Name: Mr. Sudbury - Key Period: _____ Date: _____

Title: **Periodic Trends Packet**	Video: 5.2	Textbook: 5.3 (pp 149+)

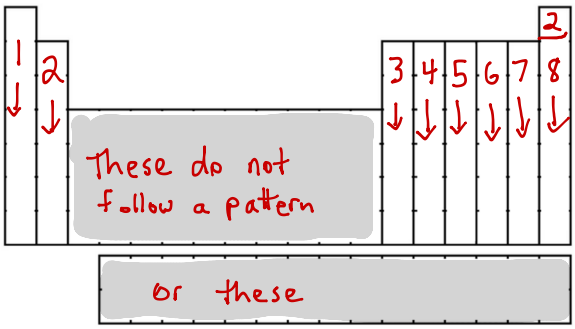
TEKS: 5B-Use the Periodic Table to *identify* and *explain* the properties of chemical families, including alkali metals, alkaline earth metals, halogens, noble gases, and transition metals; **5C**-Use the Periodic Table to *identify* and *explain* periodic trends, including atomic and ionic radii, electronegativity, and ionization energy.

The organization of elements on the periodic table leads to many trends that we can observe. In the following sections you will *define*, *identify*, and *explain* periodic trends based on element location on the periodic table.

Trend # 1: Valence Electrons

1. What is a valence electron?

A valence electron is the outer electron available to be lost, gained or shared in the formation of a chemical compound. (_s^x_p^x electrons)



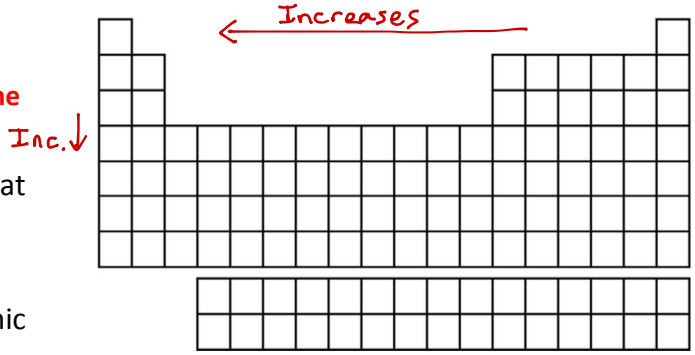
2. Label the periodic table of elements with the number of valence electrons. →
3. Shade the areas of the PT where you cannot predict the number of valence electrons.
4. How many valence electrons do the following elements have? *(You can determine this by the location of the elements.)*

a. Na = 1
b. Cl = 7
c. As = 5
d. Po = 6
e. Rb = 1
f. Cs = 1
g. Al = 3
h. H = 1
i. He = 2
j. In = 3
k. Fr = 1
l. P = 5
m. Xe = 8

Trend # 2: Atomic [Size] Radii

5. Define atomic radii:

The size of an atom. Defined as one-half the distance between the nuclei of identical atoms that are bonded together.



6. Label the PT with vertical and horizontal arrows that represents the ***increasing*** atomic radii of atoms in groups and periods.
7. Circle the element below that has the largest atomic radius: *(You can determine this by the location of the elements.)*

a. (H) or He
b. (Sr) or Sb
c. (Ba) or Mg
d. Cs or (Fr)
e. O or (Te)
f. (Pb) or Si
g. Re or (Ta)
h. (Pa) or Cm
i. (Hg) or Zn
j. Sc or (Y)
k. (Ac) or No
l. (Ga) or Ge
m. (K) or Kr

8. *Explain* why the atomic radii trend increases as it does on the periodic table.

(Left/Right) Across a period-

Increases to the left (decreases to the right) because electrons added to the same energy level experience increasing attraction to the nucleus due to successive addition of protons.

