles:
  - Aufbau Principle “work up from the bottom”
  - Hund’s Rule “don’t double book unless you have to”
  - Pauli Exclusion Principle “pair spins when you double book”
- To form an, we start with the neutral atom, and then add electrons.
  - The same rules apply.

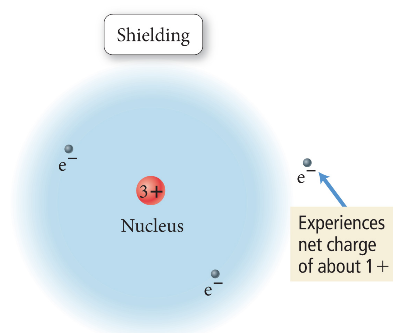
O	8 electrons	$1s^2 2s^2 2p^4$	O^{2-}	8 electrons + 2 e	$1s^2 2s^2 2p^6$
N	7 electrons	$1s^2 2s^2 2p^3$	N^{3-}	7 electrons + 3 e	$1s^2 2s^2 2p^6$
Cl	17 electrons	$1s^2 2s^2 2p^6 3s^2 3p^5$	Cl^{1-}	17 electrons + 1 e	$1s^2 2s^2 2p^6 3s^2 3p^6$

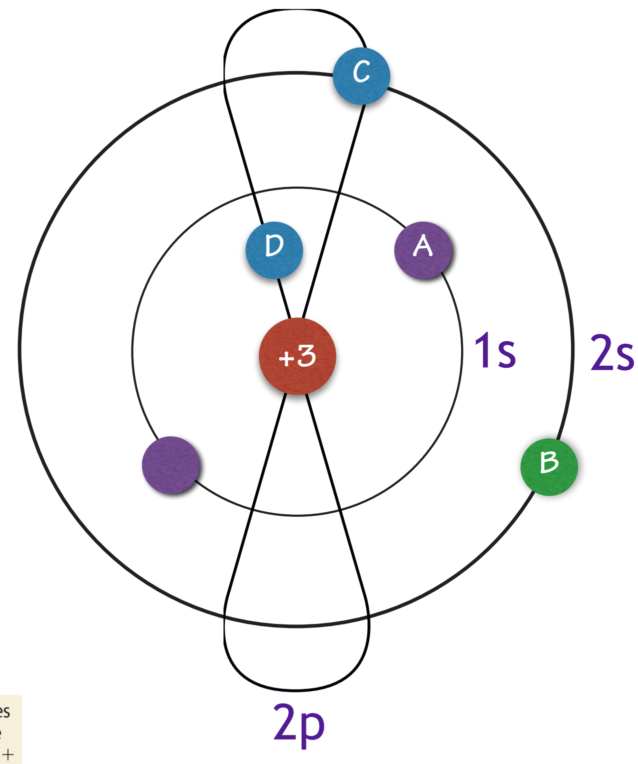


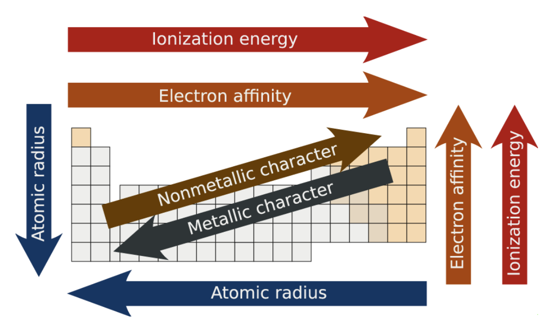
Size of Anions

- Anions have a larger atomic radius than the corresponding neutral atom.
- Anions are formed by adding electrons to the outermost electron shell.
- The electron configuration becomes the same as the configuration for the nobel gas element of the same period.
- Without any increase in nuclear charge!
- There are now more electrons in the outer most shell, pushing against each other.
- And each electron is seeing a lesser nuclear charge.
- The outer shell stretches out to accommodate and creates a larger atomic radius.





