Chemistry, The Central Science, 11th edition
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Chapter 7
Periodic Properties of the Elements

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Periodic Trends

- In this chapter, we will rationalize observed trends in
  - Sizes of atoms and ions
  - Ionization energy
  - Electron affinity
  - Electronegativity

Periodic Trends Key Words

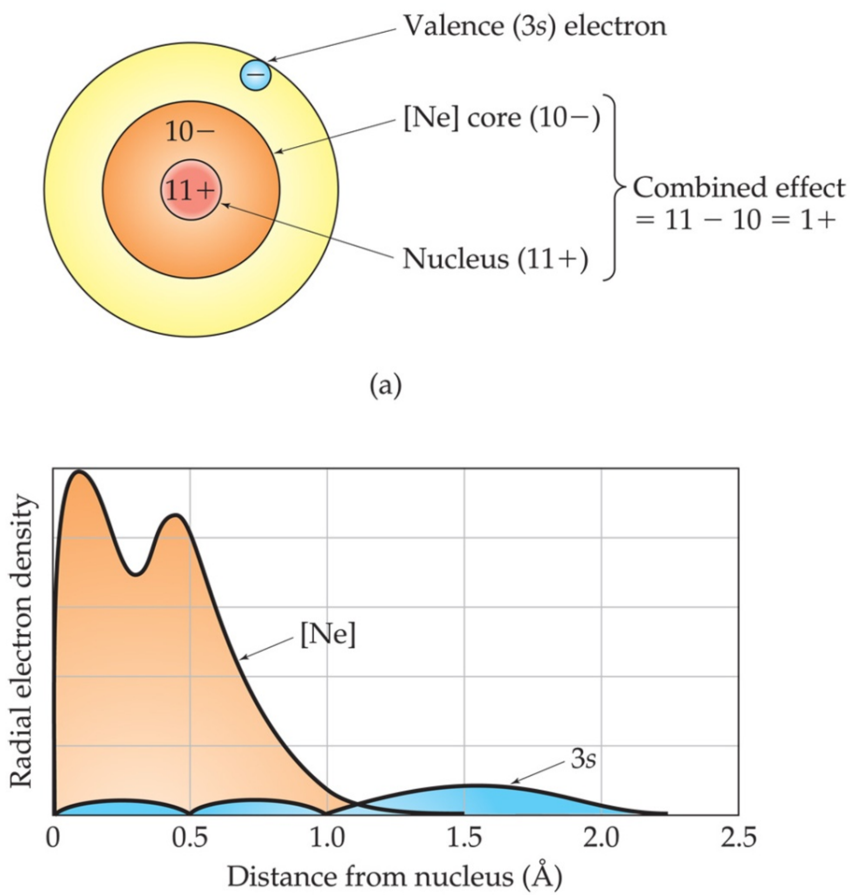
- **Principal Energy Levels:** The more the number of principal energy levels, the bigger the size of atoms.
- **Nuclear Charge (# of p in an atom):** Results into increased attraction on electrons. Causes atomic radius to decrease.
- **Shielding Effect**: Electrons present between nucleus and outermost energy level (all electrons except for valence electrons).

Periodic Trends Key Words

Shielding electrons tend to increase atomic size by reducing the attractive force on outermost electrons.

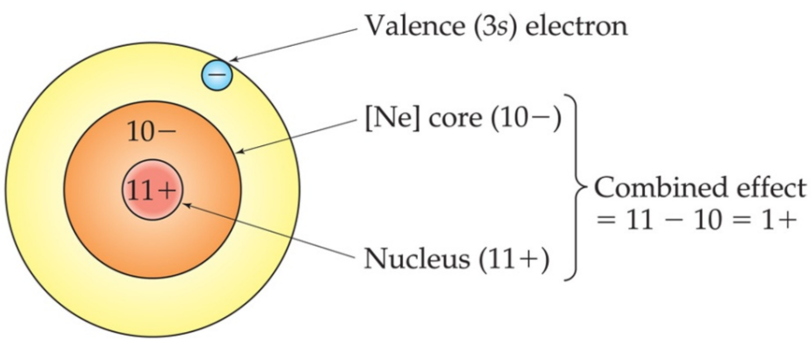
- **Effective Nuclear Charge**: Force of attraction felt by the outermost (valence e) from the protons in the nucleus. Effective nuclear charge depends upon the two counteractive factors of nuclear charge and shielding effect. A high effective nuclear charge means smaller ionic radius (greater attraction on the outermost electrons).

Effective Nuclear Charge



- In a many-electron atom, electrons are both attracted to the nucleus and repelled by other electrons.
- The nuclear charge that an electron experiences depends on both factors.

Effective Nuclear Charge



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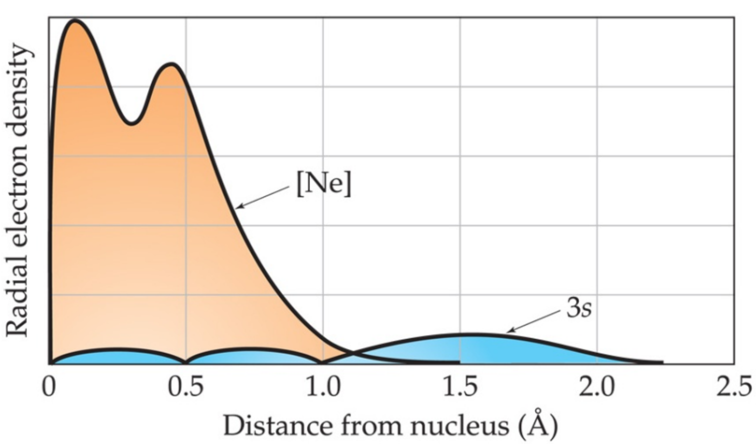
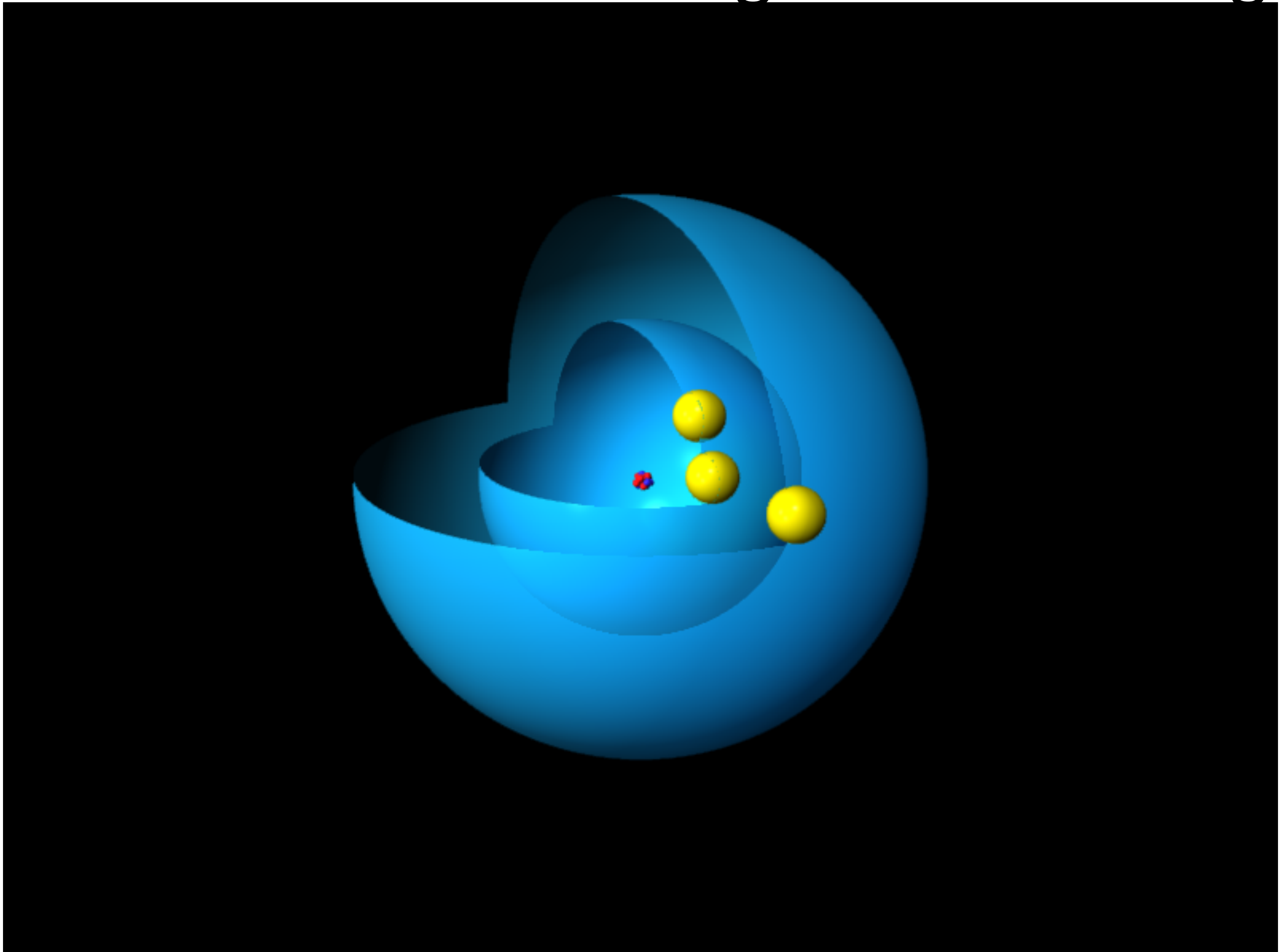


