# Grade 11 Chemistry

Matter, Trends, and Chemical Bonding

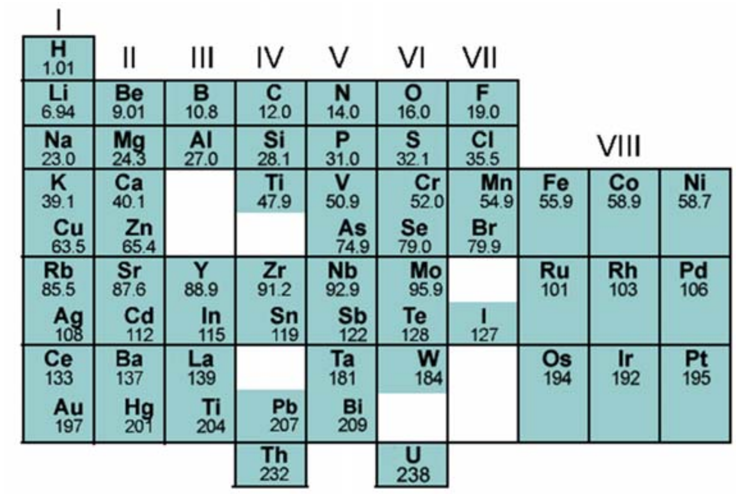
Class 3

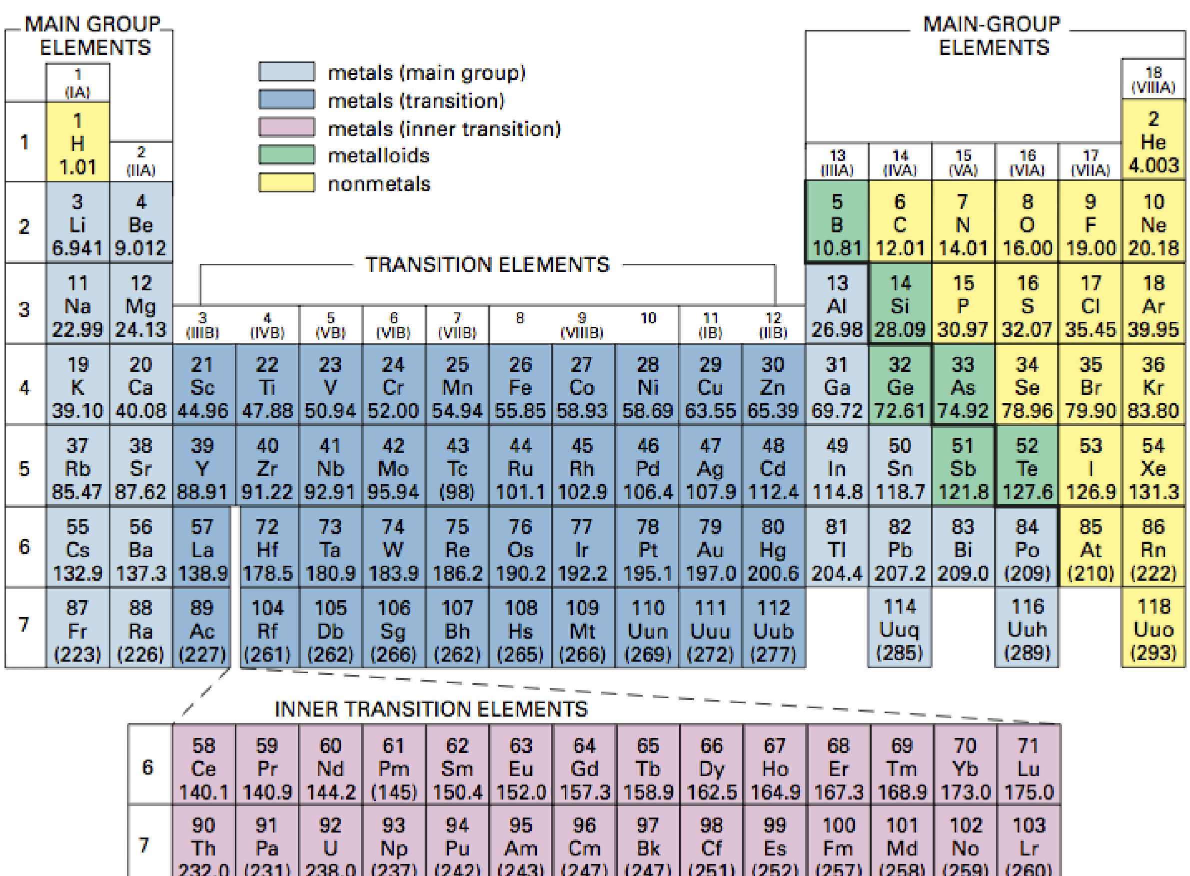
# Periodic Table

- First organized based on atomic masses
- 1864, Newlands noticed the Law of Octaves
- 1869, Mendeleev and Meyer grouped elements based on physical and chemical properties
- Modern periodic table organizes elements based on atomic number
- Patterns in the periodic table repeat → called the periodic (repeating) table

