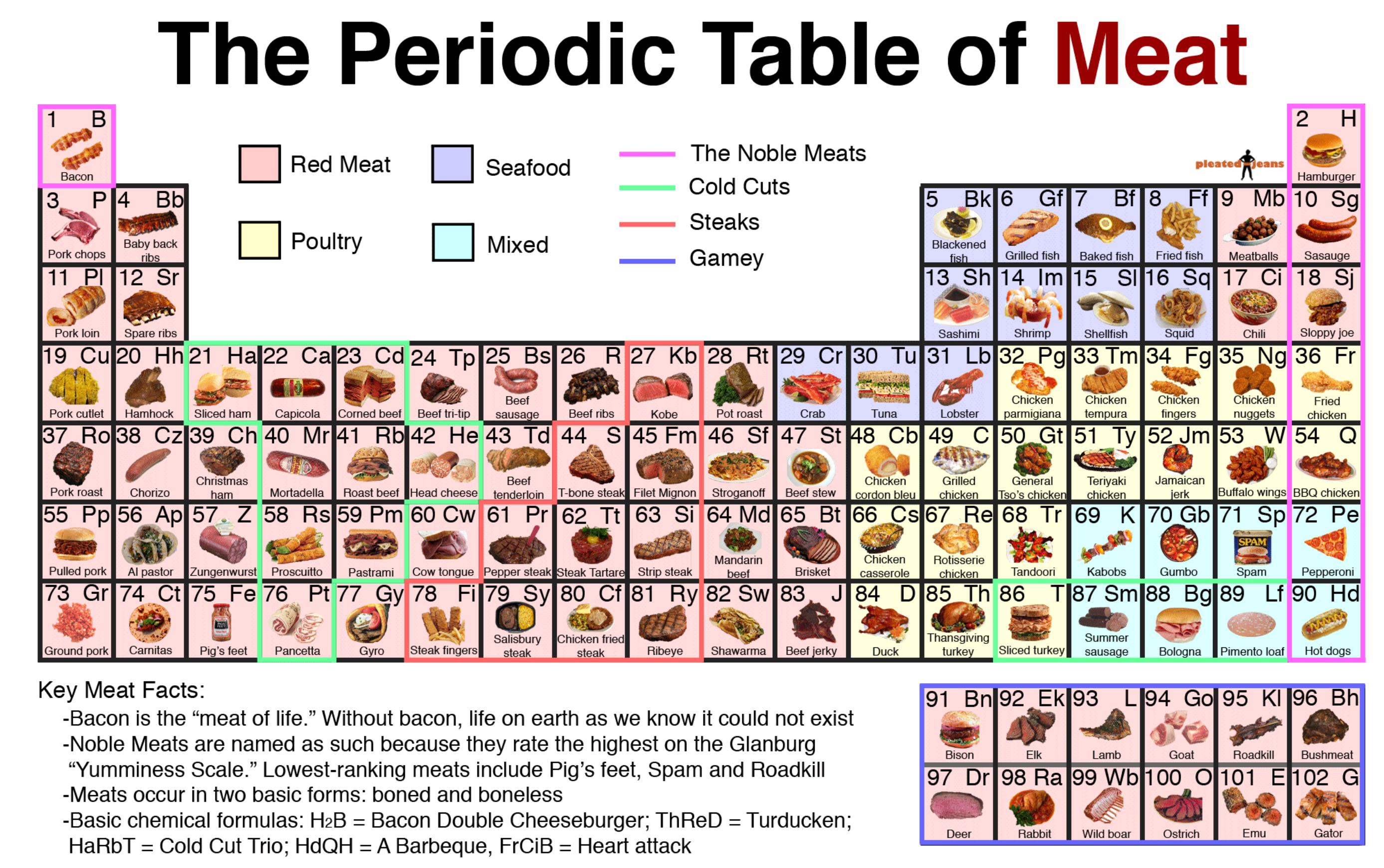
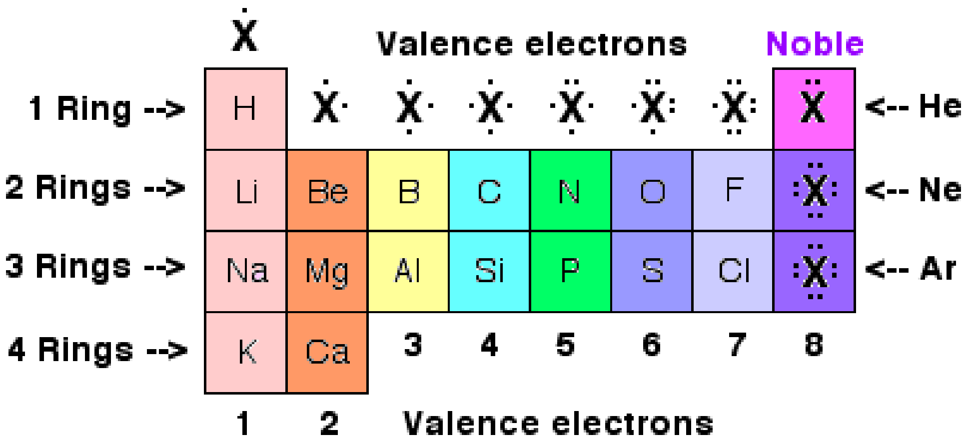
AP Chapter 7 – Periodic Properties of the Elements



Jennie L. Borders

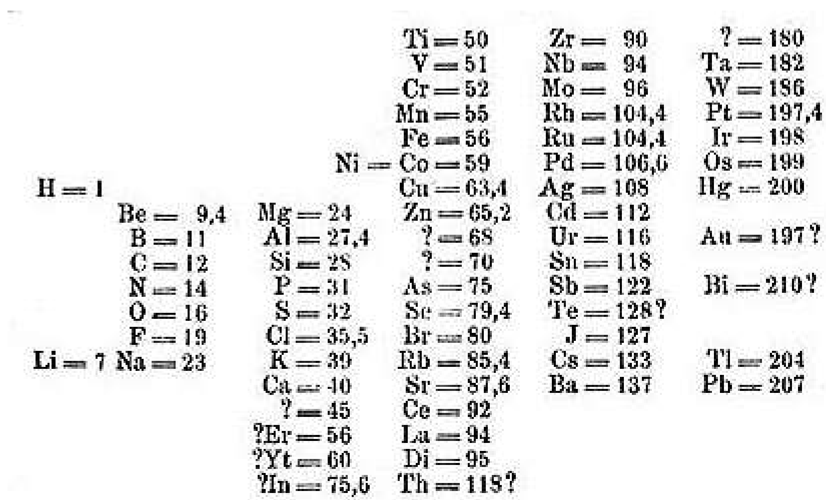
Section 7.1 – Development of the Periodic Table

- The periodic nature of the table arises from the repeating patterns in the electron configurations of the elements.
- Elements in the same column of the table contain the same number of valence electrons.



Mendeleev’s Table

- Mendeleev created the “first” periodic table by arranging the elements with increasing atomic weight.
- Mendeleev predicted properties of undiscovered elements based on their location on his periodic table.



			Ti = 50	Zr = 90	? = 180
			V = 51	Nb = 94	Ta = 182
			Cr = 52	Mo = 96	W = 186
			Mn = 55	Rh = 104,4	Pt = 197,4
			Fe = 56	Ru = 104,4	Ir = 198
			Ni — Co = 59	Pd = 106,6	Os = 199
H = 1			Cu = 63,4	Ag = 108	Hg = 200
	Be = 9,4	Mg = 24	Zn = 65,2	Cd = 112	
	B = 11	Al = 27,4	? = 68	Ur = 116	Au = 197?
	C = 12	Si = 28	? = 70	Sn = 118	
	N = 14	P = 31	As = 75	Sb = 122	Bi = 210?
	O = 16	S = 32	Se = 79,4	Te = 128?	
	F = 19	Cl = 35,5	Br = 80	J = 127	
Li = 7	Na = 23	K = 39	Rb = 85,4	Cs = 133	Tl = 204
		Ca = 40	Sr = 87,6	Ba = 137	Pb = 207
		? = 45	Ce = 92		
		?Er = 56	La = 94		
		?Yt = 60	Di = 95		
		?In = 75,6	Th = 118?		