The effective nuclear charge, Z_{eff}, is found this way:

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

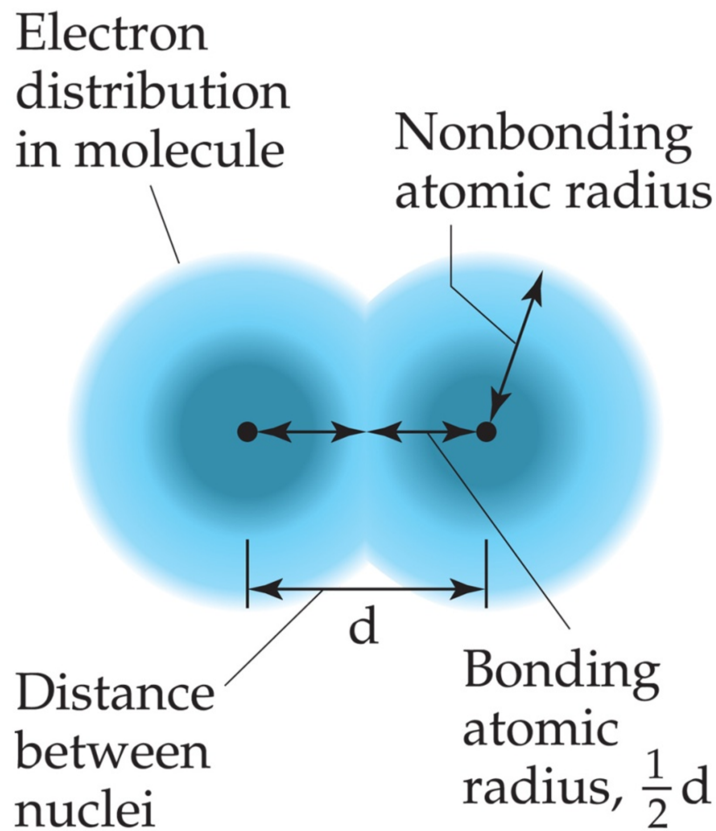
where Z is the atomic number and S is a screening constant, usually close to the number of inner electrons.

Effective Nuclear Charge and Shielding

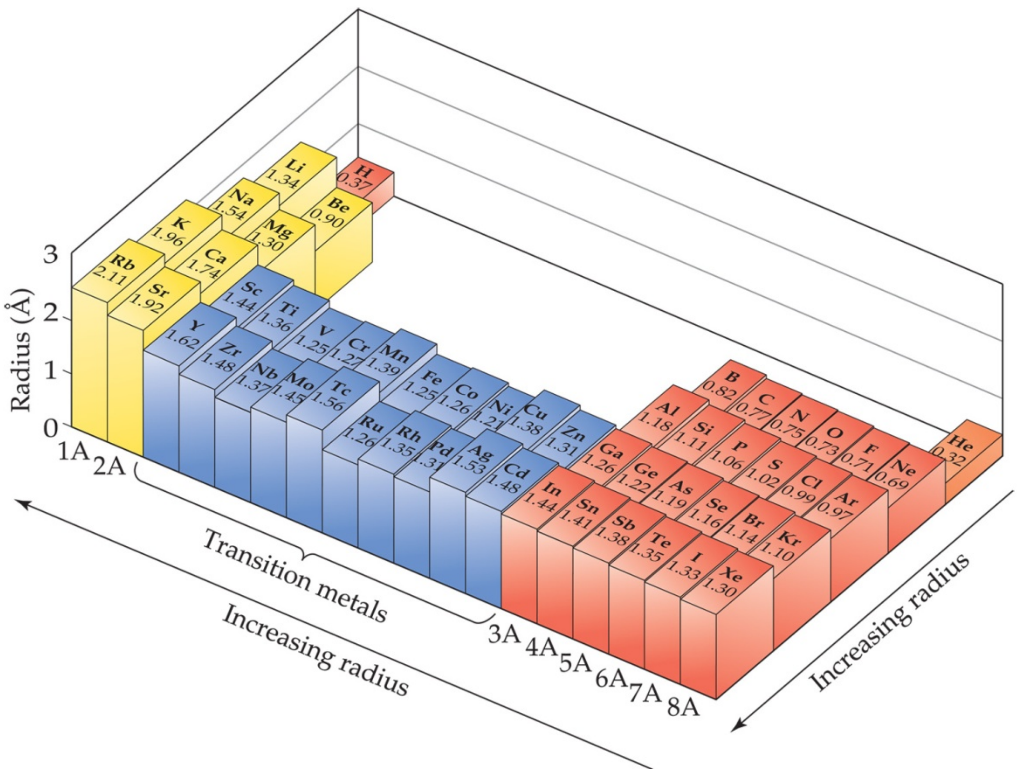


What Is the Size of an Atom?

The bonding atomic radius is defined as one-half of the distance between covalently bonded nuclei.



Sizes of Atoms



Bonding atomic radius tends to…

…decrease from left to right across a row

(due to increasing Z_{eff}).

…increase from top to bottom of a column

(due to increasing value of *n*).

http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/atomic4.swf

Periodic Trend for Atomic Radius

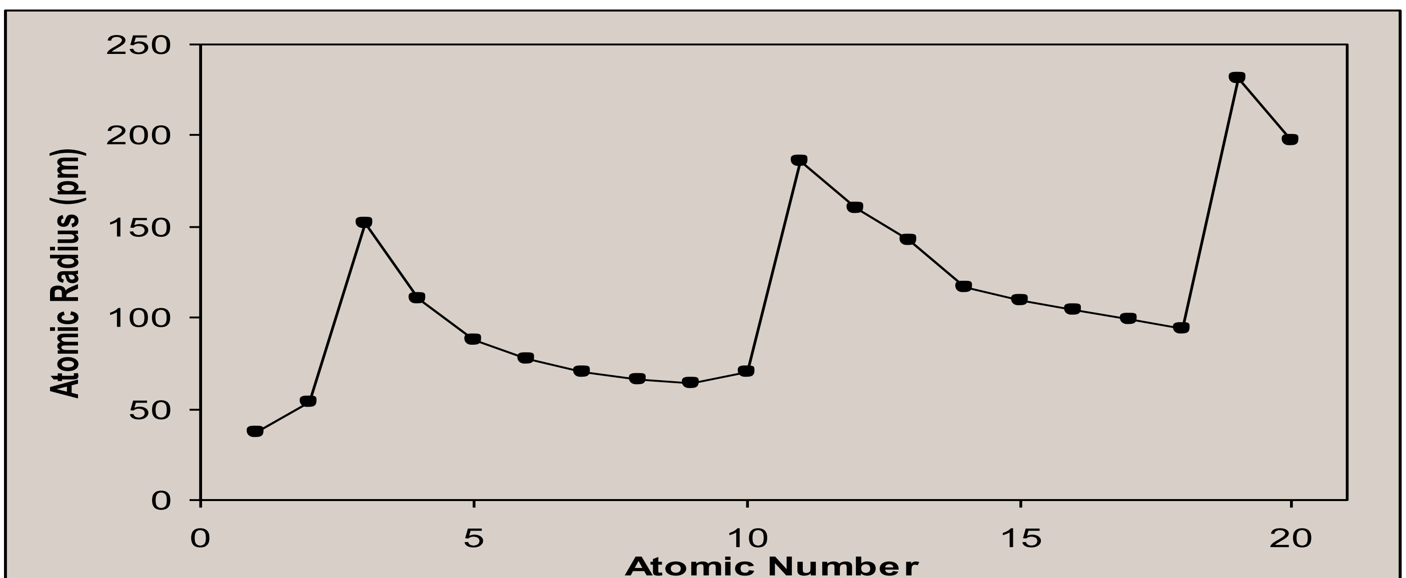
Going down a group- ENC is lower; size is larger:

Higher energy levels have larger orbitals

<u>Shielding</u> - core e^- block the attraction between the nucleus and the valence e^-

Across a group- ENC is higer; size is smaller:

Increased nuclear charge without additional shielding pulls e^- in tighter



Atomic Radius

- **Why larger going down?**
  - Higher energy levels have larger orbitals
  - Shielding - core e^- block the attraction between the nucleus and the valence e^-
- **Why smaller to the right?**
  - Increased nuclear charge without additional shielding pulls e^- in tighter