- **Periodic Law**: the chemical and physical properties of the elements repeat in a regular, periodic pattern when they are arranged according to their atomic number (Z)

Mendeleev's first published periodic table

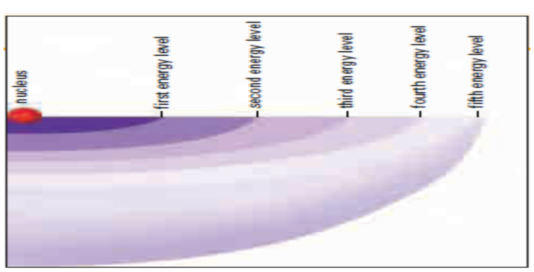


- Groups are numbered according to two different systems
  - 1-18: current number system
  - I to VIII: older system and separates them into two categories, A and B
- Elements in the A groups are the main group elements (**representative elements**)
- Elements in the B groups are the **transition elements**
- The rare earth metals (inner transition metals) fit between Group IIIB and IVB

# Trends in the Periodic Table

- Why do elements in the same group have similar properties?
  - Due to the number and arrangement of electrons in the atoms of each element
- Electron movement is restricted to fixed regions called energy shells or energy levels



- An element’s period number is the same as the number of energy levels that the electrons of its atoms occupy
  - Period 5 elements have electrons that occupy 5 energy levels

- Elements in the group have the same number of valence electrons for representative elements
  - Look at the Roman numeral


# Checkpoint


- Identify the element that is described by the following information:
  - It is a Group 14 (IVA) metalloid in the third period
  - It is a Group 15 (VA) metalloid in the fifth period
  - It is the other metalloid in Group 15 (VA)
  - It is a halogen that exists in the liquid state at room temperature

- metals (main group)
- metals (transition)
- metals (inner transition)
- metalloids
- nonmetals

MAIN GROUP ELEMENTS | TRANSITION ELEMENTS | MAIN-GROUP ELEMENTS

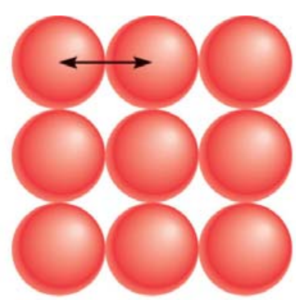
| | 1 (IA) | 2 (IIA) | 3 (IIIB) | 4 (IVB) | 5 (VB) | 6 (VIB) | 7 (VIIB) | 8 | 9 (VIIIB) | 10 | 11 (IB) | 12 (IIB) | 13 (IIIA) | 14 (IVA) | 15 (VA) | 16 (VIA) | 17 (VIIA) | 18 (VIIIA) |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 1 | 1 H 1.01 | | | | | | | | | | | | | | | | | 2 He 4.003 |
| 2 | 3 Li 6.941 | 4 Be 9.012 | | | | | | | | | | | 5 B 10.81 | 6 C 12.01 | 7 N 14.01 | 8 O 16.00 | 9 F 19.00 | 10 Ne 20.18 |
| 3 | 11 Na 22.99 | 12 Mg 24.13 | | | | | | | | | | | 13 Al 26.98 | 14 Si 28.09 | 15 P 30.97 | 16 S 32.07 | 17 Cl 35.45 | 18 Ar 39.95 |
| 4 | 19 K 39.10 | 20 Ca 40.08 | 21 Sc 44.96 | 22 Ti 47.88 | 23 V 50.94 | 24 Cr 52.00 | 25 Mn 54.94 | 26 Fe 55.85 | 27 Co 58.93 | 28 Ni 58.69 | 29 Cu 63.55 | 30 Zn 65.39 | 31 Ga 69.72 | 32 Ge 72.61 | 33 As 74.92 | 34 Se 78.96 | 35 Br 79.90 | 36 Kr 83.80 |
| 5 | 37 Rb 85.47 | 38 Sr 87.62 | 39 Y 88.91 | 40 Zr 91.22 | 41 Nb 92.91 | 42 Mo 95.94 | 43 Tc (98) | 44 Ru 101.1 | 45 Rh 102.9 | 46 Pd 106.4 | 47 Ag 107.9 | 48 Cd 112.4 | 49 In 114.8 | 50 Sn 118.7 | 51 Sb 121.8 | 52 Te 127.6 | 53 I 126.9 | 54 Xe 131.3 |
| 6 | 55 Cs 132.9 | 56 Ba 137.3 | 57 La 138.9 | 72 Hf 178.5 | 73 Ta 180.9 | 74 W 183.9 | 75 Re 186.2 | 76 Os 190.2 | 77 Ir 192.2 | 78 Pt 195.1 | 79 Au 197.0 | 80 Hg 200.6 | 81 Tl 204.4 | 82 Pb 207.2 | 83 Bi 209.0 | 84 Po (209) | 85 At (210) | 86 Rn (222) |
| 7 | 87 Fr (223) | 88 Ra (226) | 89 Ac (227) | 104 Rf (261) | 105 Db (262) | 106 Sg (266) | 107 Bh (262) | 108 Hs (265) | 109 Mt (266) | 110 Uun (269) | 111 Uuu (272) | 112 Uub (277) | | 114 Uuq (285) | | 116 Uuh (289) | | 118 Uuo (293) |