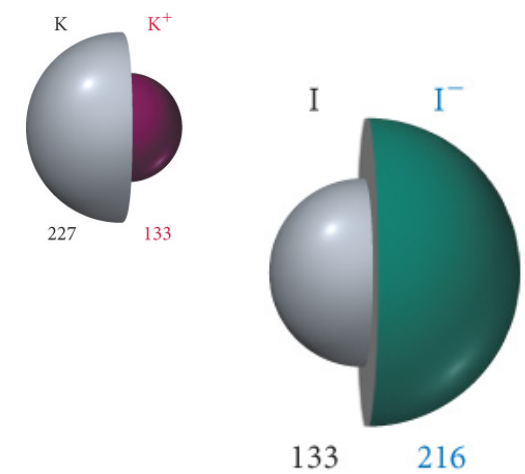


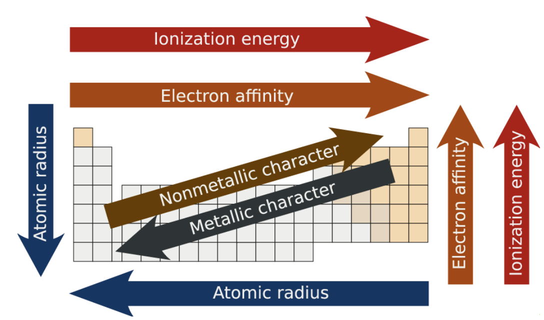


$$E = \frac{1}{4\pi\varepsilon_0} \times \frac{q_1 q_2}{r}$$

Comparing Radius

- What's bigger, B or N? **B**
- What's bigger Cl or Br? **Br**
- What's bigger Cl^{1-} or Br^{1-}? **Br^{1-}**
- What's bigger C or C^{4+}? **C**
- What's bigger C^{4-} or C? **C^{4-}**
- What's bigger C^{4-} or C^{4+}? **C^{4-}**
- Which is bigger S^{2-}, Ar, or Ca^{2+}? **S^{2-}**



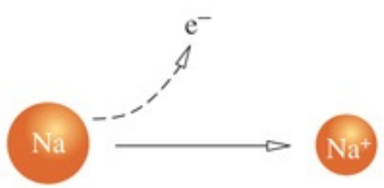


Electron Affinity (EA)

- Electron affinity is the energy released by adding an electron to an atom.
- Bonding is a result of sharing electrons, electron affinity will be a factor in covalent bond formation (chapter 9).
- Less energy is released as we go down the periodic table.
- More energy is released as we go across the periodic table (left to right).
  - Noble Gases have a positive AE, no energy is released when they accept an electron.
    - They aren't very reactive.
  - Non-metals tend to have high AE, we get a lot of energy by giving them electrons.
    - Pure non-metals tend to be very reactive, they even react with themselves.
    - N_2, O_2, Cl_2, Br_2

Electron Affinities (kJ/mol)

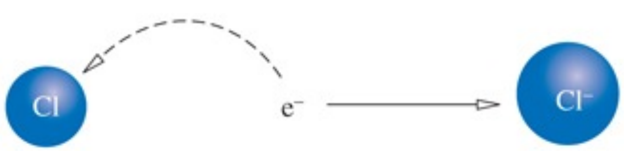
1A	2A	3A	4A	5A	6A	7A	8A
H −73							He >0
Li −60	Be >0	B −27	C −122	N >0	O −141	F −328	Ne >0
Na −53	Mg >0	Al −43	Si −134	P −72	S −200	Cl −349	Ar >0
K −48	Ca −2	Ga −30	Ge −119	As −78	Se −195	Br −325	Kr >0
Rb −47	Sr −5	In −30	Sn −107	Sb −103	Te −190	I −295	Xe >0



Ionization Energy (IE)