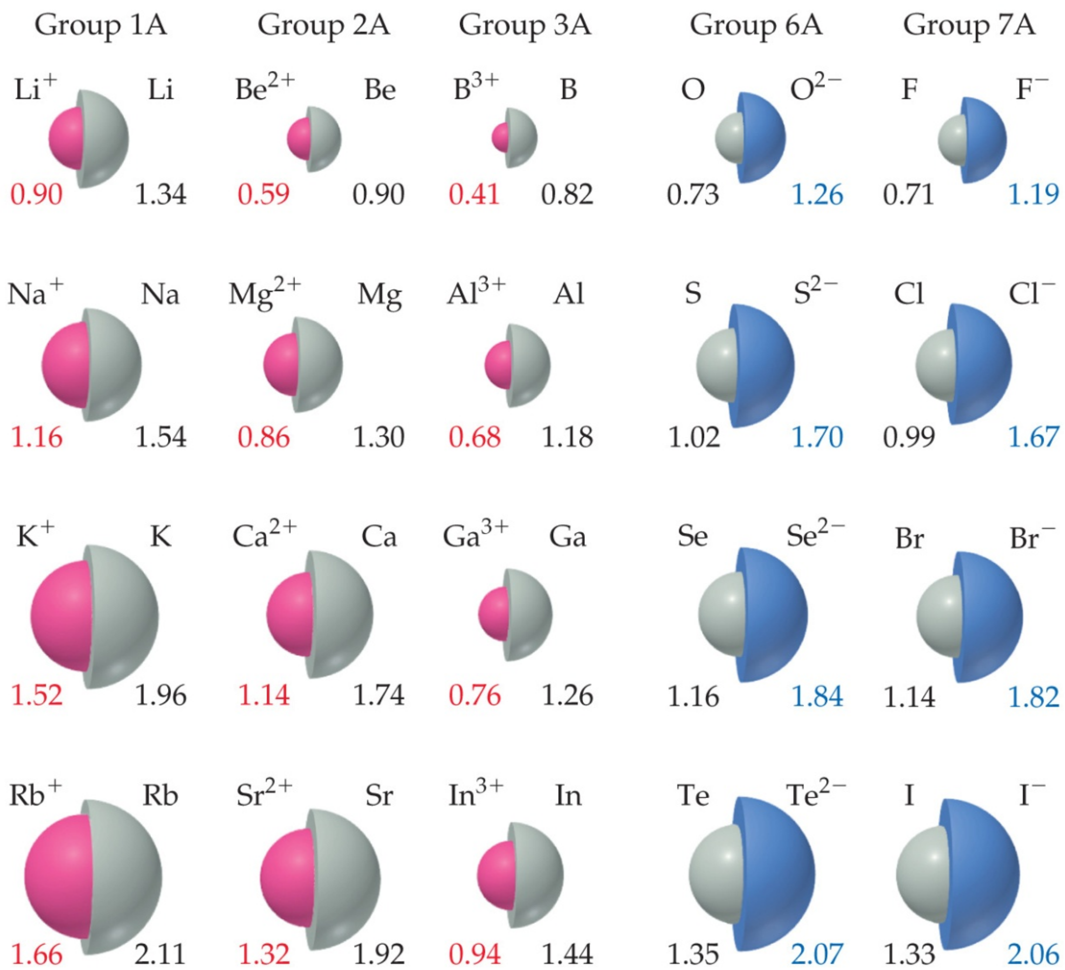
Explain to your shoulder partner the atomic radius trend:

a. Across a period
b. Down a group

Be sure to use the following key words:

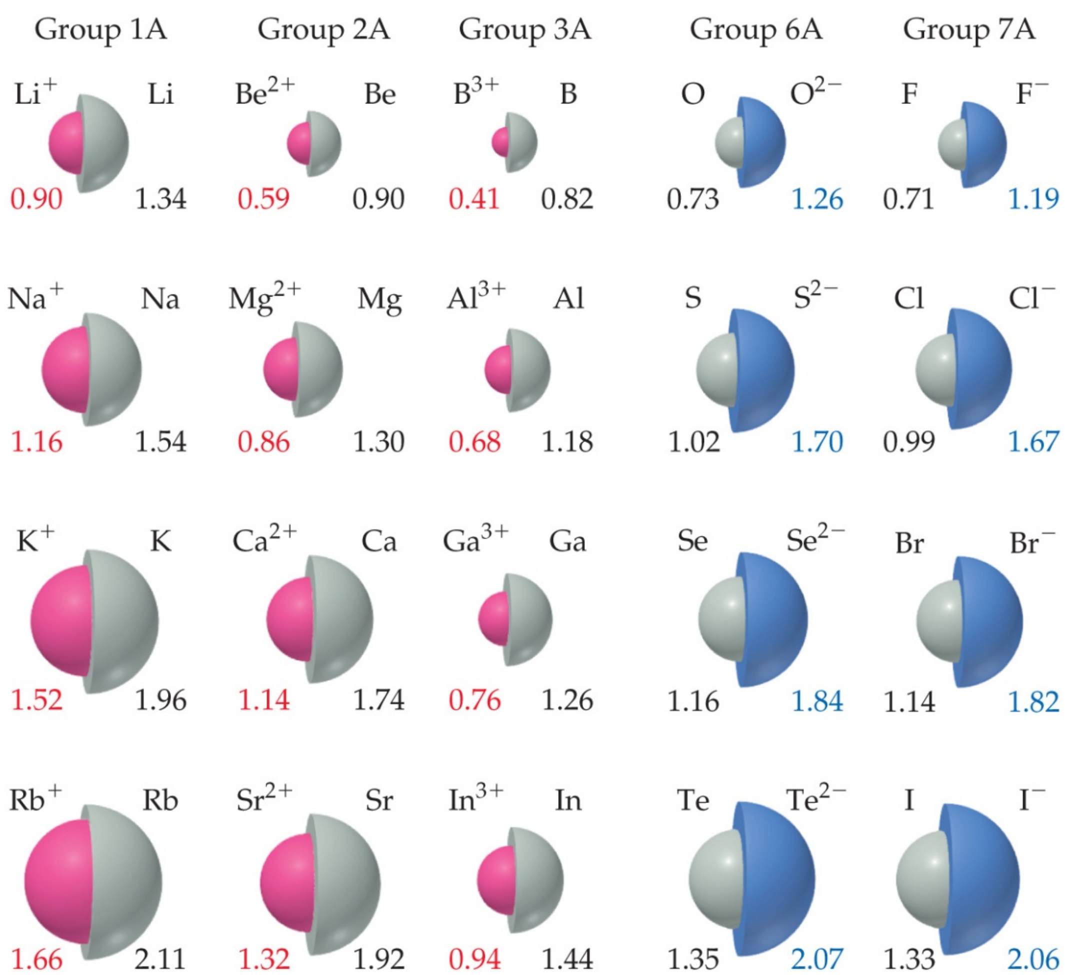
- N (number of energy levels)
- Nuclear Charge
- Shielding Effect
- Effective Nuclear Charge

Sizes of Ions



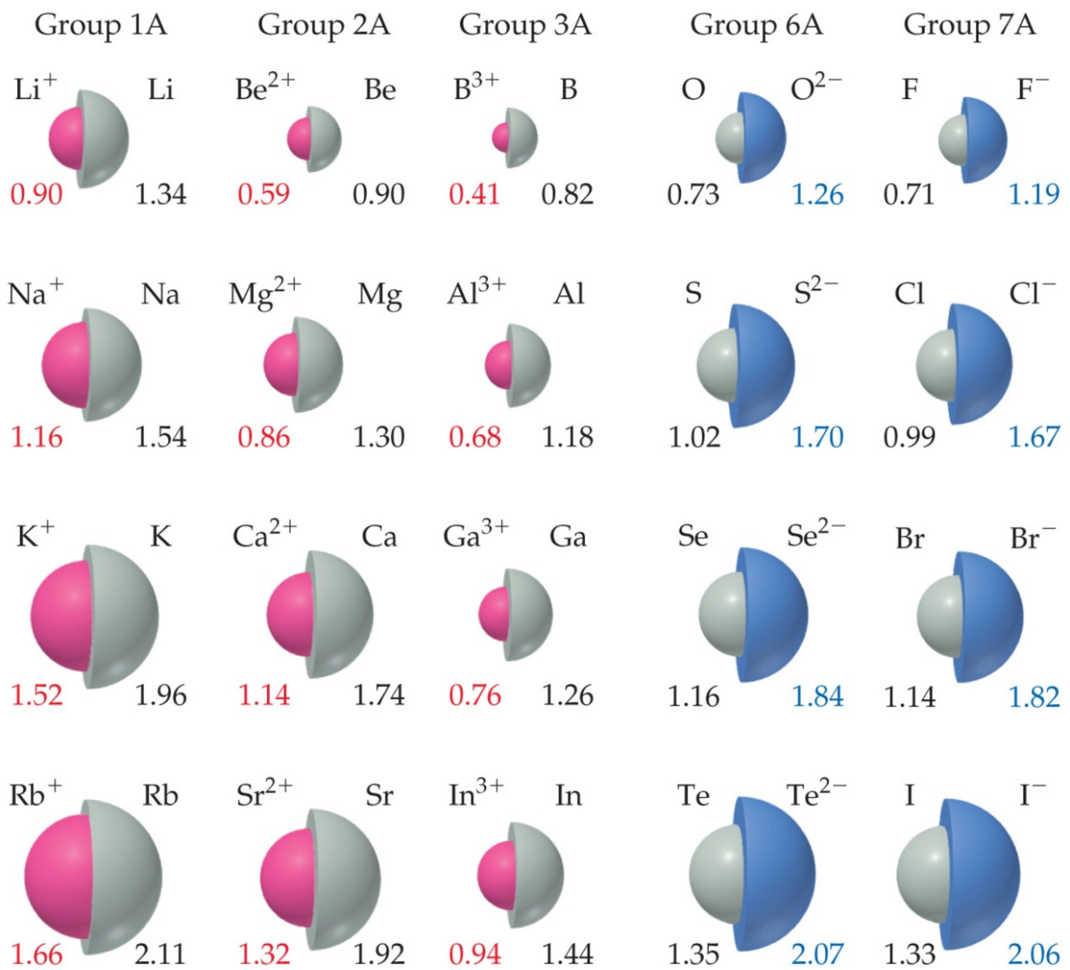
- Ionic size depends upon:
  - The nuclear charge.
  - The number of electrons.
  - The orbitals in which electrons reside.