Atomic Number

- Henry Moseley studied the X-ray frequencies produced when elements were bombarded with high-energy electrons.
- Each element had a unique frequency which increased as atomic mass

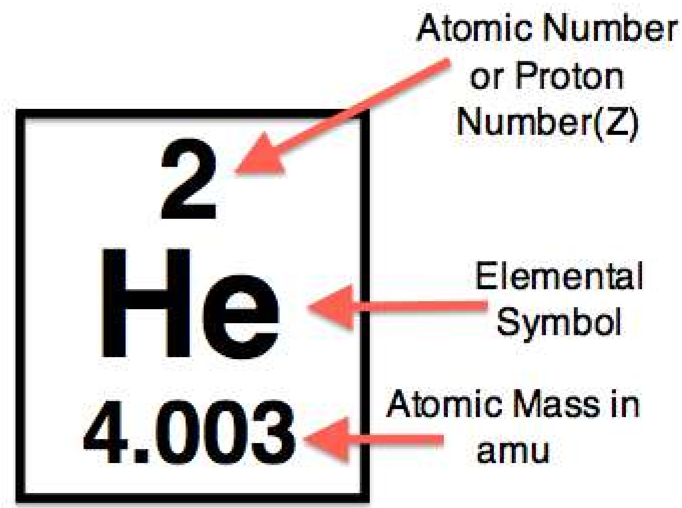
increased.

- From this data, Moseley

assigned each element

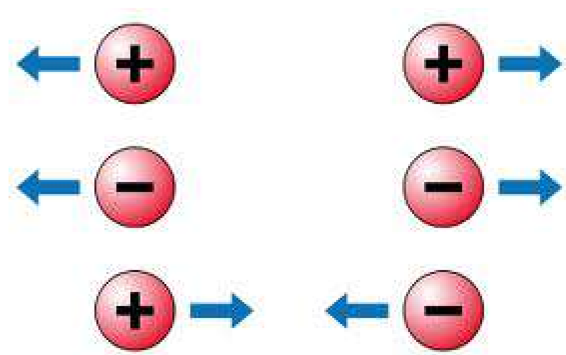
and atomic number

(the number of protons).



Section 7.2 – Effective Nuclear Charge

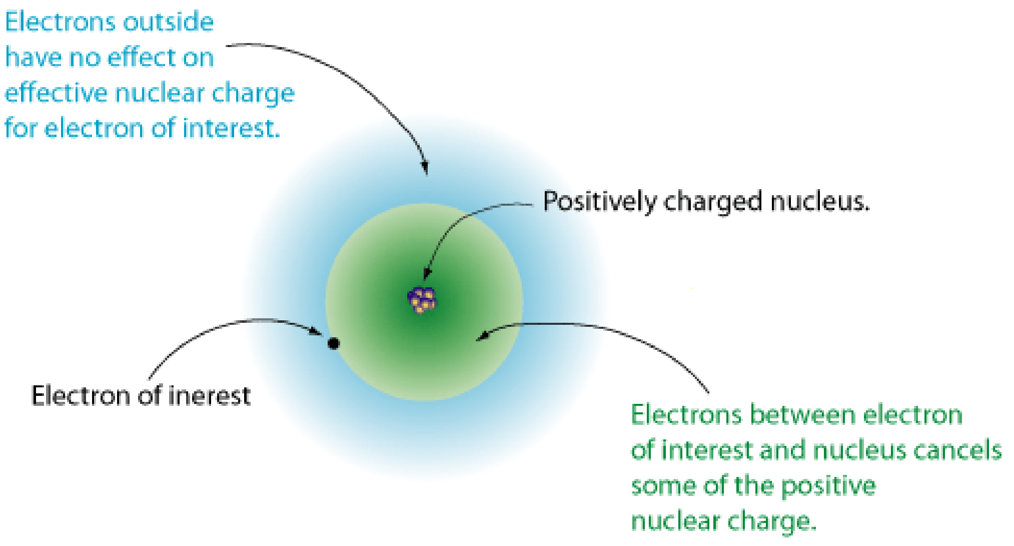
- Since electrons are negatively charged, they are attracted to nuclei, which are positively charged.
- Many of the properties of atoms depend on their electron configurations and on how strongly their outer electrons are attracted to the nucleus.



Electrons and the Nucleus

- The force of attraction between an electron and the nucleus is based on Coulomb's law and depends on two amounts:

1. The magnitude of the net nuclear charge acting on the electron
2. The average distance between the nucleus and the electron



Effective Nuclear Charge

- Effective nuclear charge is the actual nuclear charge minus the shielding of the core electrons, so Z_{eff} is always less than actual nuclear charge.

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

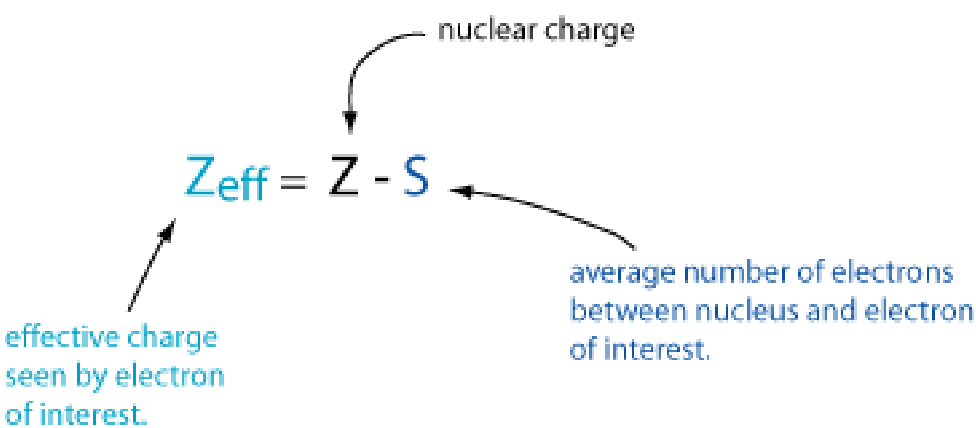
Z_{eff} = effective nuclear charge

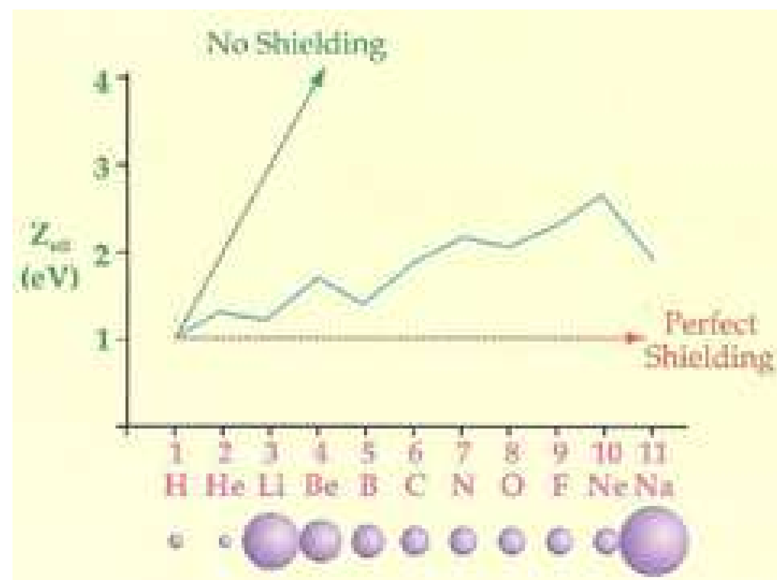
Z = actual nuclear charge (atomic number)

S = screening constant (number of core electrons)

Z_{eff} Trend Across

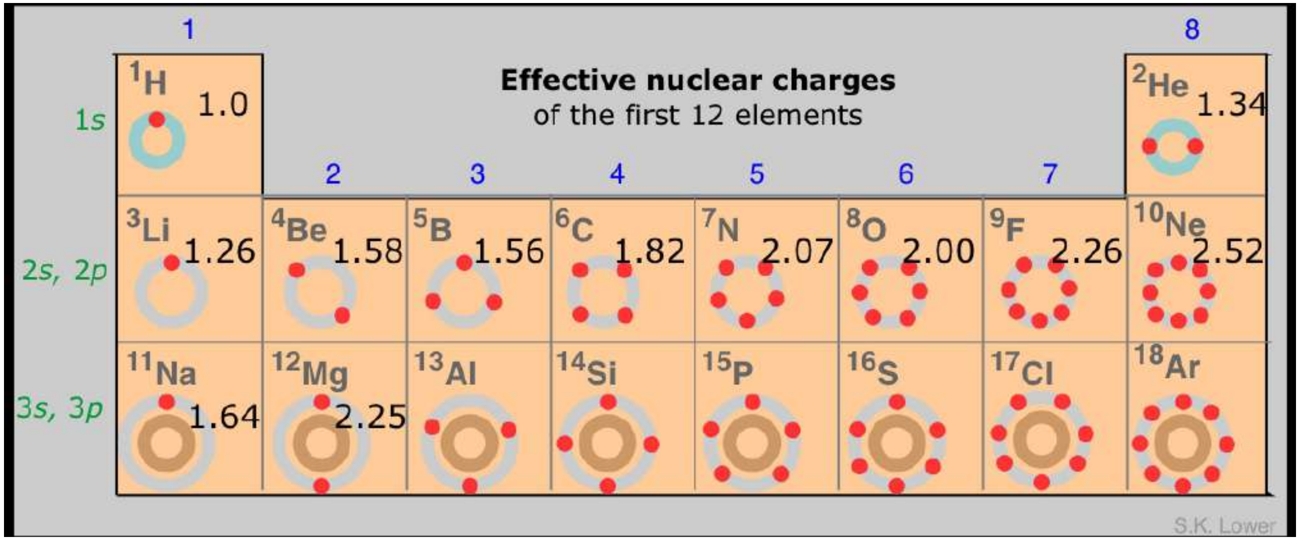
- The effective nuclear charge increases as we move across a row (period) on the table.
- Although the number of core electrons stays the same as we move across, the actual nuclear charge increases.