ΔH = +496 kJ/mol — endothermic

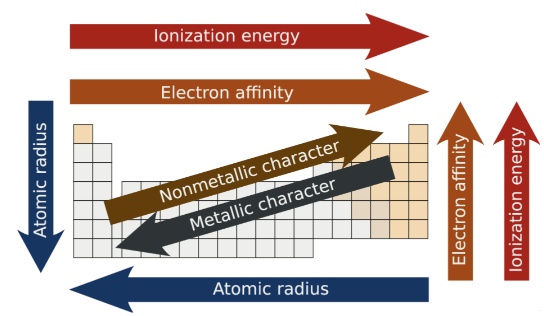
forming cations *consumes* energy



Electron Affinity (EA)

ΔH = -349 kJ/mol — exothermic

forming anions *releases* energy



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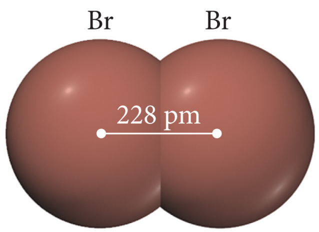
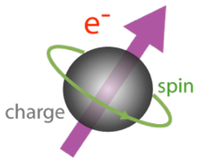
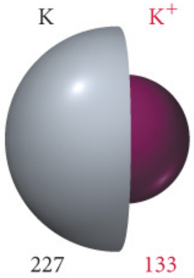
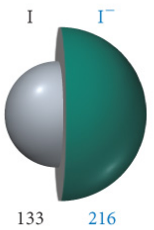
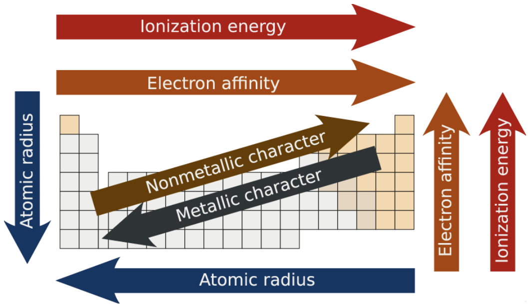