INNER TRANSITION ELEMENTS

| | | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 6 | 58 Ce 140.1 | 59 Pr 140.9 | 60 Nd 144.2 | 61 Pm (145) | 62 Sm 150.4 | 63 Eu 152.0 | 64 Gd 157.3 | 65 Tb 158.9 | 66 Dy 162.5 | 67 Ho 164.9 | 68 Er 167.3 | 69 Tm 168.9 | 70 Yb 173.0 | 71 Lu 175.0 |
| 7 | 90 Th 232.0 | 91 Pa (231) | 92 U 238.0 | 93 Np (237) | 94 Pu (242) | 95 Am (243) | 96 Cm (247) | 97 Bk (247) | 98 Cf (251) | 99 Es (252) | 100 Fm (257) | 101 Md (258) | 102 No (259) | 103 Lr (260) |

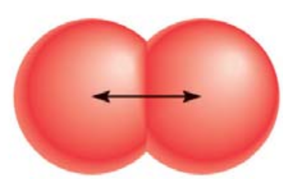
# Atomic Radius

- Atomic radius is one-half the distance between the nuclei of two atoms
- Not all atoms are bound together in the same way
  - Ionic Bonds
  - Covalent Bonds
  - Metals



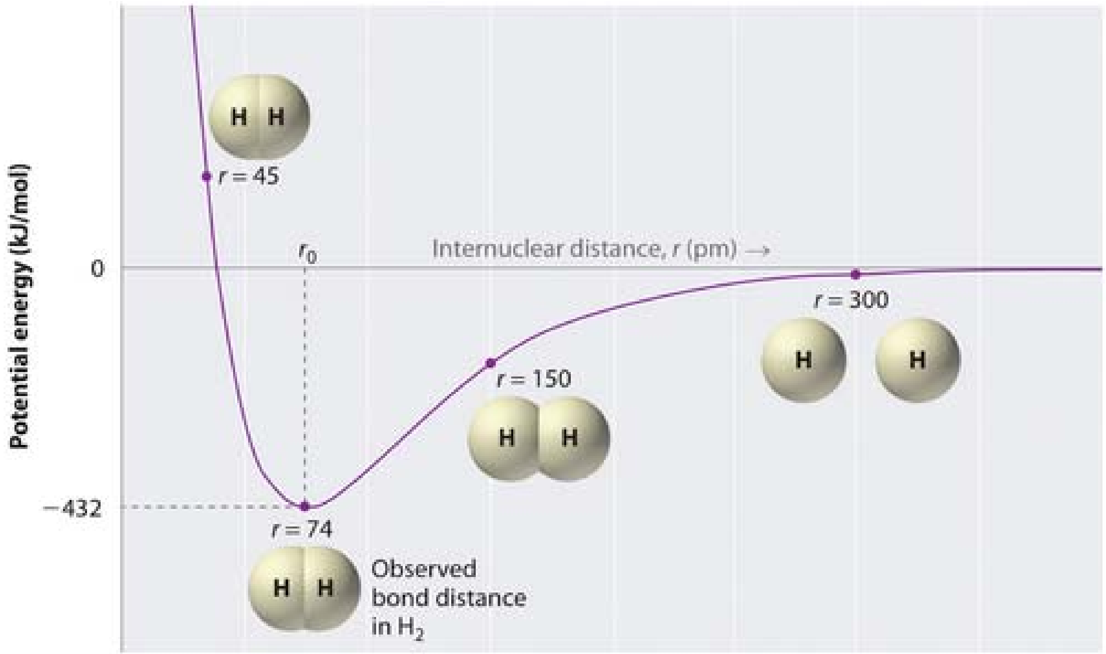
(a)

Atomic radius of metals

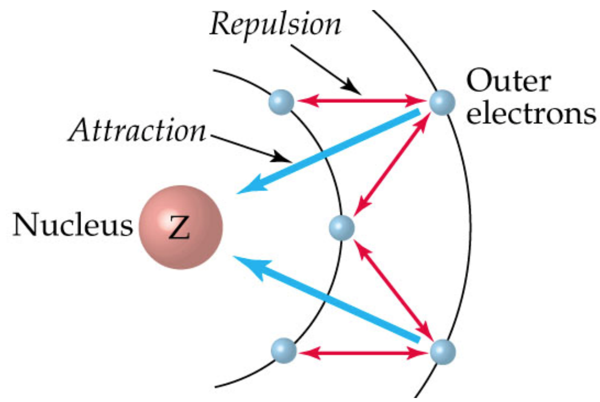


(b)