Sizes of Ions



- Cations are smaller than their parent atoms.
  - The outermost electron is removed and repulsions between electrons are reduced.

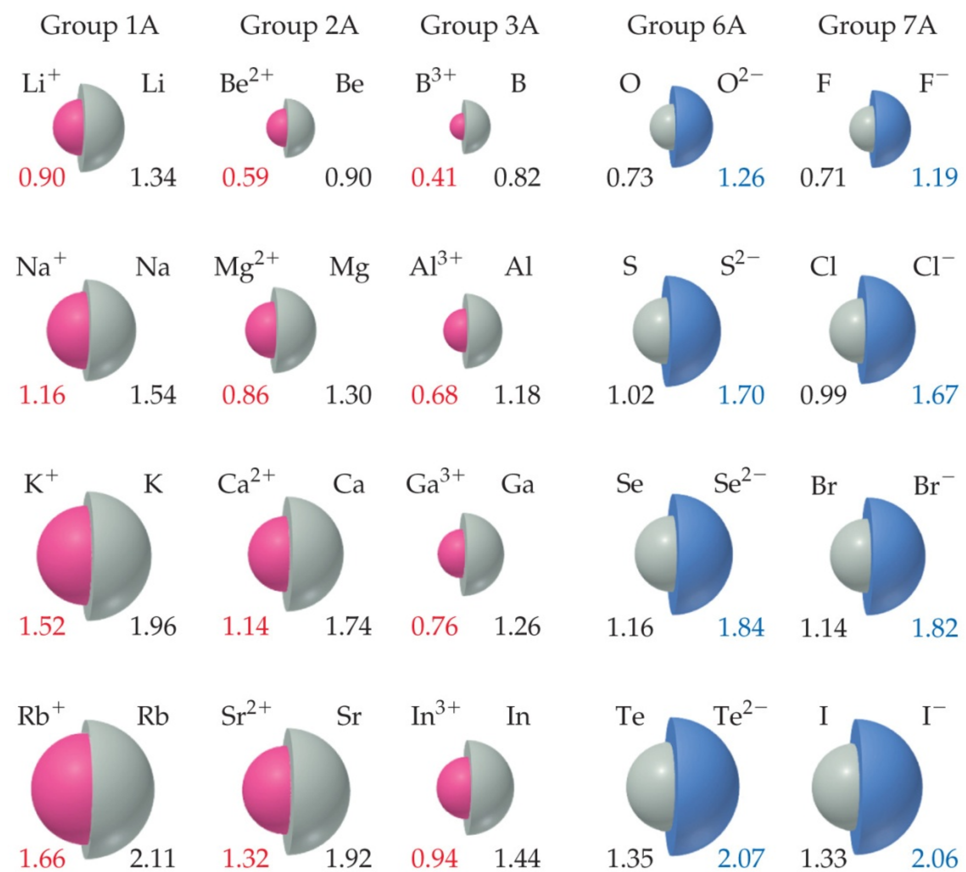
Sizes of Ions



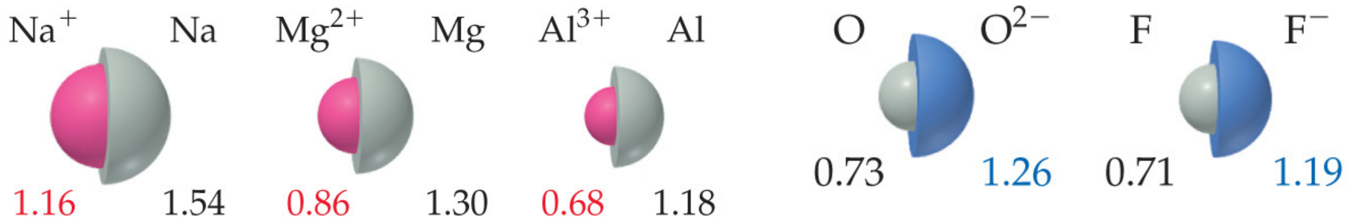
- Anions are larger than their parent atoms.
  - Electrons are added and repulsions between electrons are increased.

Sizes of Ions

- Ions increase in size as you go down a column.
  - This is due to increasing value of *n*.

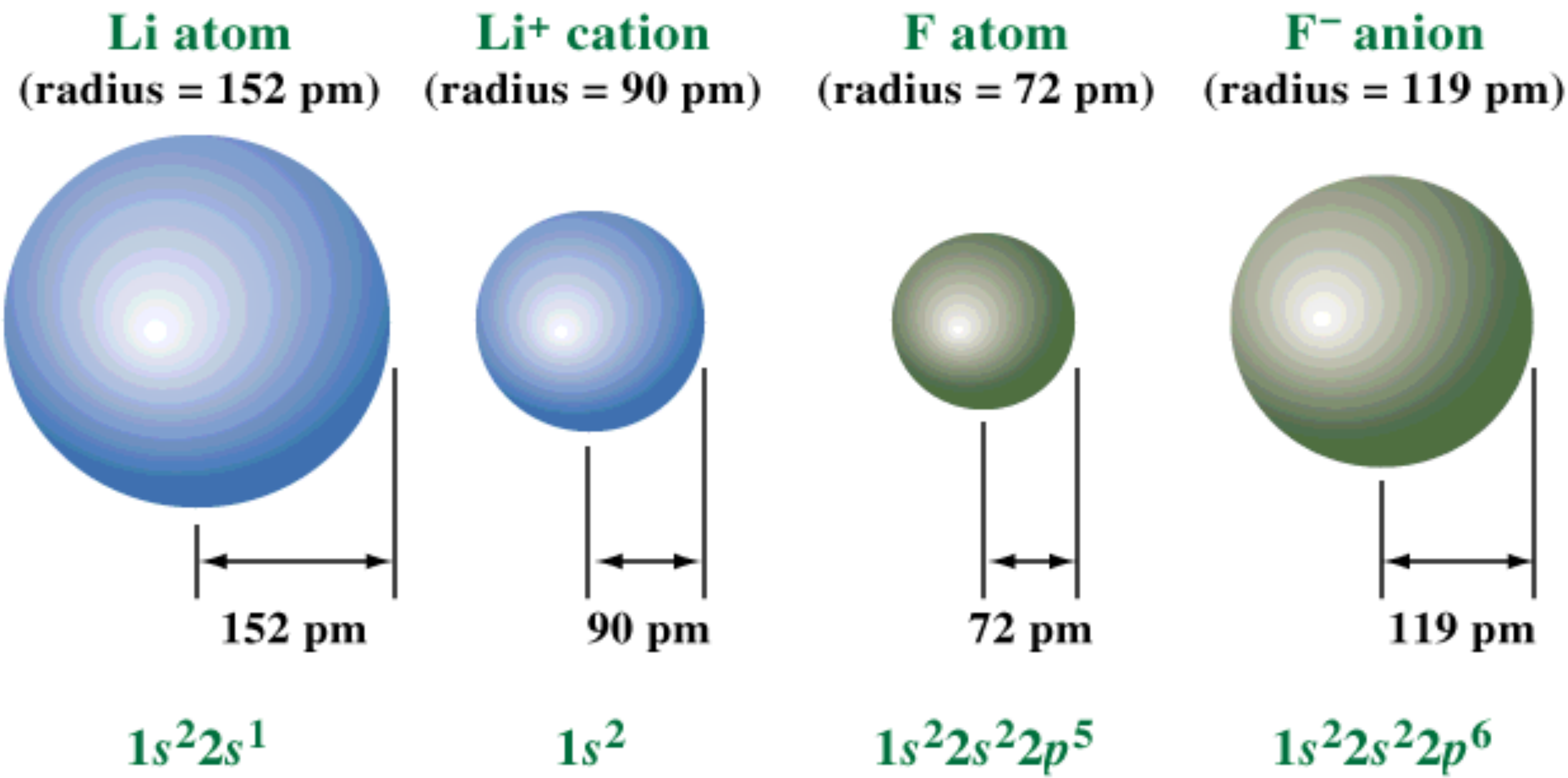


Sizes of Ions



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

Li atom and Li^+ cation; F atom and F^- anion



MR08U20.PIC

Find a new partner (someone who you have never worked with before!) and do the following pair-share activity:
Explain to your partner periodic trend for size of ions
Your partner explains to you the group trend for size of ions
Make sure to use the following key words:

- N (number of energy levels)
- Nuclear charge
- Shielding effect
- Effective Nuclear Charge

Now find another group and explain your partner's reasoning for the trend to them, while your partner explains your reasoning for the trend to them.