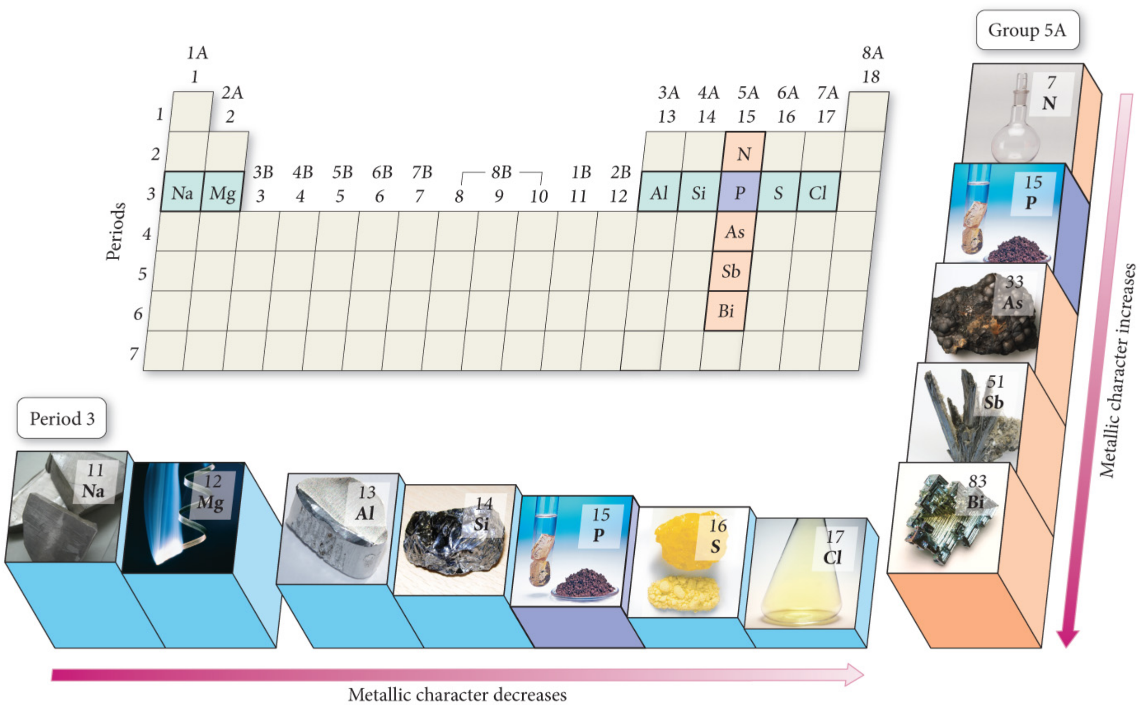


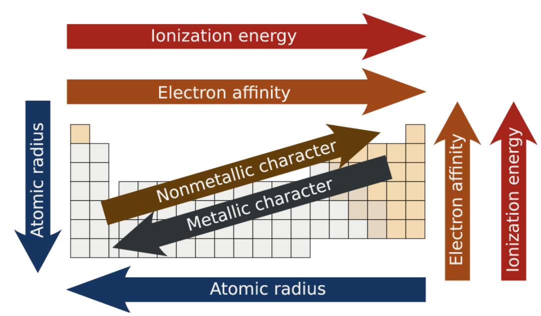




Metallic Character

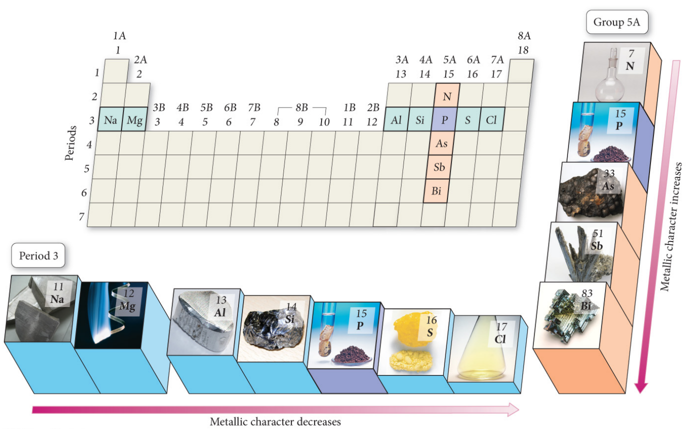
- Metallic character is a general term for how well a element matches the profile of a metallic element.
- Metallic elements are:
  - Good conductors of heat and electricity.
  - Reflective (shiny).
  - Meltable and ductile.
  - Tend to release electrons.
  - Form solids at lower temperatures.
- Metallic character decreases as you move across the periodic table.
- Metallic character increases as you move down the periodic table.

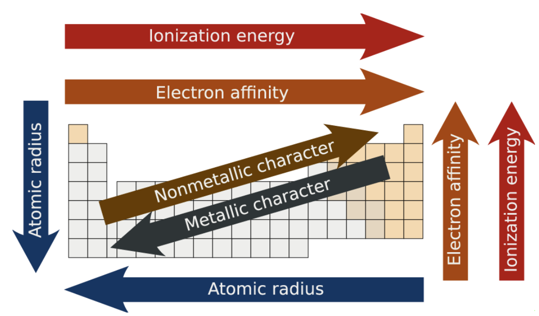




Using Metallic Character Trends

- What substance is a better electrical conductor, Iron or Zinc? **Iron**
- What substance is a better electrical conductor, Selenium or Chloride? **Selenium**
- What substance is more brittle, Sulfur or Aluminum? **Sulfur**
- What substance conducts less heat, Carbon or Copper? **Carbon**
- What substance is more shiny, Phosphorus or Silver? **Silver**