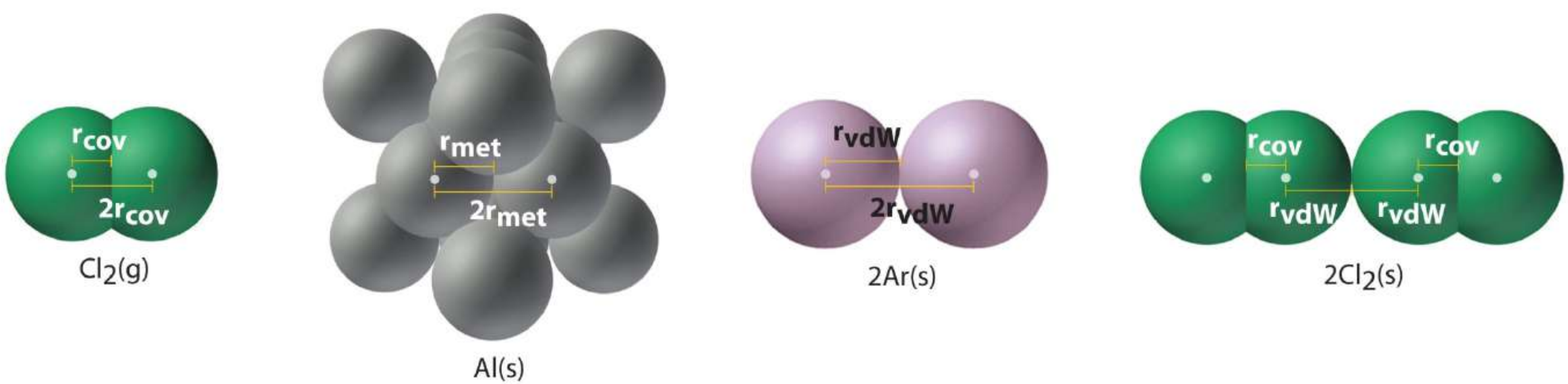
Z_{eff} Trend Down

- Going down a column, the effective nuclear charge increases slightly.
- Actual nuclear charge is increases as you move down a group, but the larger electron cores are less able the screen the valence electrons.



Section 7.3 – Sizes of Atoms and Ions

- Bonding atomic radius is measured by halving the distance between the nuclei of two bonded atoms.



(a) Covalent radius, r_{cov} **(b) Metallic radius, r_{met}** **(c) van der Waals radius, r_{vdW}** **(d) Covalent vs. vdW radii**

Sample Exercise 7.1

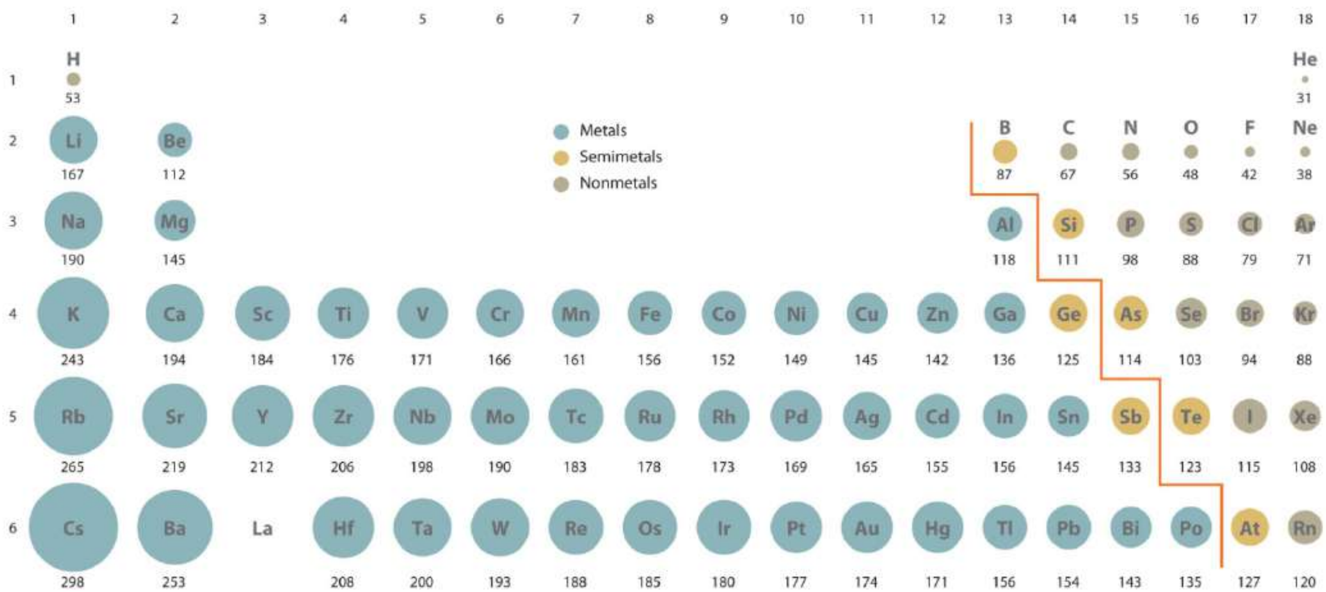
- Methyl mercaptan, CH_3SH, is used to help detect natural gas leaks. Use Figure 7.7 (p. 261) to predict the lengths of the C-S, C-H, and S-H bonds in this molecule.

Practice Exercise

- Using Figure 7.7 (p. 261), predict which will be greater, the P-Br bond length in PBr_3 or the As-Cl bond length in $AsCl_3$.

Atomic Radius Trend Down

- Within a group, atomic radius increases from top to bottom.
- This trend results primarily from the increase in the size of the energy level.



Atomic Radius Trend Across

- Within each period, atomic radius tends to decrease from left to right.
- The major factor influencing this trend is the increase in Z_{eff}. The increase in Z_{eff} draws the valence electrons closer to the nucleus.

Sample Exercise 7.2

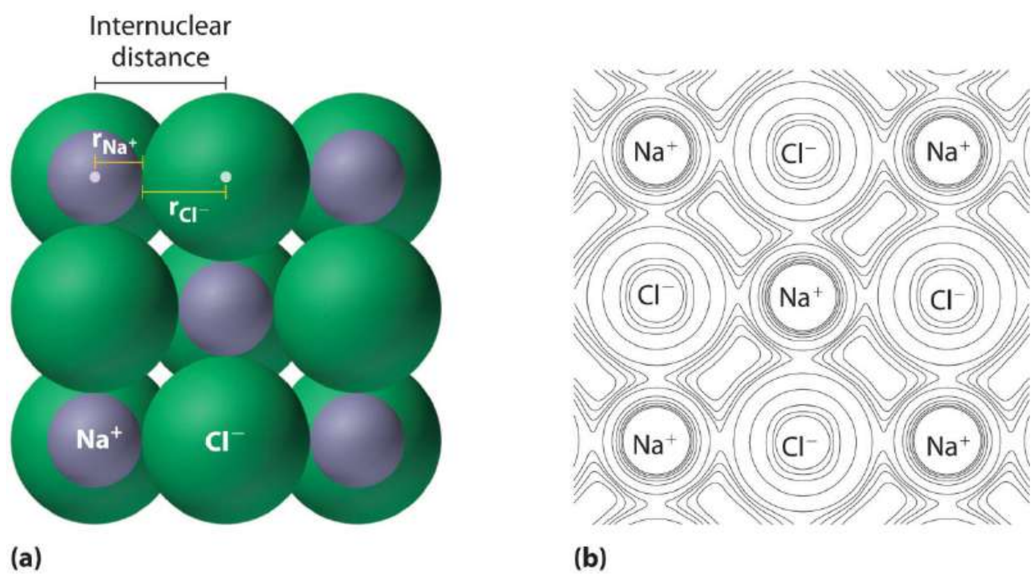
- Arrange the following atoms in order of increasing atomic size: P, S, As, and Se.

Practice Exercise

- Arrange the following atoms in order of increasing atomic radius: Na, Be, and Mg.