Atomic radius of diatomic molecules




- The attraction that an electron feels for the nucleus depends on:
  - Number of protons in the nucleus (Z)
  - The distance from the nucleus; the number of energy shells between the nucleus and the electron
  - The amount of screening by core (inner) electrons



# Effective Nuclear Charge ($Z_{eff}$)

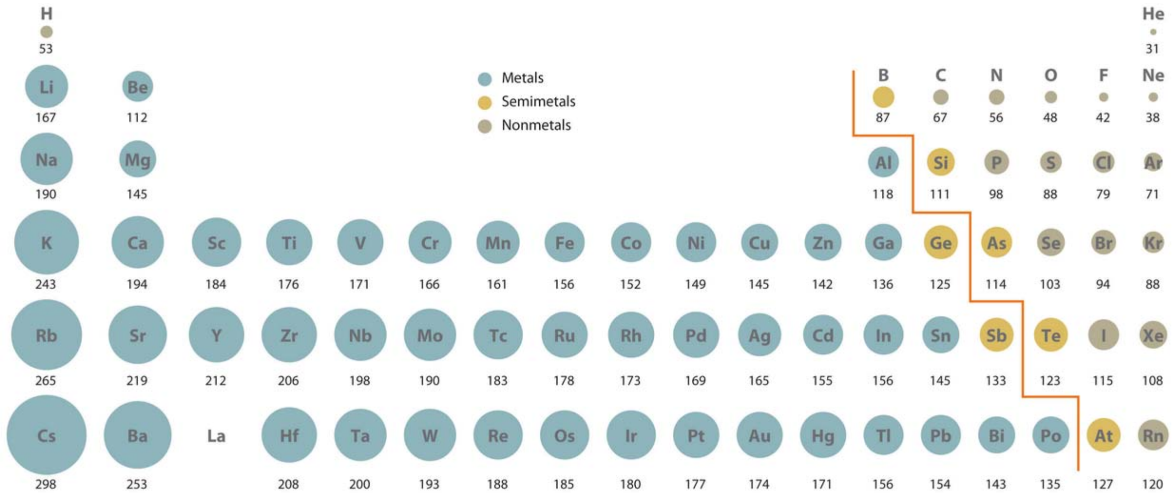
$$Z_{eff} = Z - \sigma$$

**Effective Nuclear Charge** **Actual Nuclear Charge** **Shielding Constant**

- $Z_{eff}$ is the nuclear charge felt by an electron when both the actual nuclear charge (Z) and the repulsive effects (shielding) of the other electrons are taken into account
- Core electrons shield valence electrons more than valence electrons shield one another

## Atomic Radius

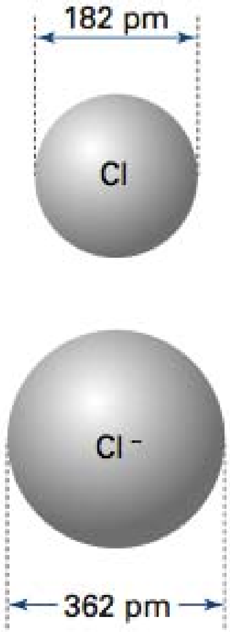
- Down a group: **Increases**
  - More energy levels; valence electrons are farther from the nucleus
- Across a period: **Decreases**
  - The number of protons increases and electrons are restricted to its outer energy level
  - Actual nuclear charge increases (protons increase) while the shielding constant remains the same; Effective Nuclear Charge increases
  - Valence electrons are pulled closer to the nucleus due to increased effective nuclear charge



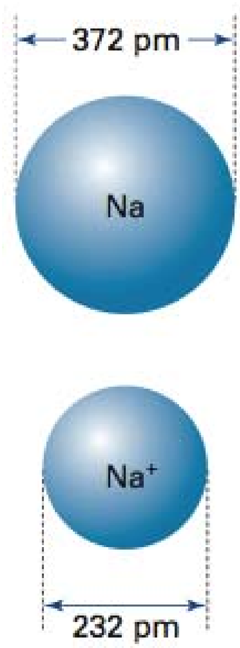


# Checkpoint


- Rank the atoms in each set by decreasing atomic size:
  a) Mg, Be, Ba
  b) Ca, Se, Ga
  c) Br, Rb, Sr



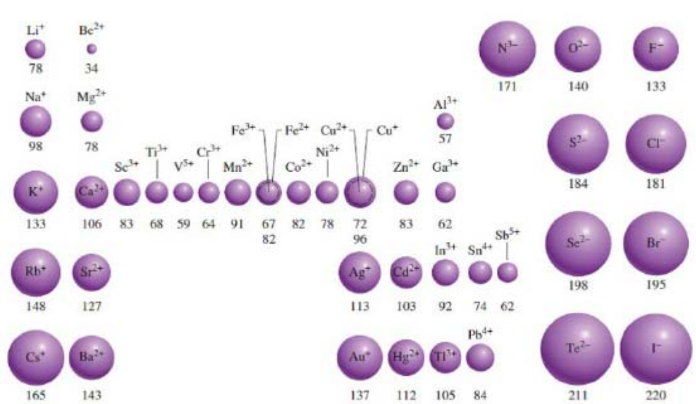

- **Ionic Radius** – radius of a cation or an anion
  – If an atom forms an **anion**, its size increases because the nuclear charge remains the same but repulsion enlarges the electron cloud
  – If an atom forms a **cation**, its size decreases because the nuclear charge remains the same and there is less repulsion