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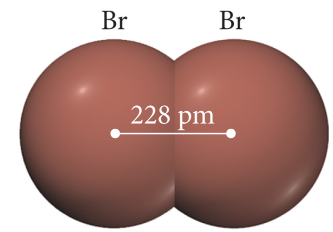
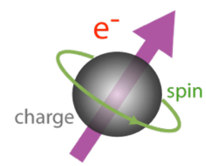
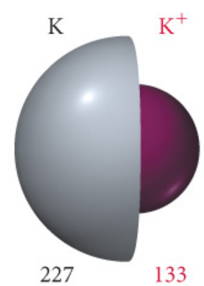
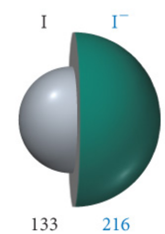
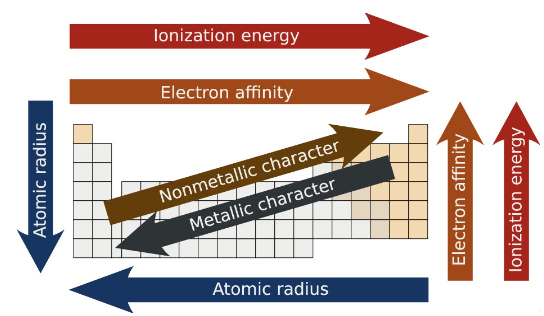