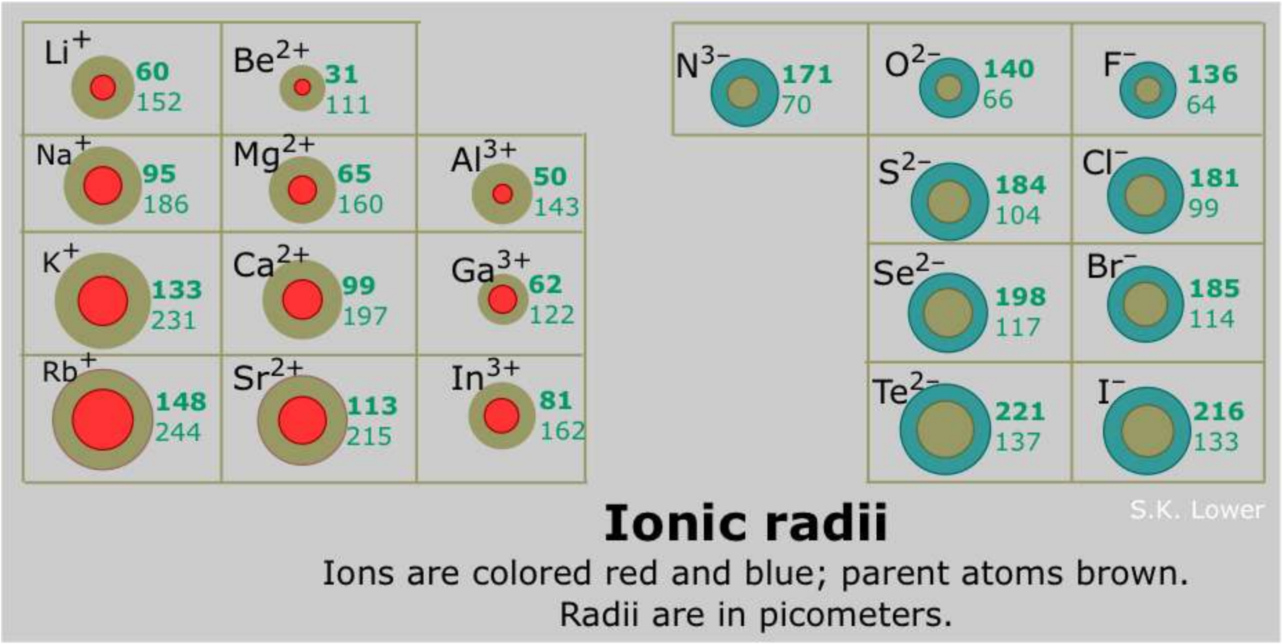
Cation Radii

- Cations are smaller than their parent atoms because electrons have been removed from the outer energy level, so the remaining electrons are pulled closer by the nucleus.



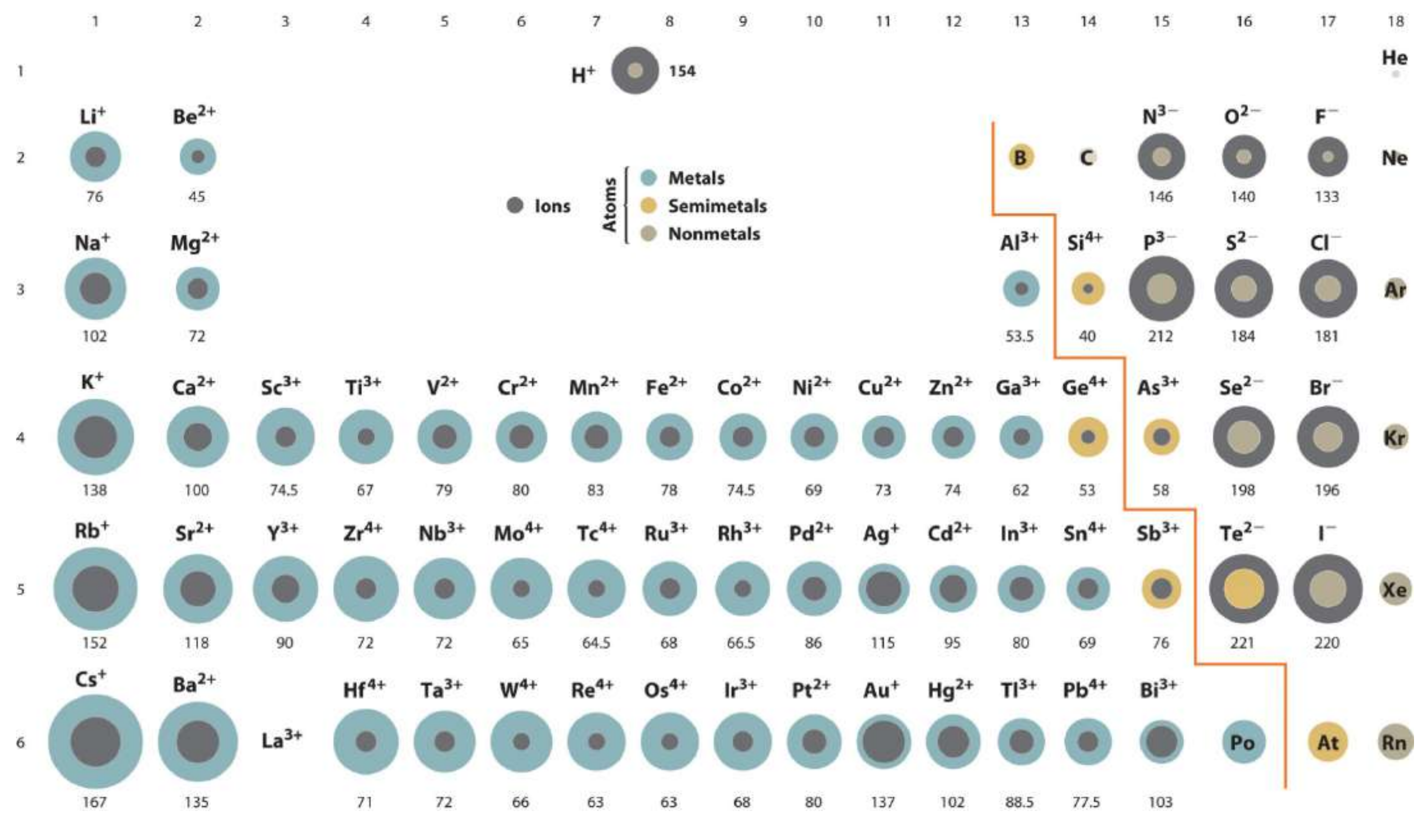
Anion Radii

- Anions are larger than their parent atoms because electrons are added which increase the electron-electron repulsion causing the electrons to spread further apart.



Ionic Radii Trend Down

- Ionic radii increases as we move down a group because the energy level increases.



Sample Exercise 7.3

- Arrange these atoms in order of decreasing size: Mg^{2+}, Ca^{2+}, and Ca.

Practice Exercise

- Which of the following atoms and ions is largest: S^{2-}, S, O^{2-}?

Isoelectronic Series

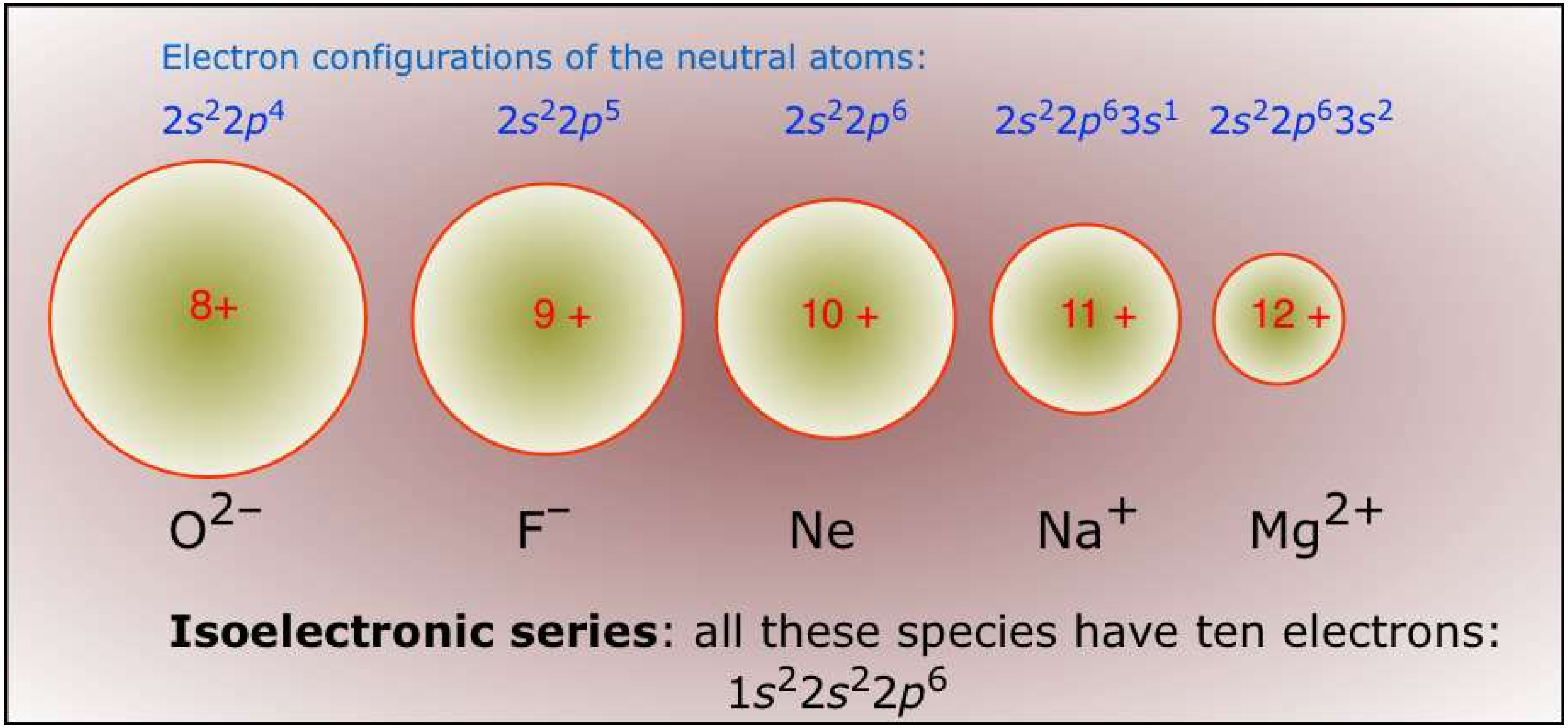
- An isoelectronic series is a group of ions all containing the same number of electrons.
- In an isoelectronic series we can list the members in order of increasing atomic number and the radius will decrease.