(Up/Down) In a group-

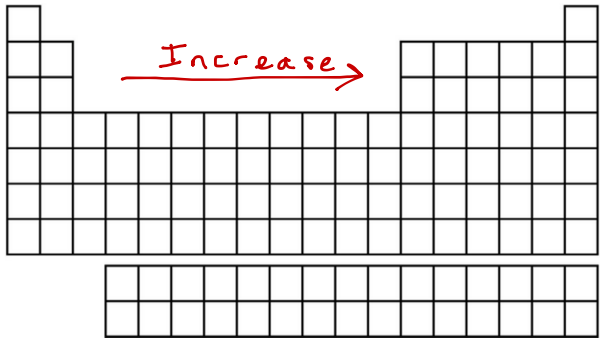
Increases down a group because each period on the table adds a new outer energy level to the electron cloud.

Trend # 3: Ionization Energy

8. Define ionization energy:

Ionization energy is the energy required to remove one electron from a neutral atom of an element. Each atom has multiple ionization energies for removing the first, second, third and so on electrons.

9. Label the PT with vertical and horizontal arrows that represents the ***increasing*** ionization energy in atoms in groups and periods.



10. Circle the element below that has the largest ionization energy: *(Hint: You are circling the element that wants to hold on to their electrons the most…. Keep in mind that some elements want to give one (or more) electrons away and some elements (like the noble gases) don't want to give any away because they have a full octet.)*

a. K or Ca
b. Ba or Cs
c. Ga or Se
d. S or Te
e. F or Br
f. Ne or Ar
g. Fr or Li
h. Ca or As
i. Po or Pb
j. O or Se
k. Nb or Ru
l. Si or Cl
m. He or H

9. *Explain* why the ionization energy trend increases as it does on the periodic table.

(Left/Right) Across a period-

Increases to the right across a period because electrons are harder to move from smaller atoms because they are closer to the nucleus.

(Up/Down) In a group-

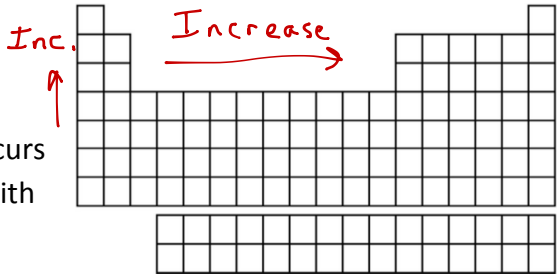
Decreases down a group because the electrons are easier to remove from larger atoms are easier to remove because they are farther from the nucleus.

Trend # 4: Electron Affinity

11. Define electron affinity:

Electron affinity is the energy change that occurs when an electron is acquired by a neutral atom. Electron affinity is related to ionization energy.

12. Electron affinity *is very similar* to ionization energy. Ionization energy is basically how tightly an atom holds on to its "first" or one electron and electron affinity is the energy change that occurs when an electron is acquired by a neutral atom. Label the PT with vertical and horizontal arrows that represents the ***increasing*** electron affinity in groups and periods.

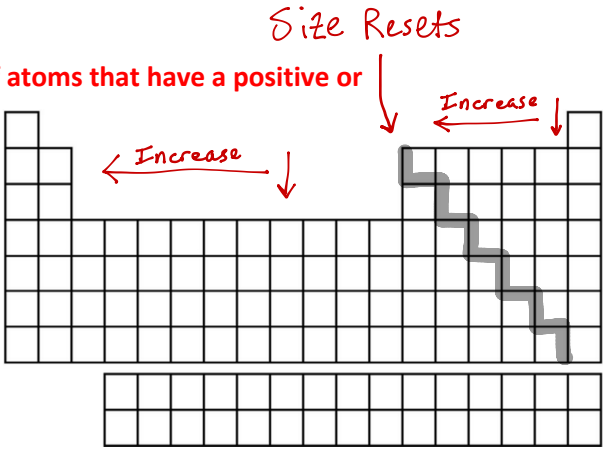


Trend # 5: Ionic Radii

13. Define ionic radii:

Ionic radius is the radius of an ion. An ion is an atom or group of atoms that have a positive or negative charge.