Ionization Energy

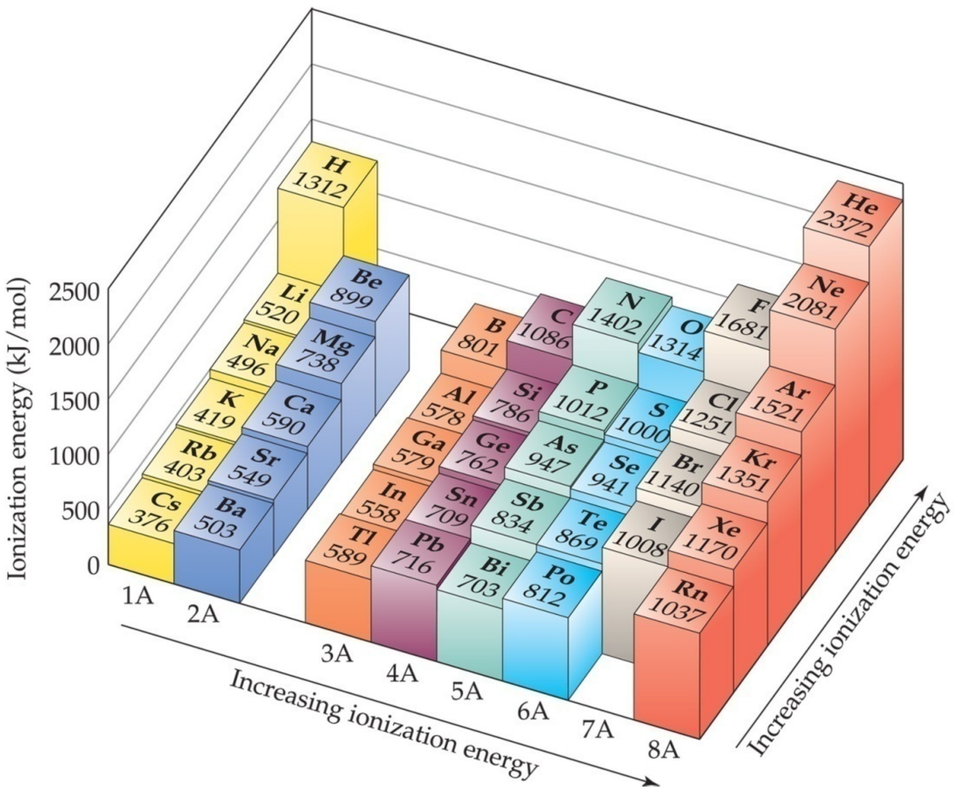
- The ionization energy is the amount of energy required to remove an electron from the ground state of a gaseous atom or ion.
- A(g)→ A+ + e
  - The first ionization energy is that energy required to remove first electron.
  - The second ionization energy is that energy required to remove second electron, etc.

Ionization Energy

- It requires more energy to remove each successive electron.
- When all valence electrons have been removed, the ionization energy takes a quantum leap. Why? *Hint: where are the electrons being removed from?*

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

Trends in First Ionization Energies

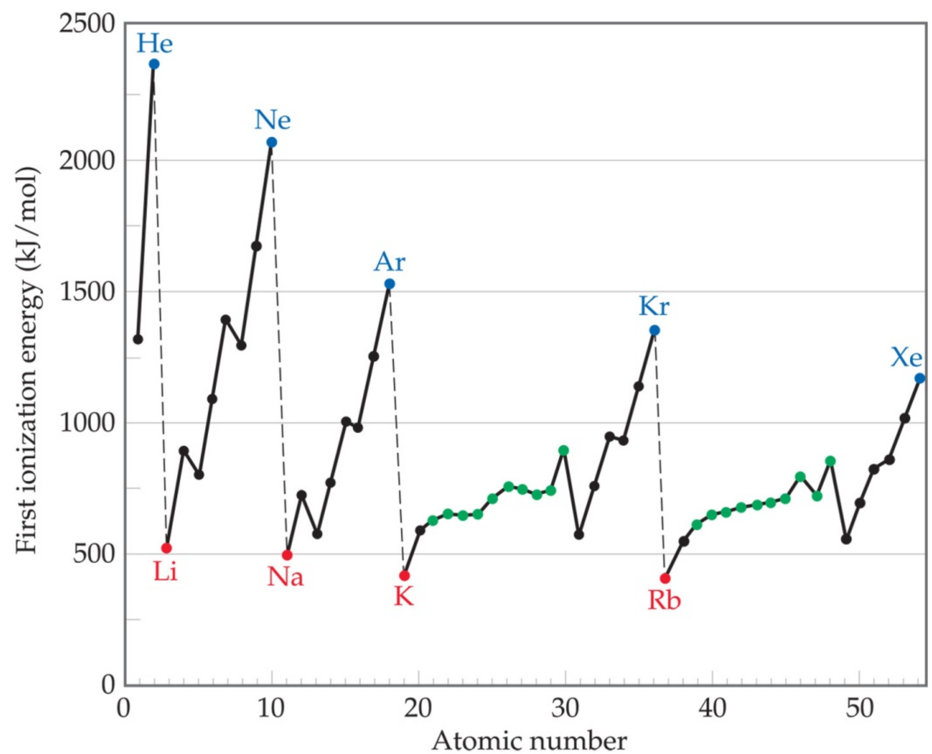


- As one goes down a column, less energy is required to remove the first electron.
  - For atoms in the same group, Z_{eff} is essentially the same, but the valence electrons are farther from the nucleus.

http://nuweb.neu.edu/bmaheswaran/phyu121/data/ch09/anim/anim0903.htm

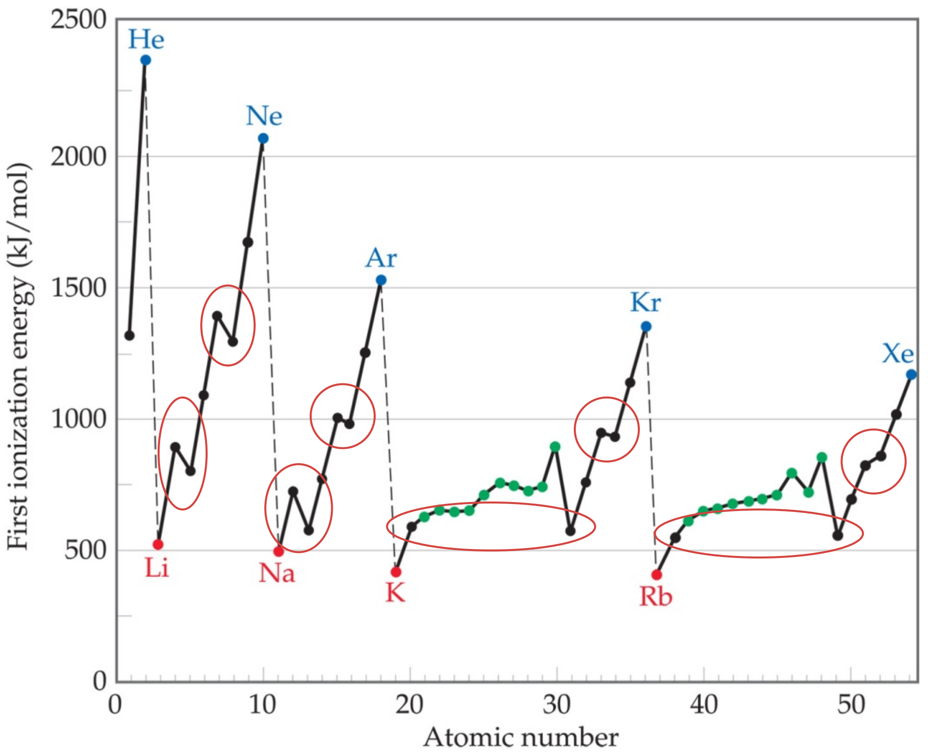
Trends in First Ionization Energies

- Generally, as one goes across a row, it gets harder to remove an electron.
  - As you go from left to right, Z_{eff} increases.



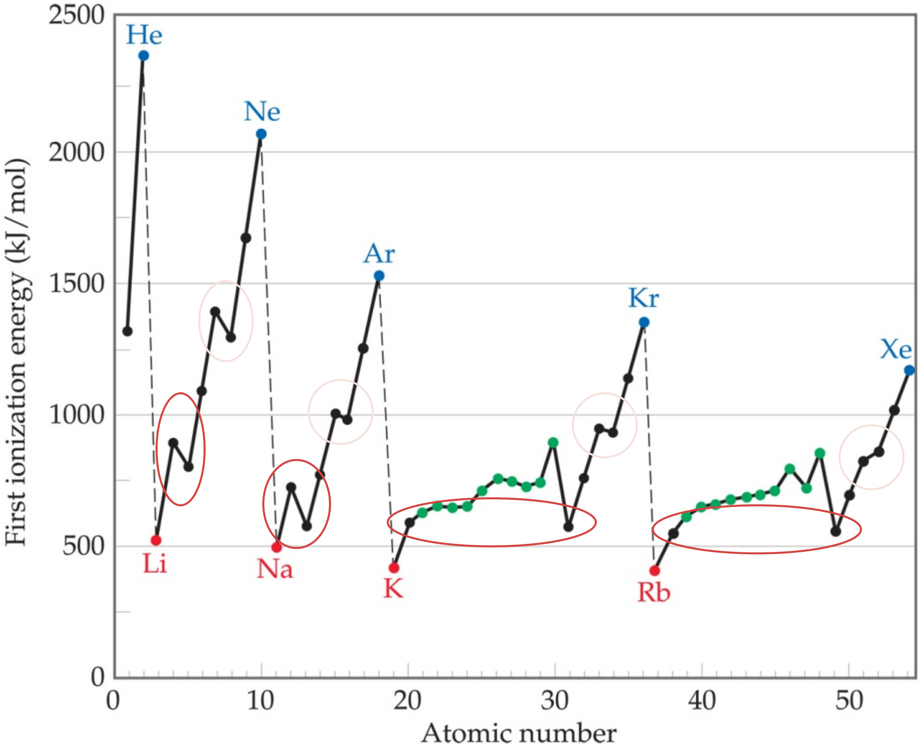
Trends in First Ionization Energies

However, there are two apparent discontinuities in this trend.