→Increasing atomic number→

O^{2-} F^{-} Na^{+} Mg^{2+} Al^{3+}

→Decreasing atomic radius→

Isoelectronic Series



Sample Exercise 7.4

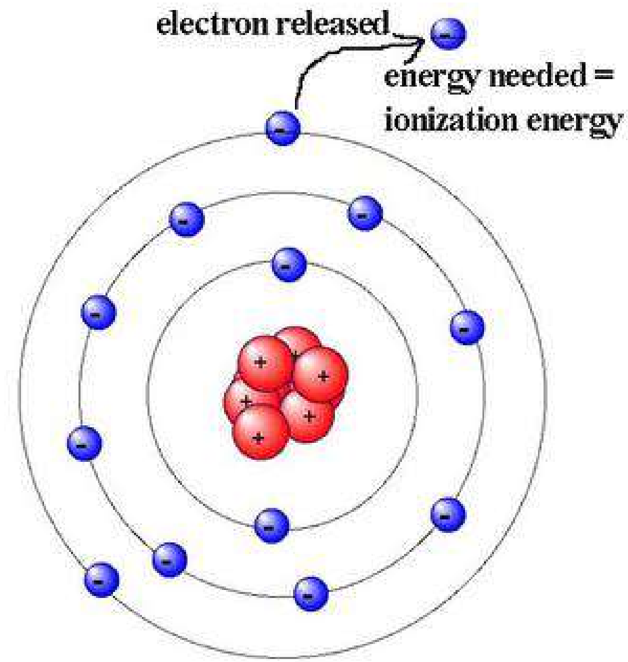
- Arrange the ions K^+, Cl^-, Ca^{2+}, and S^{2-} in order of decreasing size.

Practice Exercise

- Which of the following ions is largest, Rb^+, Sr^{2+}, or Y^{3+}?

Section 7.4 – Ionization Energy

- The ionization energy of an atom or ion is the amount of energy required to remove one electron from the ground state of the isolated gaseous atom or ion.



Ionization Energies

- Notice the values for a given element increase as successive electrons are removed: $I_1<I_2<I_3$, and so forth.
- This trend exists because each extra electron is being removed from an increasingly positive ion.

General increase

General decrease

Successive Ionization Energies for Period 3 Elements

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	498	4560	6910	9540	13 400	16 600	20 100
Mg	736	1445	7730	10 600	13 600	18 000	21 700
Al	577	1815	2740	11 600	15 000	18 310	23 290
Si	787	1575	3220	4350	16 100	19 800	23 800
P	1063	1890	2905	4950	6270	21 200	25 400
S	1000	2260	3375	4565	6950	8490	27 000
Cl	1255	2295	3850	5160	6560	9360	11 000
Ar	1519	2665	3945	5770	7230	8780	12 000

Ionization Energies

- There is a sharp increase in the ionization energy that occurs when an inner-shell electron is removed.
- This is the reason that only valence electrons are involved in bonding.

Table 2.2 The Ionization Energies ($kJ\ mol^{-1}$) of the First Twenty Elements

Z	Element	First	Second	Third	Fourth	Fifth	Sixth
1	H	1,312					
2	He	2,373	5,251				
3	Li	520	7,300	11,815			
4	Be	899	1,757	14,850	21,005		
5	B	801	2,430	3,660	25,000	32,820	
6	C	1,086	2,350	4,620	6,220	38,000	47,261
7	N	1,400	2,860	4,580	7,500	9,400	53,000
8	O	1,314	3,390	5,300	7,470	11,000	13,000
9	F	1,680	3,370	6,050	8,400	11,000	15,200
10	Ne	2,080	3,950	6,120	9,370	12,200	15,000
11	Na	495.9	4,560	6,900	9,540	13,400	16,600
12	Mg	738.1	1,450	7,730	10,500	13,600	18,000
13	Al	577.9	1,820	2,750	11,600	14,800	18,400
14	Si	786.3	1,580	3,230	4,360	16,000	20,000
15	P	1,012	1,904	2,910	4,960	6,240	21,000
16	S	999.5	2,250	3,360	4,660	6,990	8,500
17	Cl	1,251	2,297	3,820	5,160	6,540	9,300
18	Ar	1,521	2,666	3,900	5,770	7,240	8,800
19	K	418.7	3,052	4,410	5,900	8,000	9,600
20	Ca	589.5	1,145	4,900	6,500	8,100	11,000

Sample Exercise 7.5

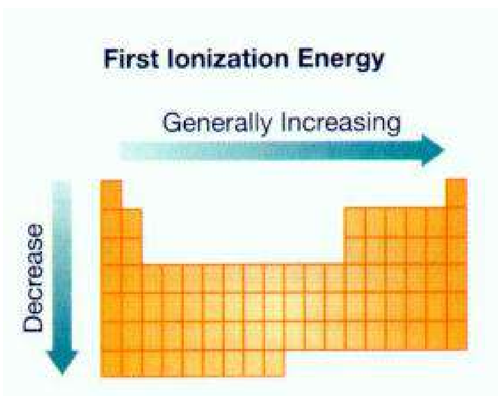
- Based on the locations of sodium, calcium, and sulfur, predict the one with the largest second ionization energy.

Practice Exercise

- Which will have the greater third ionization energy, Ca or S?

Ionization Energy Trend Across

- Ionization energy generally increases as you move across a period.
- This trend occurs because moving across the periodic table, Z_{eff} increases and the radius of the atom decreases, so more energy is required to remove an electron.



Ionization Energy Trend Down

- Ionization Energy decreases as you move down a group.
- This trend occurs because the electrons are farther from the nucleus since they are in larger energy levels,

so the energy

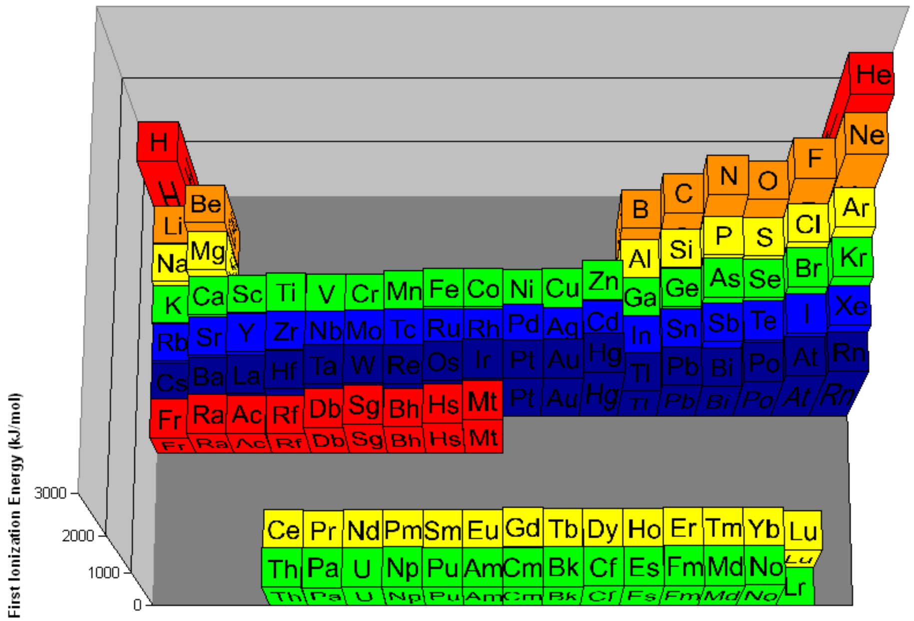
required to remove

one is lower.