14. Label the PT with vertical and horizontal arrows that represents the ***increasing*** ionic radii in the groups and periods. It is important to recognize that the trend resets when you change from cations (+ ions, the metals) to the anions (- ions, the nonmetals). *All cations are larger than their regular atoms because they have fewer electrons than protons, the remaining electrons are held closer by the nucleus. Anions, on the other hand, are larger than their regular atoms because they have extra electrons and the protons in the nucleus can't hold them as close since the electrons outnumber the protons.*



15. Circle the element below that has the largest ionic radius. *(Hint: You are comparing the ionic radius of one ion to the ionic radius of another ion.)*

a. Na^{+} or Mg^{2+}

b. Li^{+} or Cs^{+}

c. Cs^{+} or Pt^{2+}

d. Ac^{3+} or Ra^{2+}

e. Ag^{+} or Au^{+}

f. Sn^{2+} or Sb^{3+}

g. N^{3-} or P^{3-}

h. F^{-} or I^{-}

i. Cl^{-} or P^{3-}

j. Rb^{+} or Sr^{2+}

k. Be^{2+} or Li^{+}

l. Zn^{2+} or Cd^{2+}

m. O^{2-} or Se^{2-}

16. Why don't the noble gases have an ionic radius listed? Because they don't form ions.

17. *Explain* why the ionic radii trend increases as it does on the periodic table.

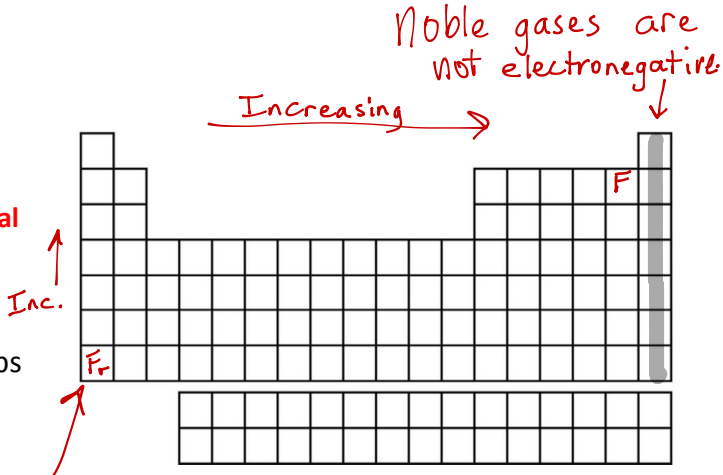
(Left/Right) Across a period-

Anions (nonmetals) are larger than their neutral atoms. The radius of anions increases to the left across the period. Cations are typically smaller than their neutral atoms and the radius of cations also increases to the right.

(Up/Down) In a group-

Anions and cations both increases as you go down a group.

Trend # 6: Electronegativity



18. Define electronegativity:

Electronegativity is a measure of the ability of an atom in a chemical compound to attract electrons.

19. Label the PT with vertical and horizontal arrows that represents the ***increasing*** electronegativity in the groups and periods.

20. What is the most electronegative element? F

21. What is the least electronegative element? Fr

22. Circle the element below that has the highest electronegativity:

a. F or Cl

b. Pd or Ag

c. Y or Sn

d. At or Br

e. Mn or Ge

f. Er or Tm

g. Be or Ra

h. Si or Cl

i. Cd or Sn

j. Pb or C

k. Na or P

l. Mn or Co

23. *Explain* why the electronegativity trend increases as it does on the periodic table.

(Left/Right) Across a period-

Electronegativity increases to the right across a period because shared electrons are closer in the nucleus in smaller atoms.

(Up/Down) In a group-

Electronegativity decreases down a group because shared electrons are farther from the nucleus in larger atoms.

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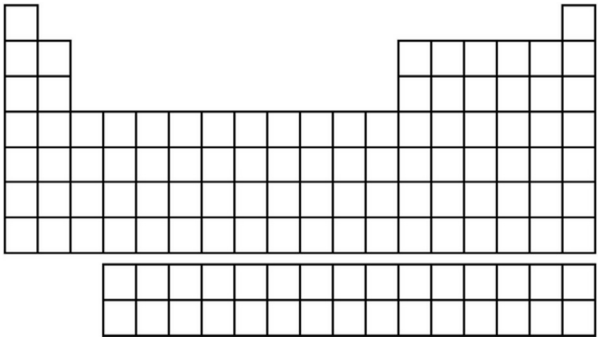

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