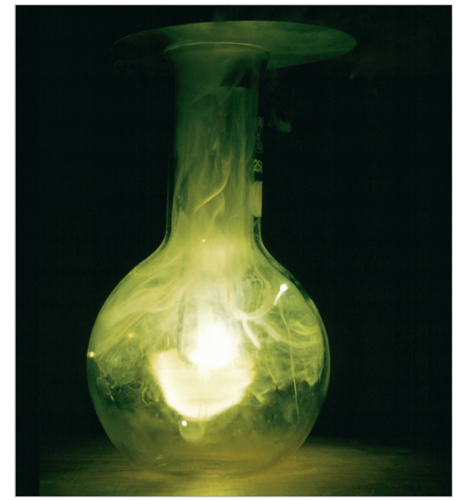


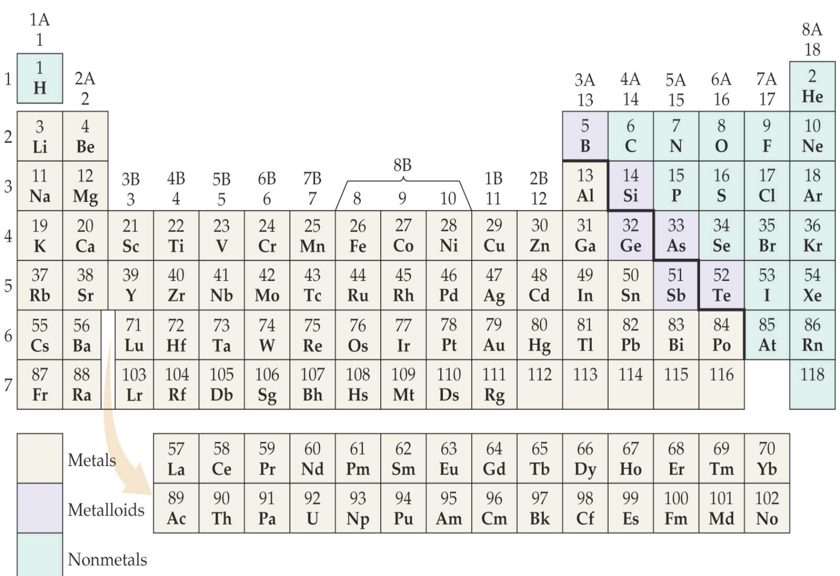


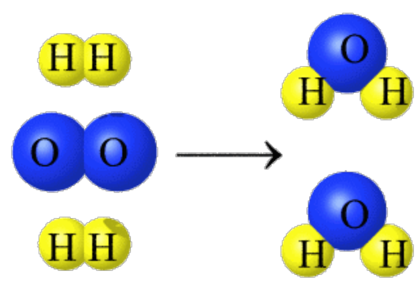


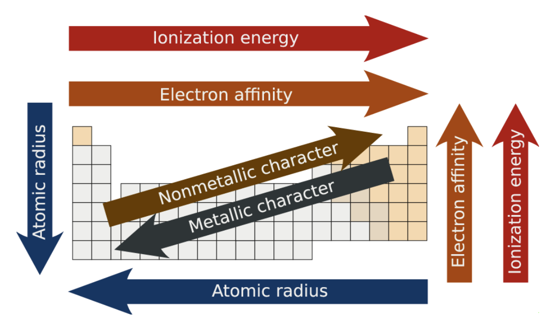
Chemical Properties & Reactivity

- Electron configuration predicts chemical properties.
- Elements with similar valence electron configuration predicts chemical reactivity.
  - Chemical reactivity is a result of interactions between atoms.
  - We see similar chemical reactions with elements that have similar electronic configurations.
  - Atoms interact when their outermost electron shells entangle (valence electrons).









Alkali Metals

1A

3 **Li** $2s^1$
11 **Na** $3s^1$
19 **K** $4s^1$
37 **Rb** $5s^1$
55 **Cs** $6s^1$
87 **Fr** $7s^1$

Alkali metals

Lithium

Sodium

Potassium

- They have a single electron in their valence shell.
- Alkali Metals have a low ionization energy.
  - They tend to form +1 ions.
  - They have strongly metallic properties.
    - Solids, Conductors, Maleable, etc
  - They are strong reducing agents.
  - They don't react with other metals.
  - Their reactivity increases as you move down the periodic table.

$$2\ Na + H_2O \rightarrow 2\ NaOH + H_{2(g)}$$

$$2\ Na + Cl_2 \rightarrow 2\ NaCl$$

TABLE 8.2 Properties of the Alkali Metals*

Element	Electron Configuration	Atomic Radius (pm)	IE_1 (kJ/mol)	Density at 25 °C (g/cm^3)	Melting Point (°C)
Li	[He] $2s^1$	152	520	0.535	181
Na	[Ne] $3s^1$	186	496	0.968	102
K	[Ar] $4s^1$	227	419	0.856	98
Rb	[Kr] $5s^1$	248	403	1.532	39
Cs	[Xe] $6s^1$	265	376	1.879	29

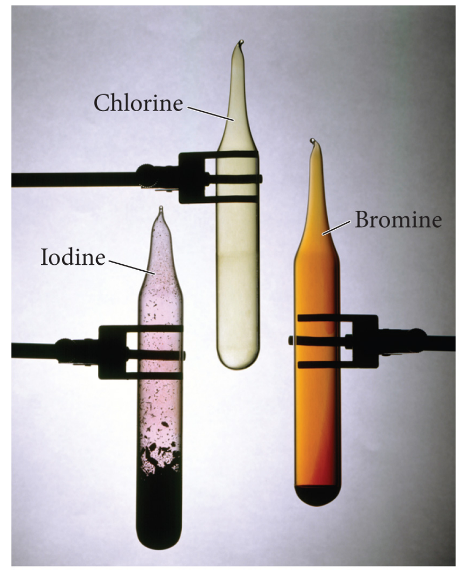
Halogens

- They have seven electrons in their valence shell.
- Their valence shell has room for only one more electron.
- Alkali Metals have high electron affinity.
  - They tend to form -1 ions.
  - They have weak metallic properties.
    - Most are not solids at room temperature, poor conductors, etc.
  - They are strong oxidizing agents.
  - They are highly reactive with everything (even themselves).
  - Their reactivity decreases as you move down the periodic table.

$$2\ Fe + 3\ X_2 \rightarrow 2\ FeX_{3\ (s)}$$

$$H_2 + X_2 \rightarrow 2\ HX$$

$$F_2 + Cl_2 \rightarrow 2\ ClF$$



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7A

7A
9 **F** $2s^2 2p^5$
17 **Cl** $3s^2 3p^5$
35 **Br** $4s^2 4p^5$
53 **I** $5s^2 5p^5$
85 **At** $6s^2 6p^5$