Ionization Energy Trend d-block

- The change in ionization energy is greater in the s and p blocks of the periodic table. The energies only increases slightly across the d and f blocks.

Periodic Trends in First Ionization Energy



Sample Exercise 7.6

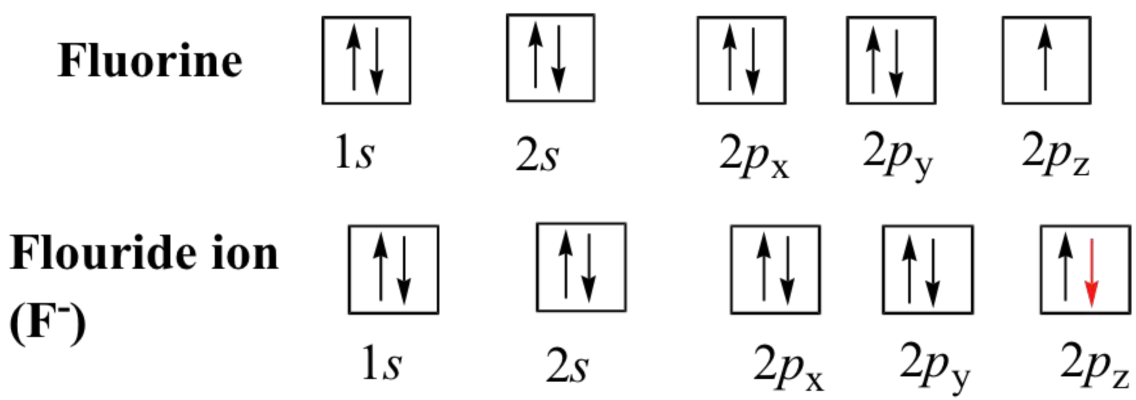
- Arrange the following atoms in order of increasing first ionization energy: Ne, Na, P, Ar, K.

Sample Exercise

- Which has the lowest first ionization energy, B, Al, C, or Si? Which has the highest ionization energy?

Electron Configuration for Ions

- When electrons are removed from a cation, they are removed from the highest sublevel in the highest principle energy level.
- When electrons are added to an anion, they are added to the last partially filled orbital or the next sublevel.



Sample Exercise 7.7

- Write the electron configuration for the following:

a. Ca^{2+}

b. Co^{3+}

c. S^{2-}

Sample Exercise

- Write the electron configuration for the following:

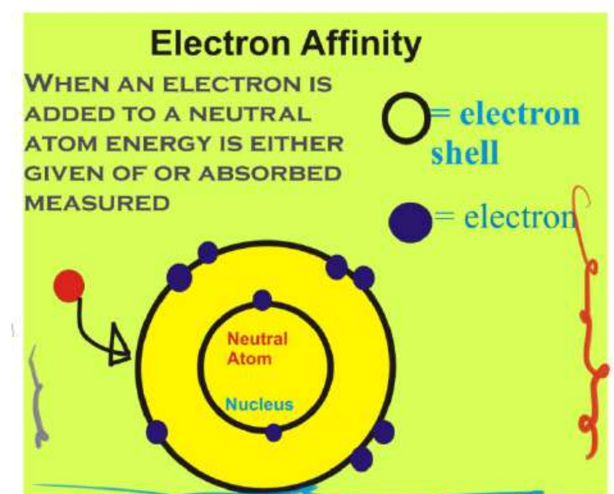
a. Ga^{3+}

b. Cr^{3+}

c. Br^{-}

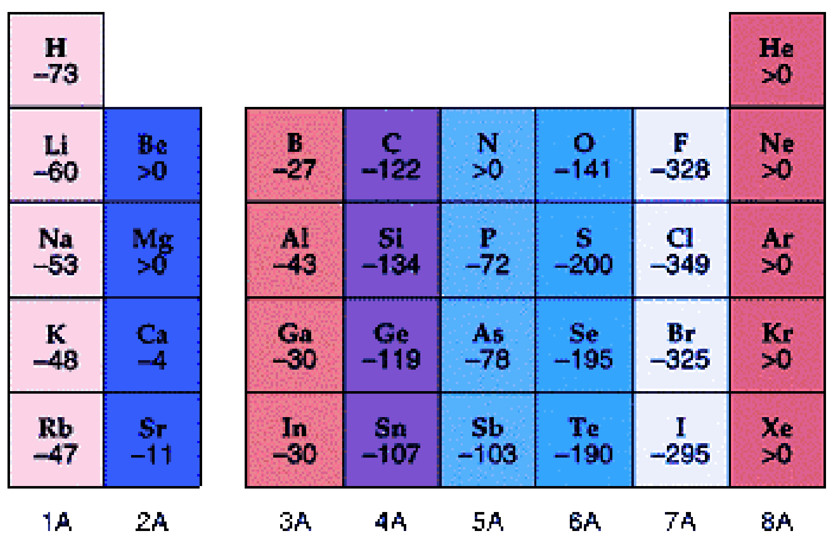
Section 7.5 – Electron Affinities

- The energy change that occurs when an electron is added to a gaseous atom is called the electron affinity. If measure the attraction of the atom for the added electron.
- For most atoms, energy is released when an electron is added.



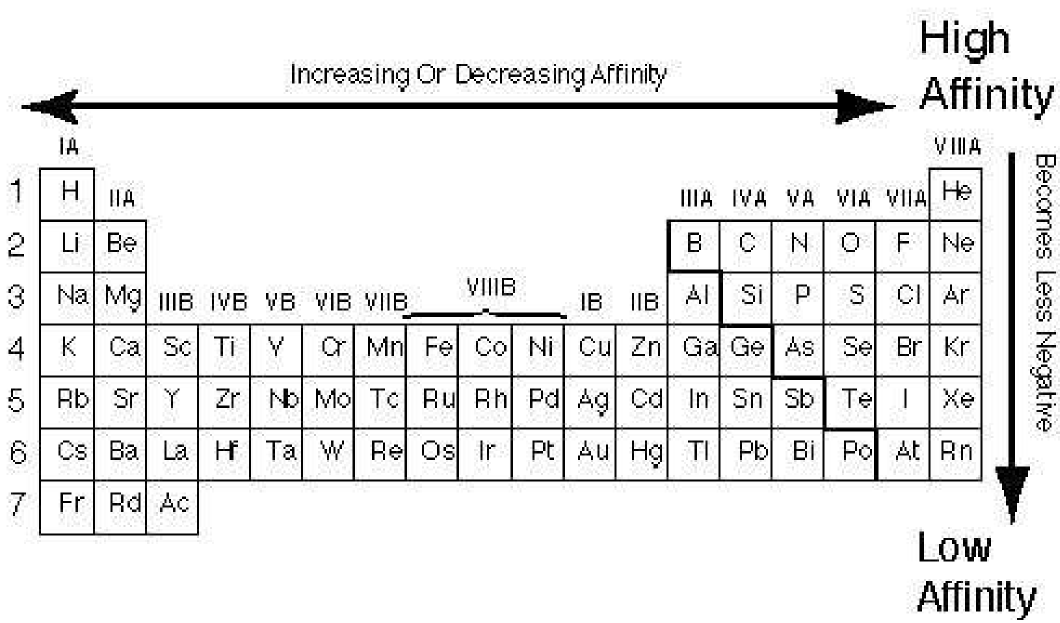
Electron Affinity Trend Across

- Electron affinity does not have a consistent trend going across a period. It does overall increase as you move across.
- This trend occurs due to the increased Z_{eff} and smaller atomic radius.