Trends in First Ionization Energies

- The first discontinuity occurs between Groups 2 and 13.
- In group 13, the electron is removed from a *p*-orbital rather than an *s*-orbital.
  - The electron removed is farther from nucleus.
  - There is also a small amount of repulsion by the *s* electrons.



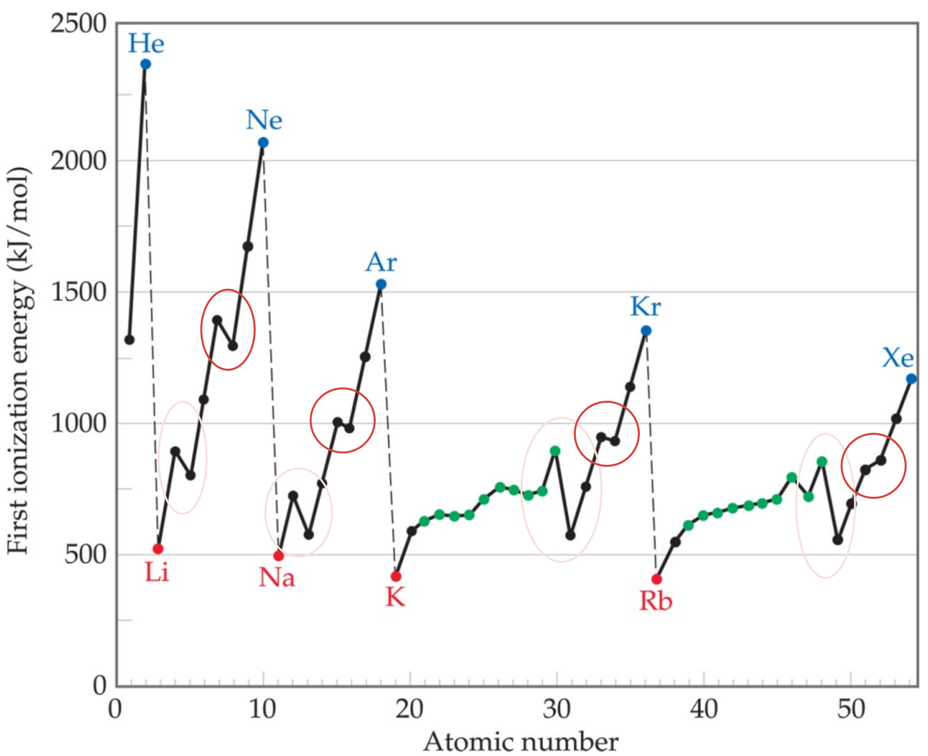
Pair-Share Activity

With your elbow partner, discuss the following question. You might be called upon to share your explanation with the whole class:

- Which element has a higher Ionization energy- Be or B? Why?

Trends in First Ionization Energies

- The second anomaly occurs between Groups 5 and 6. Following the trend, the IE of group 6 elements should be higher than group 5, but IE of group 6 elements is lower than that of group 5.
- The electron removed comes from doubly occupied orbital.
- Repulsion from the other electron in the orbital aids in



Sample Problem

- Write the answer to the following sample problem on a piece of paper. You will be grading your elbow partner's paper at the end.
- Question: Which element has a higher Ionization energy- N or O? Why?

Electron Affinity

Electron affinity is the energy change accompanying the addition of an electron to a gaseous atom:

$$Cl\ (g) + e^{-} \longrightarrow Cl^{-}$$

Trends in Electron Affinity

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

In general, electron affinity becomes more exothermic (larger – value) as you go from left to right across a row.

http://www.youtube.com/watch?v=bPB0xThmpkg&feature=related

Trends in Electron Affinity

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
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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

There are again, however, two discontinuities in this trend.

Trends in Electron Affinity

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

- The first occurs between Groups 1 and 2, where EA of group 2 is lesser than group 1.
  - The electron is farther from nucleus and feels repulsion from the *s*-electrons.

Trends in Electron Affinity

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

- The second occurs between Groups 14and 15.
  - Group 15 has no empty orbitals (only half filled p orbitals).
  - The extra electron must go into an already occupied orbital, creating repulsion.

Sample Problem

- Why is the general periodic trend for EA? What is the group trend for EA? Why?

- Which of the following elements has a higher EA?
- Na or Mg?
- P or Si?
- Why?

Electronegativity

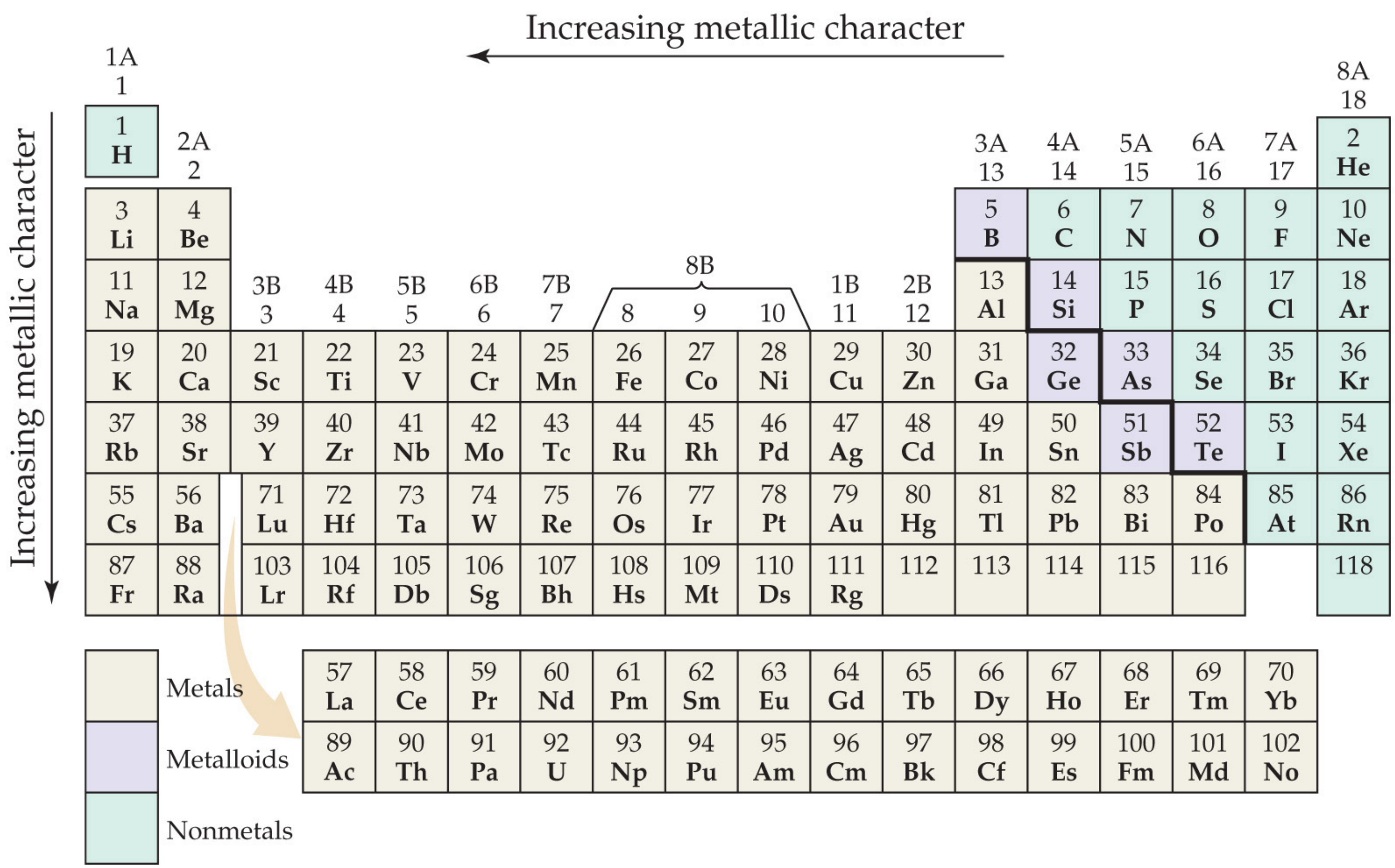
•Increases from L to R across a period and decreases down a group.

•Electronegativity is defined as tendency to attract electrons but it is different from electron affinity in that electro negativity is used in context of an element BONDED IN A COVALENT COMPOUND, while electron affinity is generally attributed to an atom by itself.