- Down a group: Ionic Radius increases
- Cations are smaller than anions if the ions are isoelectronic

$$+3 < +2 < +1 < -1 < -2, \text{ etc.}$$

- **Isoelectronic** – having the same number of electrons



# Checkpoint

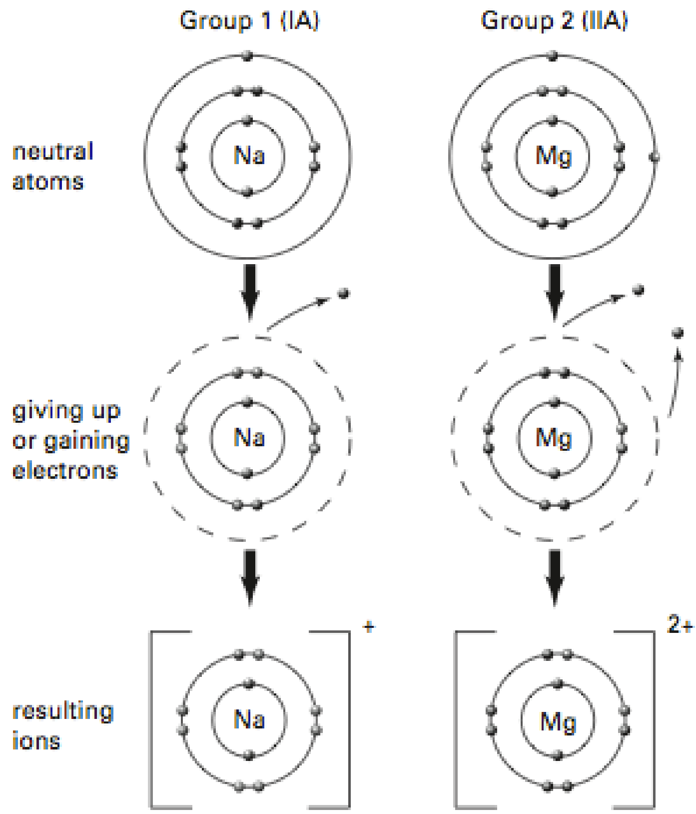
Rank the following in order from smallest to largest ionic radius:

a) $K^+$, $P^{3-}$ , $S^{2-}$, $Cl^-$, $Mg^{2+}$

b) $K^+$ , $Al^{3+}$, $Na^+$, $O^{2-}$, $Mg^{2+}$

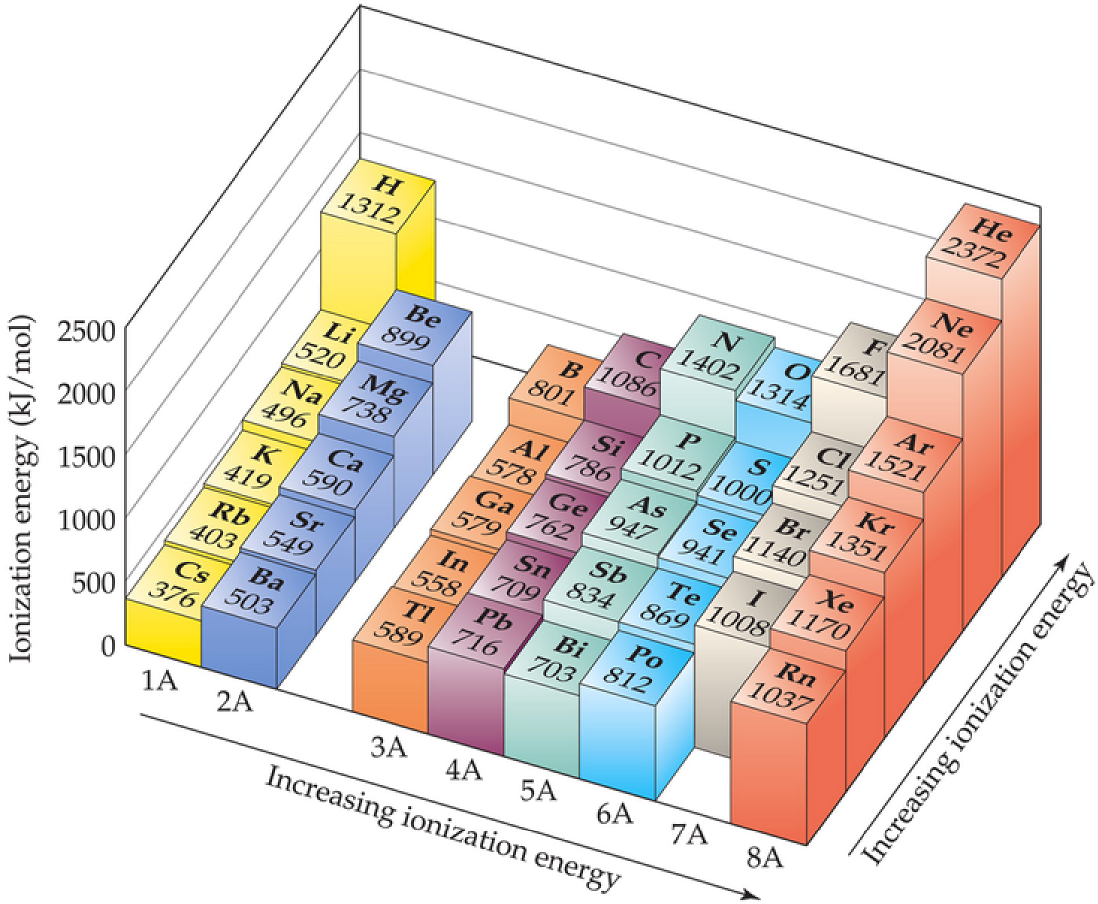
# Ionization Energy

- Minimum energy required to remove an electron from a gaseous atom in its ground state: **$M(g)$ + heat → $M^+(g) + e^-$**
- **First ionization energy ($IE_1$)** – the energy needed to remove the first valence electron
- Measured in kJ/mol
- Low ionization energy – easy to remove
- High ionization energy – hard to remove




- Down a group: **Decreases**
  - Electrons in the valence shell are farther from the nucleus, more electron shielding
- Across a period: **Increases**
  - $Z_{eff}$ increases and the valence electrons are held more tightly

- Noble Gases have the highest ionization energy in each period




# Checkpoint

- Rank the elements in each set by increasing ionization energy:
  a) Xe, He, Ar
  b) Sn, In, Sb
  c) Sr, Ca, Ba

# Electron Affinity

- Ability of an atom to accept one or more electrons
- Measure of the change in energy that occurs when an electron is added to the valence shell to form a negative ion

$$\mathbf{X(g) + e^- \rightarrow X^-(g)}$$

- If energy is released when an atom gains an electron, electron affinity is high (expressed as a negative integer)
- If energy is absorbed when an atom gains an electron, electron affinity is low (expressed as a positive integer)

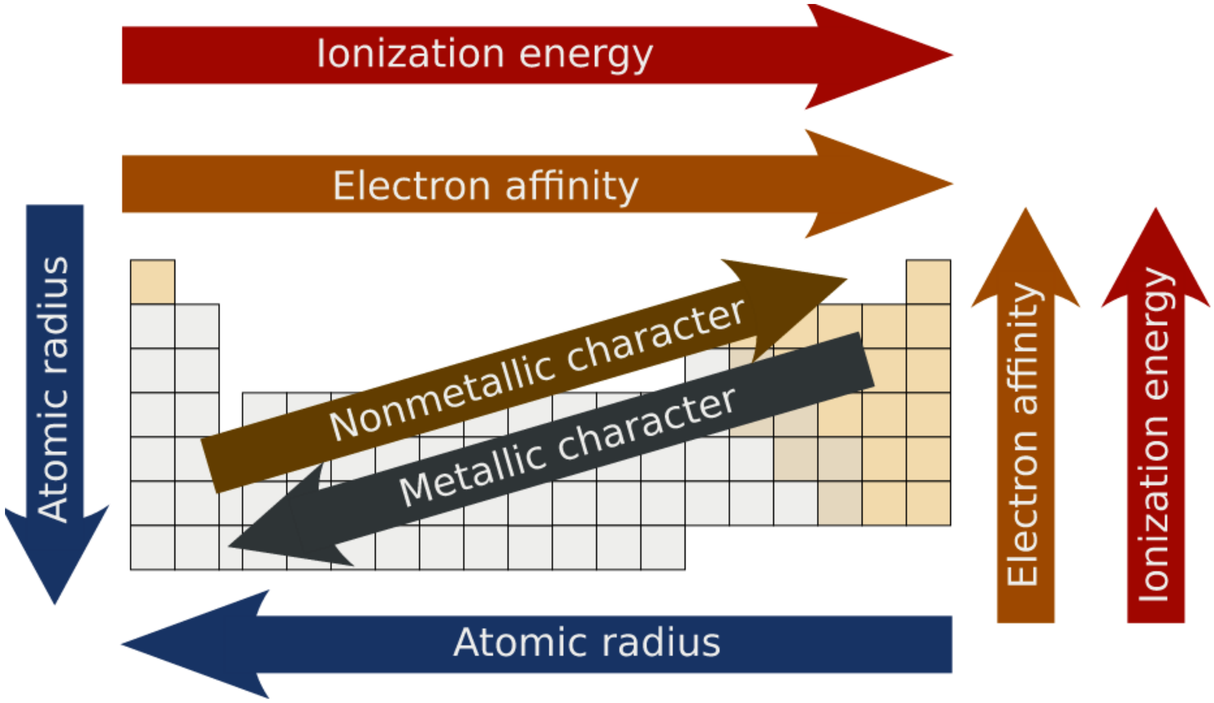
- Down a Group: **Decreases**
- Across a Period: **Increases**