Halogens

TABLE 8.3 Properties of the Halogens*

Element	Electron Configuration	Atomic Radius (pm)	EA (kJ/mol)	Melting Point (°C)	Boiling Point (°C)	Density of Liquid (g/cm^3)
F	[He] $2s^2 2p^5$	72	−328	−219	−188	1.51
Cl	[Ne] $3s^2 3p^5$	99	−349	−101	−34	2.03
Br	[Ar] $4s^2 4p^5$	114	−325	−7	59	3.19
I	[Kr] $5s^2 5p^5$	133	−295	114	184	3.96

Noble Gases

8A
2 **He** $1s^2$
10 **Ne** $2s^2 2p^6$
18 **Ar** $3s^2 3p^6$
36 **Kr** $4s^2 4p^6$
54 **Xe** $5s^2 5p^6$
86 **Rn** $6s^2 6p^6$

Noble gases

- They have a full valence shell.
- They have poor (endothermic) electron affinity.
  - You can't give them electrons.
- They have high ionization energy.
  - It's hard to take their electrons.
- They tend to be inert.
  - Don't form many ions.
  - They make few compounds.
  - They have weak metallic properties.
    - Most are not solids at room temperature, poor conductors, etc.
  - They are weak oxidizing agents.
  - They are essentially unreactive, with everything.

TABLE 8.4 Properties of the Noble Gases*

Element	Electron Configuration	Atomic Radius (pm)**	IE_1 (kJ/mol)	Boiling Point (K)	Density of Gas (g/L at STP)
He	$1s^2$	32	2372	4.2	0.18
Ne	$[He]2s^2 2p^6$	70	2081	27.1	0.90
Ar	$[Ne]3s^2 3p^6$	98	1521	87.3	1.78
Kr	$[Ar]4s^2 4p^6$	112	1351	119.9	3.74
Xe	$[Kr]5s^2 5p^6$	130	1170	165.1	5.86

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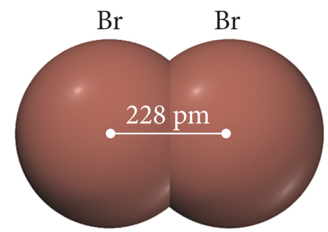
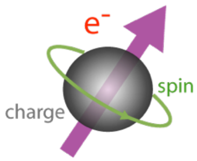
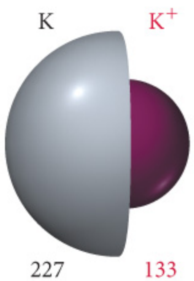
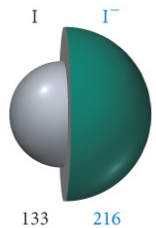
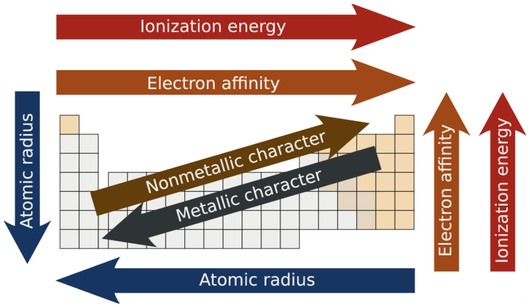


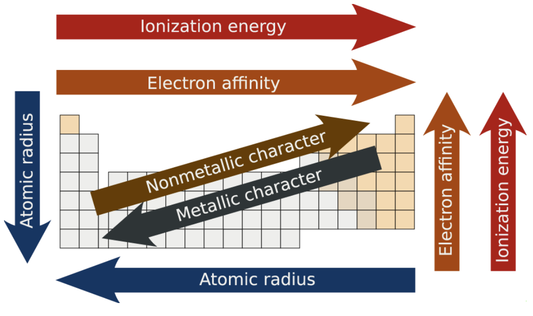








Questions?


Ionization energy
Electron affinity
Atomic radius
Nonmetallic character
Metallic character
Electron affinity
Ionization energy
Atomic radius