•Another difference is that electro negativity is a measure of affinity for electrons in debye scale, while electron affinity is the actual amount of energy released.
http://www.youtube.com/watch?v=93G_FqpGFGY
http://www.youtube.com/watch?v=WyfwRPBw62s

Good Link on Periodic Trends

Properties of Metal, Nonmetals, and Metalloids



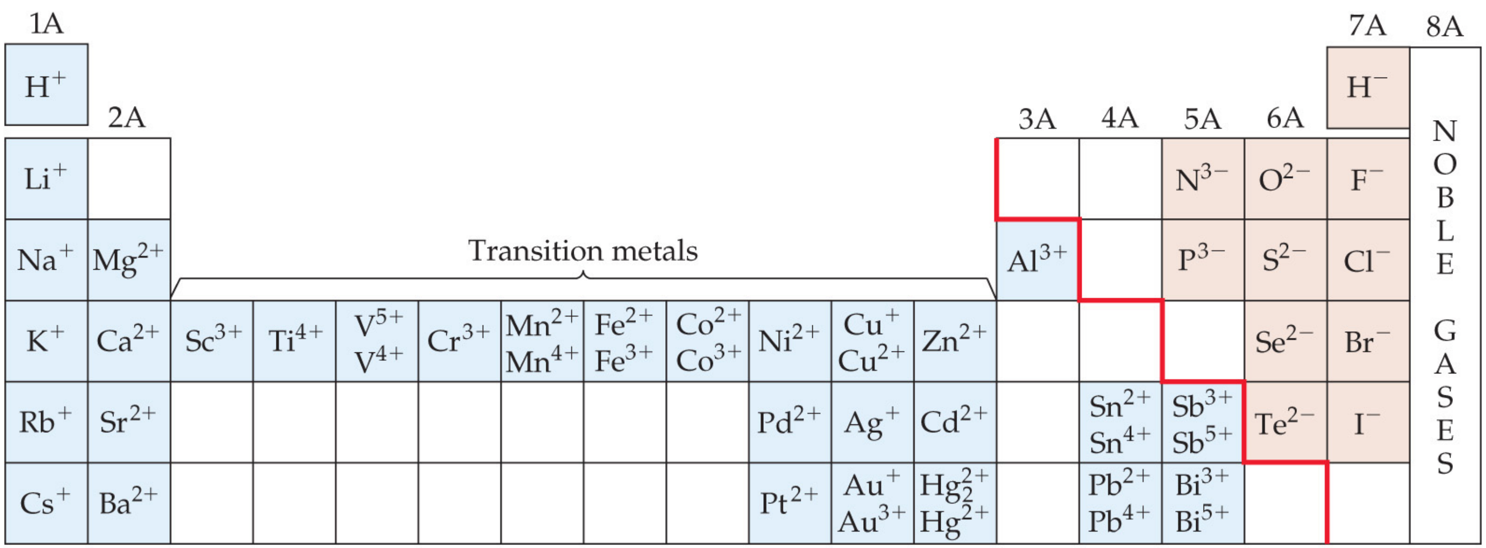
Metals versus Nonmetals

Metals	Nonmetals
Have a shiny luster; various colors, although most are silvery	Do not have a luster; various colors
Solids are malleable and ductile	Solids are usually brittle; some are hard, some are soft
Good conductors of heat and electricity	Poor conductors of heat and electricity
Most metal oxides are ionic solids that are basic	Most nonmetal oxides are molecular substances that form acidic solutions
Tend to form cations in aqueous solution	Tend to form anions or oxyanions in aqueous solution

Differences between metals and nonmetals tend to revolve around these properties.

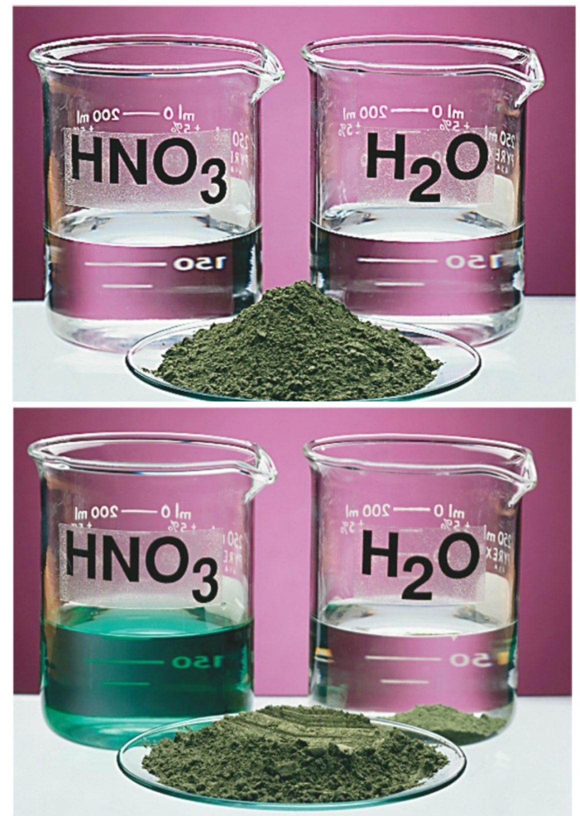
Metals versus Nonmetals

- Metals tend to form cations.
- Nonmetals tend to form anions.
- Metallic character increases down a group and decreases across a period.

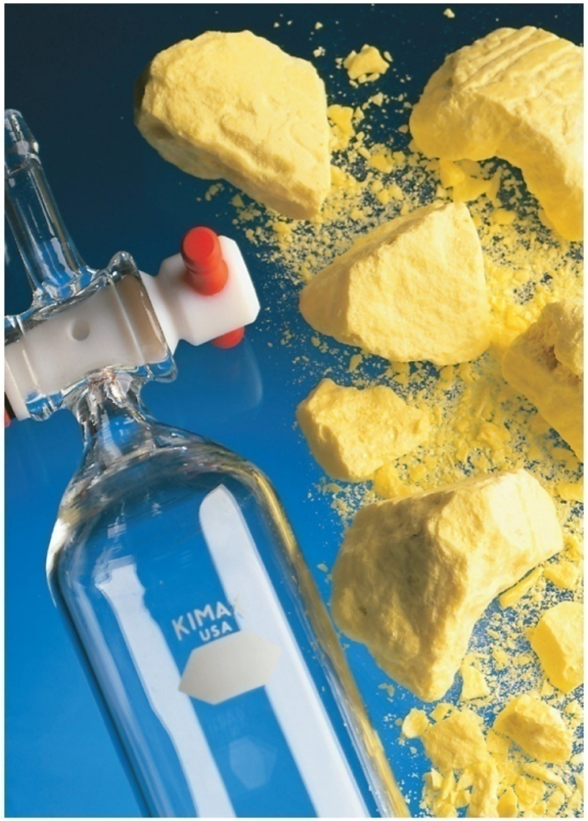


Metals

- Compounds formed between metals and nonmetals tend to be ionic.
- Metal oxides tend to be basic.



Nonmetals



- These are dull, brittle substances that are poor conductors of heat and electricity.
- They tend to gain electrons in reactions with metals to acquire a noble gas configuration.

Nonmetals

- Substances containing only nonmetals are molecular compounds.
- Most nonmetal oxides are acidic.

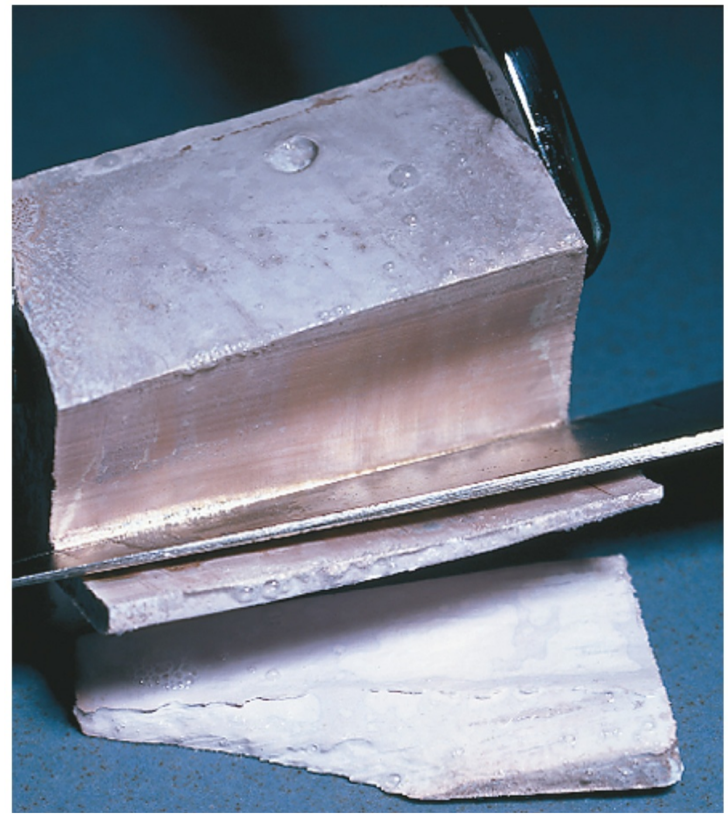
Metalloids

- These have some characteristics of metals and some of nonmetals.
- For instance, silicon looks shiny, but is brittle and fairly poor conductor.

Group Trends

Alkali Metals

- Alkali metals are soft, metallic solids.
- The name comes from the Arabic word for ashes.



Alkali Metals

- They are found only in compounds in nature, not in their elemental forms.
- They have low densities and melting points.
- They also have low ionization energies.