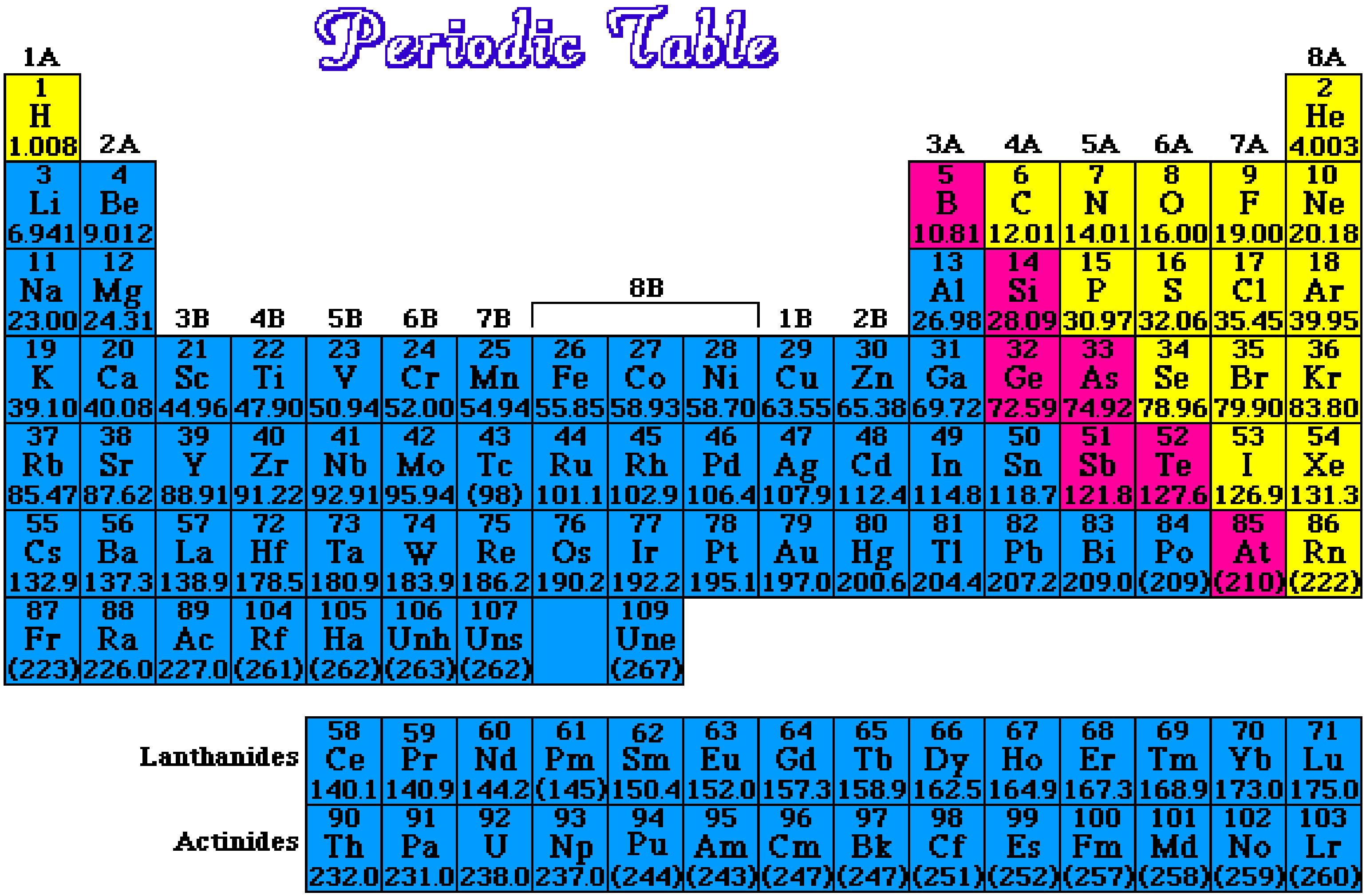


Electron Affinity Trend Down

- The electron affinity slightly decreases as you move down a group.
- This trend is due to the larger energy levels that are farther from the nucleus.



Section 7.6 – Metals, Nonmetals, and Metalloids



Properties of Metals

- Solids at room temperature (except Hg)
- Low ionization energies, so they form cations
- Metals are oxidized (lost electrons) in reactions
- Malleable
- Ductile
- Good conductors

- Form ionic compounds with nonmetals
- Metal oxides are basic (ex. Na_2O)

Sample Exercise 7.8

a. Would you expect scandium oxide to be a solid, liquid, or gas at room temperature?

b. Write the balanced equation for the reaction of scandium oxide (Sc^{3+}) with nitric acid.

Practice Exercise

- Write the balance equation for the reaction between copper (II) oxide and sulfuric acid.

Properties of Nonmetals

- Tend to be gases at room temperature
- High electron affinities, so they form anions
- Nonmetals are reduced (gain electrons) in reactions
- Brittle
- Poor conductors
- Form molecular compounds with nonmetals
- Nonmetal oxides are acidic

Sample Exercise 7.9

- Write the balanced equation for the reactions of solid selenium dioxide and the following:

a. water

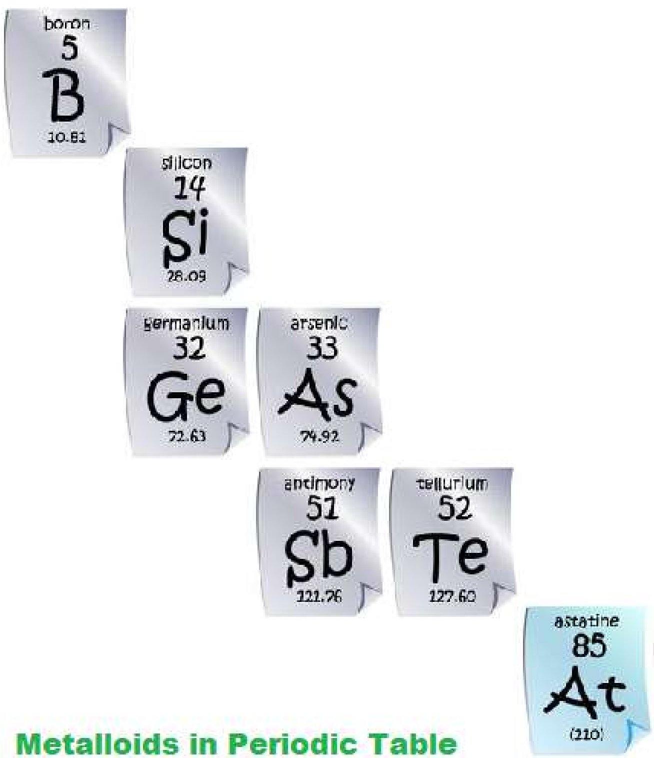
b. aqueous sodium hydroxide

Practice Exercise

- Write the balanced equation for the reaction of solid tetraphosphorus hexoxide with water.

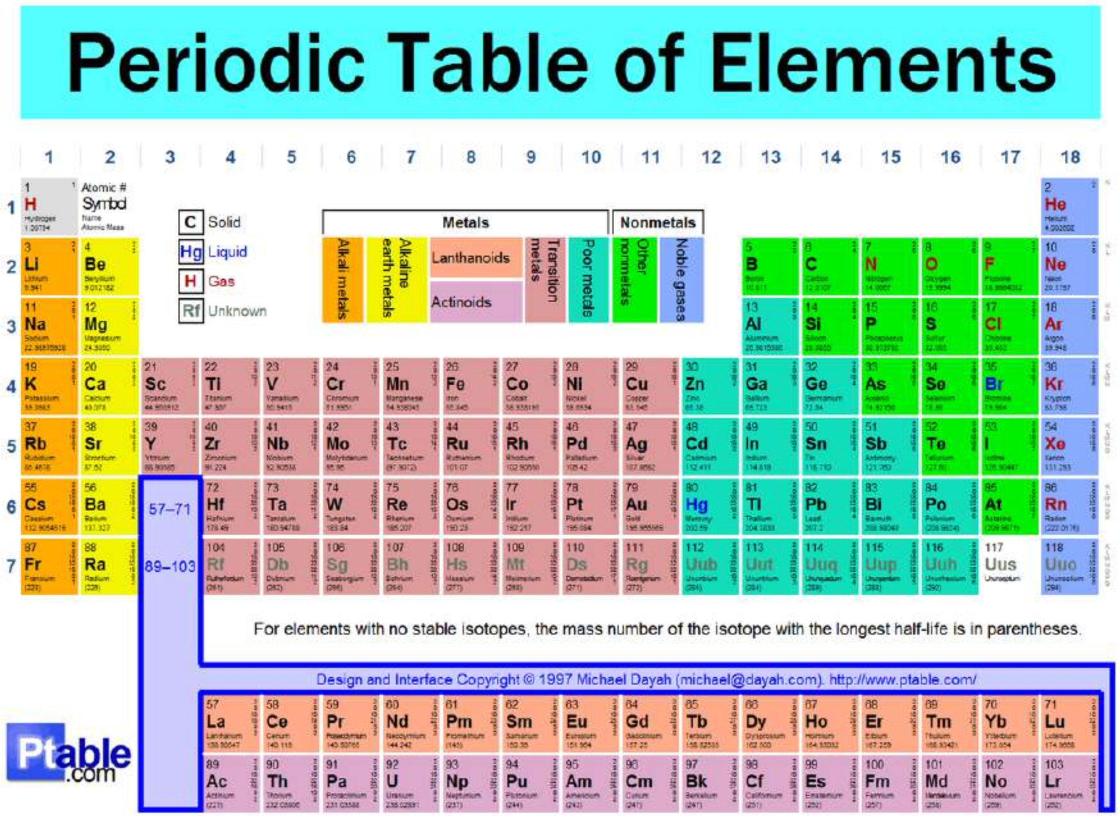
Metalloids

- Metalloids have properties intermediate between those of metals and nonmetals.