- Halogens have high (negative) electron affinities
- Metals have low (becomes more positive) electron affinities
- Units: kJ/mol

| 1 (IA) | 2 (IIA) | 13 (IIIA) | 14 (IVA) | 15 (VA) | 16 (VIA) | 17 (VIIA) | 18 (VIIIA) |
|---|---|---|---|---|---|---|---|
| **H** −72.8 | | | | | | | **He** (+21) |
| **Li** −59.6 | **Be** (+241) | **B** −26.7 | **C** −122 | **N** 0 | **O** −141 | **F** −328 | **Ne** (+29) |
| **Na** −52.9 | **Mg** (+230) | **Al** −42.5 | **Si** −134 | **P** −72.0 | **S** −200 | **Cl** −349 | **Ar** (+34) |
| **K** −48.4 | **Ca** (+156) | **Ga** −28.9 | **Ge** −119 | **As** −78.2 | **Se** −195 | **Br** −325 | **Kr** (+39) |
| **Rb** −46.9 | **Sr** (+167) | **In** −28.9 | **Sn** −107 | **Sb** −103 | **Te** −190 | **I** −295 | **Xe** (+40) |
| **Cs** −45.5 | **Ba** (+52) | **Tl** −19.3 | **Pb** −35.1 | **Bi** −91.3 | **Po** −183 | **At** −270 | **Rn** (+41) |


# Checkpoint

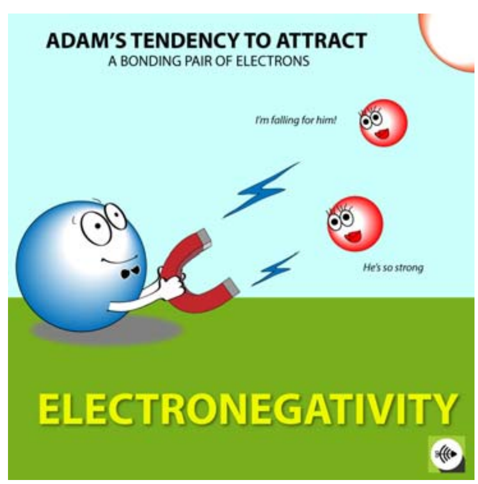

- Which element in the following pairs will have the lower electron affinity?
  a) K or Ca
  b) Cs or F
  c) O or Li