Element	Electron Configuration	Melting Point (°C)	Density (g/cm^3)	Atomic Radius (Å)	I_1 (kJ/mol)
Lithium	$[He]2s^1$	181	0.53	1.34	520
Sodium	$[Ne]3s^1$	98	0.97	1.54	496
Potassium	$[Ar]4s^1$	63	0.86	1.96	419
Rubidium	$[Kr]5s^1$	39	1.53	2.11	403
Cesium	$[Xe]6s^1$	28	1.88	2.25	376

Alkali Metals

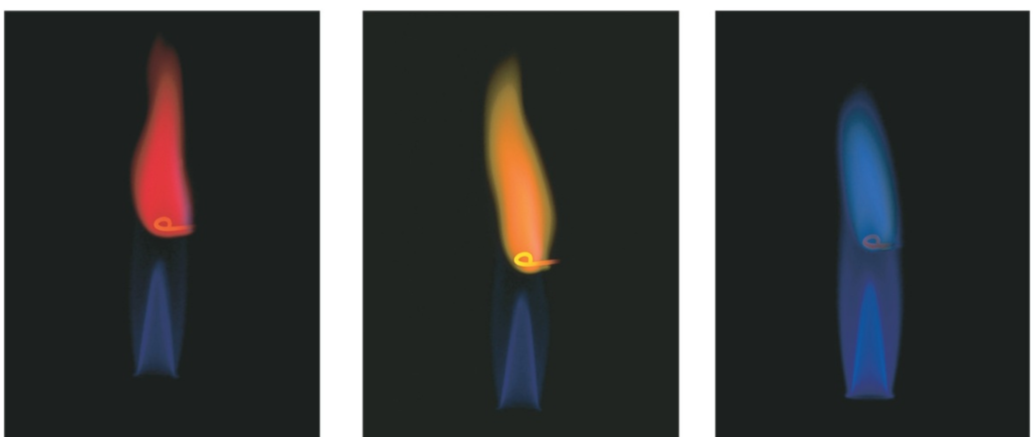
Their reactions with water are famously exothermic.

Alkali Metals

- Alkali metals (except Li) react with oxygen to form peroxides.
- K, Rb, and Cs also form superoxides:

$$K + O_2 \longrightarrow KO_2$$

- They produce bright colors when placed in a flame.



Alkaline Earth Metals

Element	Electron Configuration	Melting Point (°C)	Density (g/cm^3)	Atomic Radius (Å)	I_1 (kJ/mol)
Beryllium	$[He]2s^2$	1287	1.85	0.90	899
Magnesium	$[Ne]3s^2$	650	1.74	1.30	738
Calcium	$[Ar]4s^2$	842	1.55	1.74	590
Strontium	$[Kr]5s^2$	777	2.63	1.92	549
Barium	$[Xe]6s^2$	727	3.51	1.98	503

- Alkaline earth metals have higher densities and melting points than alkali metals.
- Their ionization energies are low, but not as low as those of alkali metals.

Alkaline Earth Metals

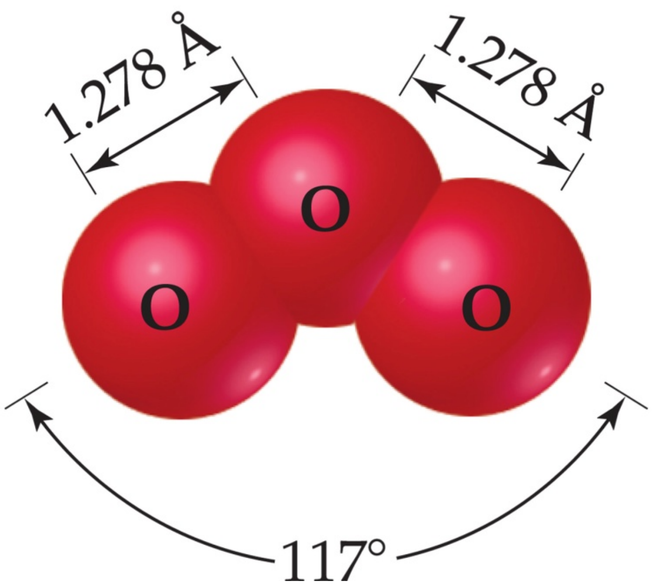
- Beryllium does not react with water and magnesium reacts only with steam, but the others react readily with water.
- Reactivity tends to increase as you go down the group.

Group 6A

Element	Electron Configuration	Melting Point (°C)	Density	Atomic Radius (Å)	I_1 (kJ/mol)
Oxygen	$[He]2s^2 2p^4$	−218	1.43 g/L	0.73	1314
Sulfur	$[Ne]3s^2 3p^4$	115	1.96 g/cm^3	1.02	1000
Selenium	$[Ar]3d^{10} 4s^2 4p^4$	221	4.82 g/cm^3	1.16	941
Tellurium	$[Kr]4d^{10} 5s^2 5p^4$	450	6.24 g/cm^3	1.35	869
Polonium	$[Xe]4f^{14} 5d^{10} 6s^2 6p^4$	254	9.20 g/cm^3	—	812

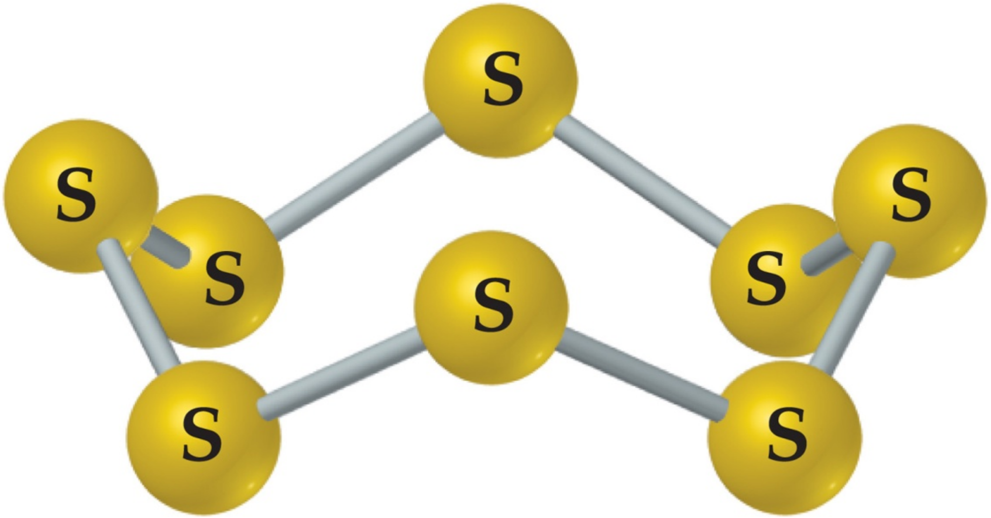
- Oxygen, sulfur, and selenium are nonmetals.
- Tellurium is a metalloid.
- The radioactive polonium is a metal.

Oxygen



- There are two allotropes of oxygen:
  - O_2
  - O_3, ozone
- There can be three anions:
  - O^{2-}, oxide
  - O_2^{2-}, peroxide
  - O_2^{1-}, superoxide
- It tends to take electrons from other elements (oxidation).

Sulfur



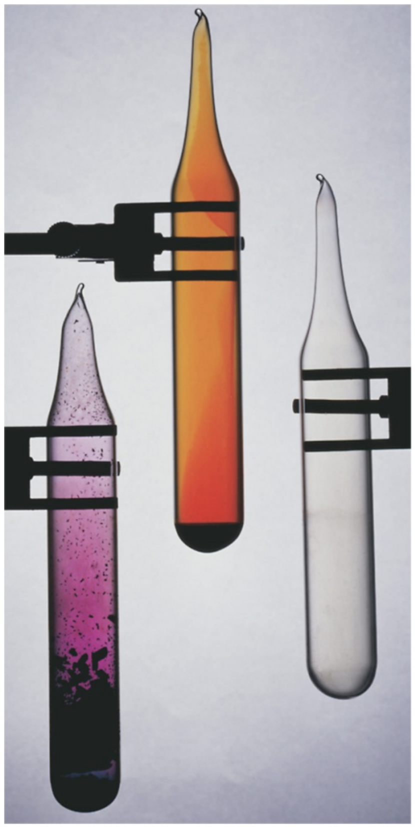
- Sulfur is a weaker oxidizer than oxygen.
- The most stable allotrope is S_8, a ringed molecule.

Group VIIA: Halogens

Element	Electron Configuration	Melting Point (°C)	Density	Atomic Radius (Å)	I_1 (kJ/mol)
Fluorine	$[He]2s^2 2p^5$	−220	1.69 g/L	0.71	1681
Chlorine	$[Ne]3s^2 3p^5$	−102	3.12 g/L	0.99	1251
Bromine	$[Ar]3d^{10} 4s^2 4p^5$	−7.3	3.12 g/cm^3	1.14	1140
Iodine	$[Kr]4d^{10} 5s^2 5p^5$	114	4.94 g/cm^3	1.33	1008

- The halogens are prototypical nonmetals.
- The name comes from the Greek words *halos* and *gennao*: "salt formers".

Group VIIA: Halogens



- They have large, negative electron affinities.
  - Therefore, they tend to oxidize other elements easily.
- They react directly with metals to form metal halides.
- Chlorine is added to water supplies to serve as a disinfectant

Group VIIIA: Noble Gases

Element	Electron Configuration	Boiling Point (K)	Density (g/L)	Atomic Radius* (Å)	I_1 (kJ/mol)
Helium	$1s^2$	4.2	0.18	0.32	2372
Neon	$[He]2s^22p^6$	27.1	0.90	0.69	2081
Argon	$[Ne]3s^23p^6$	87.3	1.78	0.97	1521
Krypton	$[Ar]3d^{10}4s^24p^6$	120	3.75	1.10	1351
Xenon	$[Kr]4d^{10}5s^25p^6$	165	5.90	1.30	1170
Radon	$[Xe]4f^{14}5d^{10}6s^26p^6$	211	9.73	1.45	1037

*Only the heaviest of the noble-gas elements form chemical compounds. Thus, the atomic radii for the lighter noble-gas elements are estimated values.

- The noble gases have astronomical ionization energies.
- Their electron affinities are positive.
  - Therefore, they are relatively unreactive.
- They are found as monatomic gases.

Group VIIIA: Noble Gases

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