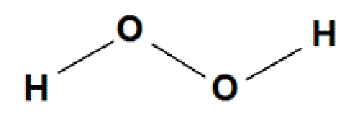
Section 7.8 – Group Trends for the Active Metals

- The alkali metals (group 1) and the alkaline earth metals (group 2) are considered the active metals.



Properties of Alkali Metals

- Soft metallic solids with low melting points
- When bonded with hydrogen, the hydrogen has a -1 charge (hydride). Ex: LiH
- React vigorously with water
- When bonded with oxygen, they can form the following:
  a. Oxide (O^{2-}) = Li_2O
  b. Peroxide (O_2^{2-}) = Na_2O_2
  c. Superoxide (O_2^{-}) = KO_2



HYDROGEN PEROXIDE

Sample Exercise 7.10

- Write a balanced equation that predicts the reaction of cesium metal with the following:

a. Cl_2

b. H_2O

c. H_2

Practice Exercise

- Write a balanced equation for the reaction between potassium metal and sulfur.

Properties of Alkaline Earth Metals

- Harder and more dense than alkali metals
- Higher melting point than alkali metals
- Less reactive with water than alkali metals

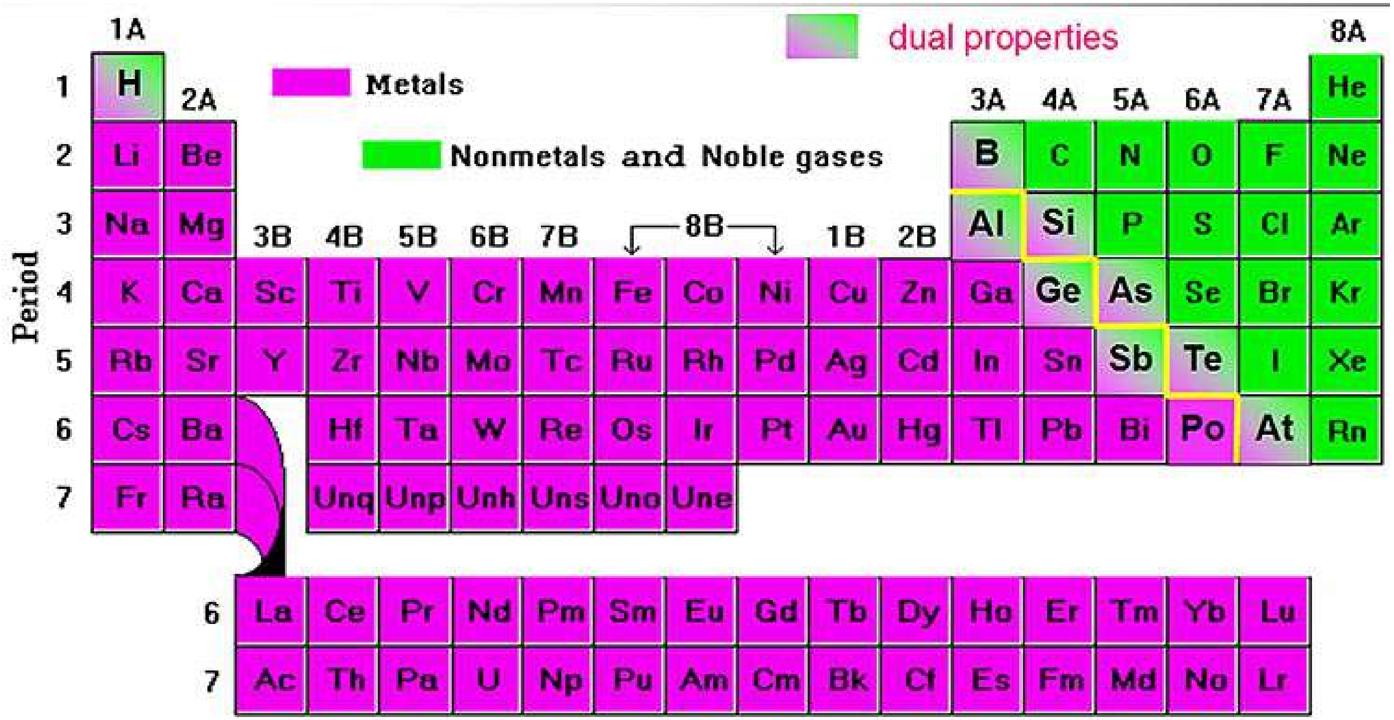
alkali metalalkaline earth metal

Section 7.8 – Group Trends for Selected Nonmetals

- Hydrogen is in the alkali metal group even though it is a nonmetal. It can be metallic under extremely high pressures.
- It belongs in group 1 because it has 1 valence electron and can form a +1 charge.
- It belongs in group 17 because it is a nonmetal, it can form a -1 charge, and it only needs 1 more electron to achieve noble gas confiuguration.

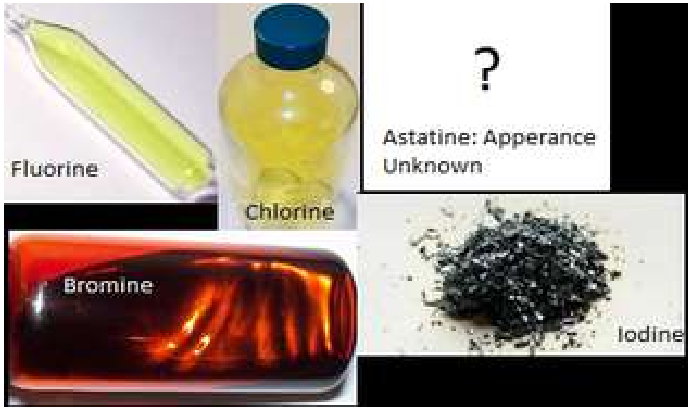
Nonmetal Groups

- When going down a group of nonmetals, the elements go from nonmetallic to metallic in nature.



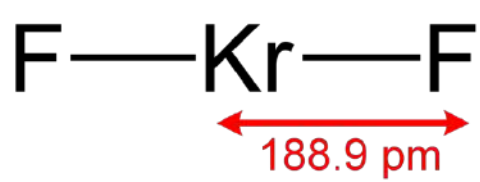
Group 17 - Halogens

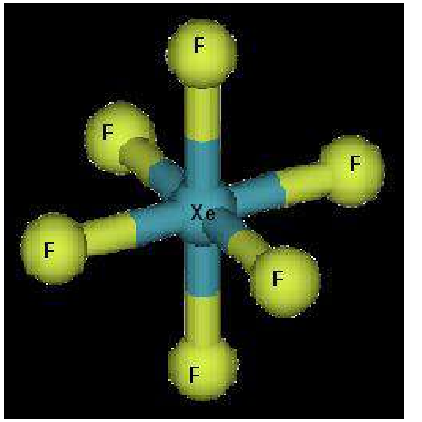
- Fluorine is a pale yellow gas.
- Chlorine is a greenish-yellow gas.
- Bromine is a reddish-brown liquid.
- Iodine is a gray/black solid that forms purple vapors.
- Halogens are very reactive.



Group 18 – Noble Gases

- Noble gases have full outer energy levels, so they are very unreactive.
- They rarely form compounds, but compounds have been formed with xenon, krypton, and argon. Most of these compounds contain fluorine, since it is highly reactive.





Sample Integrative Exercise

a. Bismuth subsalicylate is the active ingredient in Pepto-Bismol. The covalent atomic radii of thallium and lead are 1.48 angstrom and 1.47 angstrom. Using these values, predict the radius of the element bismuth.

Sample Integrative Exercise

b. What accounts for the general increase in atomic radius going down group 15?

Sample Integrative Exercise

c. Bismuth is used in low-melting alloys. The element itself is a brittle white crystalline solid. How do these characteristics fit with the fact that bismuth is in the same periodic group with such nonmetallic elements like nitrogen and phosphorus?

Sample Integrative Exercise

d. Bi_2O_3 is a basic oxide. Write a balanced equation for its reaction with dilute nitric acid. If 6.77g of Bi_2O_3 is dissolved in dilute acidic solution to make 0.500L of solution, what is the molarity of the solution of Bi^{3+} ion?

Sample Integrative Exercise

e. ^{209}Bi is the heaviest stable isotope of any element. How many protons and neutrons are present in the nucleus?

Sample Integrative Exercise

f. The density of Bi at 25°C is 9.808 g/cm^3. How many Bi atoms are present in a cube of the element that is 5.00cm on each edge? How many moles of the element are present?