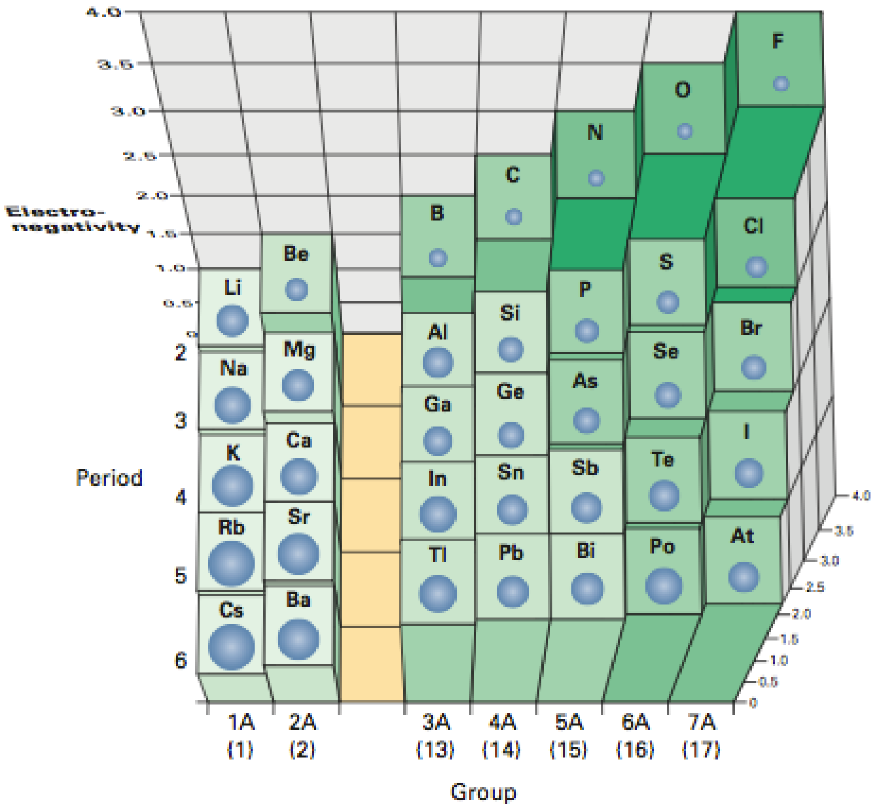



# Electronegativity

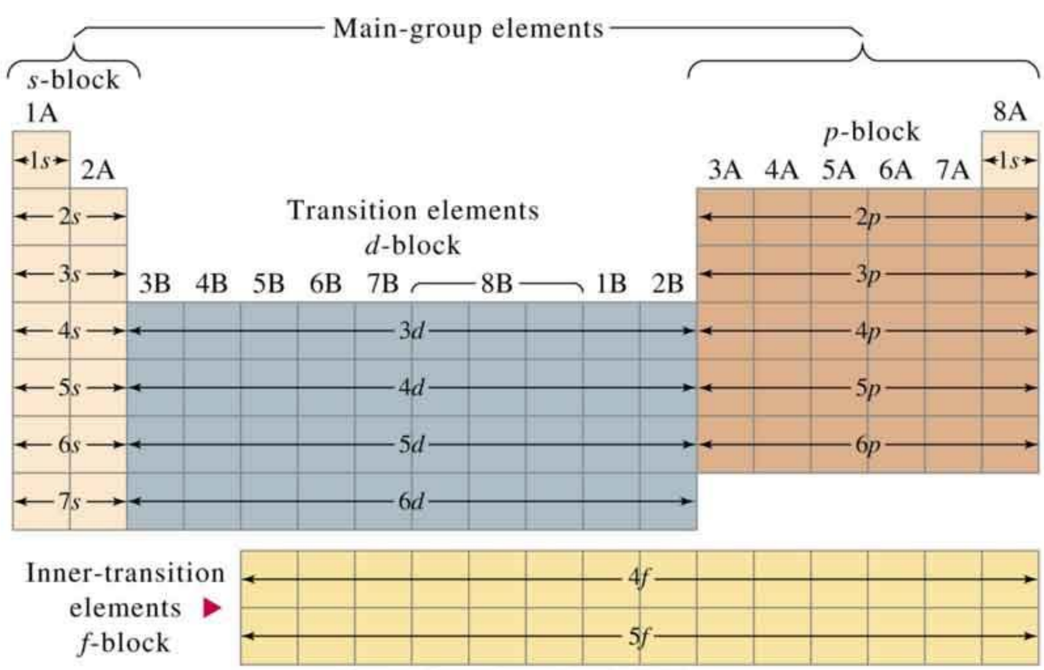
- The ability of an atom to attract electrons to itself; the pull of the negative charge
  - Down a Group: **Decreases**
  - Across a Period: **Increases**

**F>O>Cl>N>Br>I>S>C>H**






# Electron Configurations and the Periodic Table



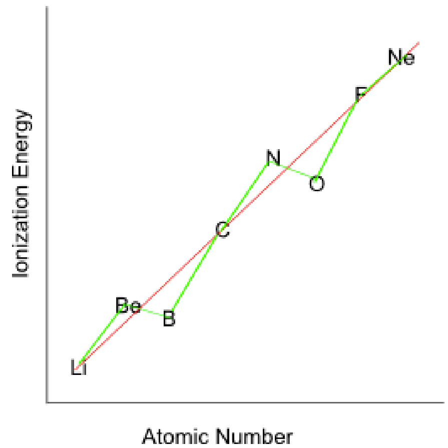
# Orbital Filling Diagrams

1. Draw the order of subshells.
2. How many electrons are in the atom?
3. Write out the electronic configuration in the order of the subshells.
   – s → 2 electrons
   – p → 6 electrons
   – d → 10 electrons
   – f → 14 electrons

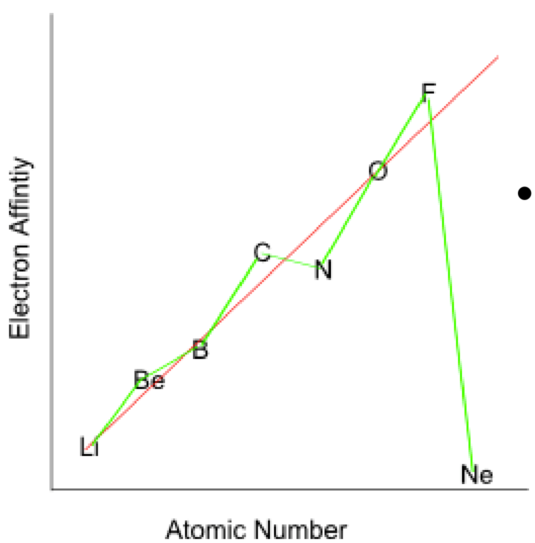
4. Draw boxes or lines to represent each subshell
   – s → 1 box/line
   – p → 3 boxes/lines
   – d → 5 boxes/lines
   – f → 7 boxes/lines
5. Fill in the electrons using up and down arrows, filling one subshell at a time. Single arrows point up by convention.

# Exceptions to Trends

- Reasons:




- Electron configurations due to half-filled stability (ex: oxygen and nitrogen)




- Electron configurations due to half-filled stability (ex: carbon and nitrogen)