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Lesson 3: The Periodic Trend (1.7)

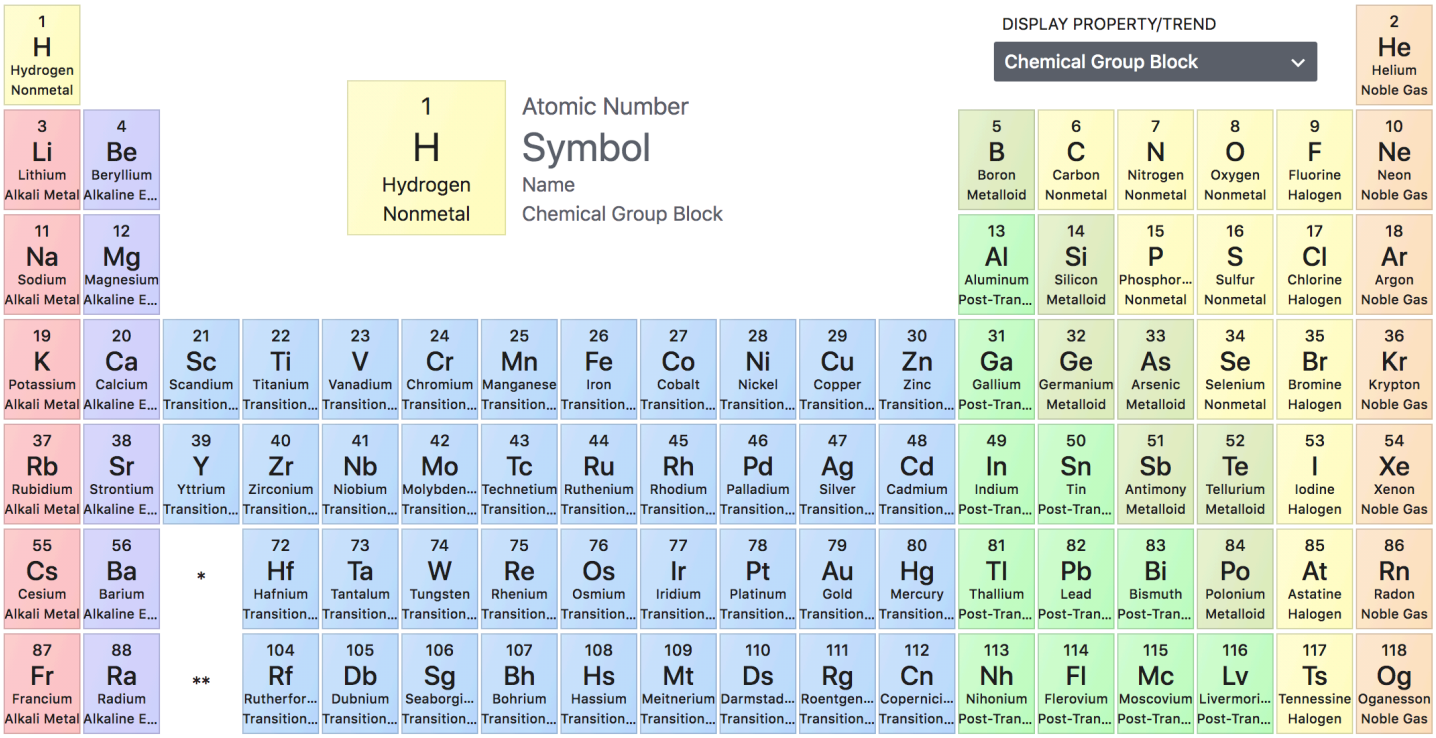
Atomic Radius (AR)

How is atomic radius measure?

The definition of atomic radius: is the distance from the centre of an atom to the boundary within which the electrons spend 90percent of their time.

Generally, as you go down a group in the periodic table, atomic radius _____ because…

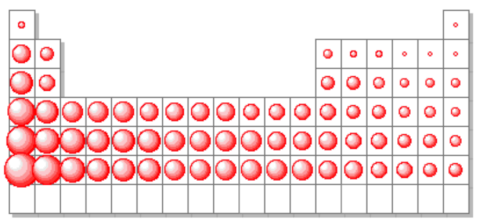
Generally, as you go across a period in the periodic table atomic radius _____ because…



Practice Problems

7. Using only their location in the periodic table, rank the atoms in each set by decreasing atomic size. Explain your answers.

(a) Mg, Be, Ba
(b) Ca, Se, Ga
(c) Br, Rb, Kr
(d) Se, Br, Ca
(e) Ba, Sr, Cs
(f) Se, Br, Cl
(g) Mg, Ca, Li
(h) Sr, Te, Se
(i) In, Br, I
(j) S, Se, O



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Ionic radii for metal cations and non-metal anions

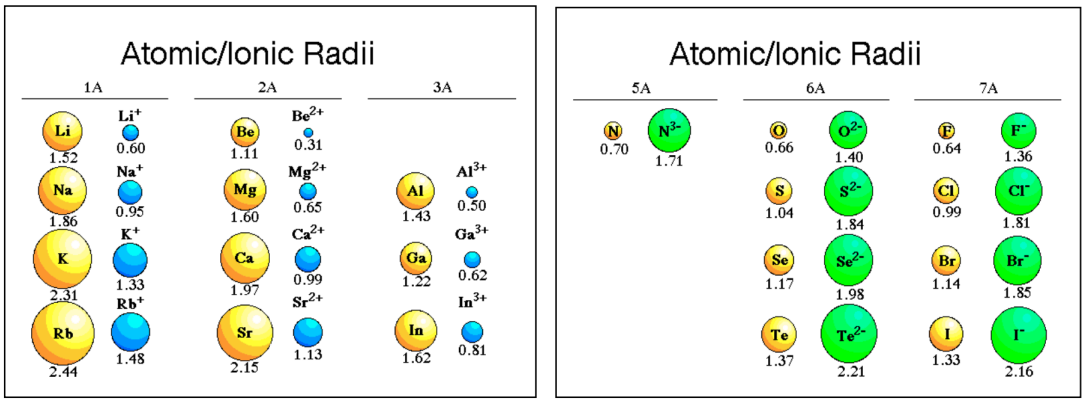

Figure 2: Periodic trend of ionic radii for metal cations and non-metal anions

Ionization Energy (IA)

Define ionization energy: _____.

Generally, as you go down a group in the periodic table, ionization energy _____ because…

Generally, as you go across a period in the periodic table, ionization energy _____ because…

How many ionization energies can an atom have?

First ionization energy: $\mathrm{A(g)} + \mathrm{energy} \rightarrow A^{+}(g) + e^{-}$

Second ionization energy: $A^{+}\mathrm{(g)} + \mathrm{energy} \rightarrow A^{2+}(g) + e^{-}$

Practice Problems

8. Using only a periodic table, rank the elements in each set by increasing ionization energy. Explain your answers.

(a) Xe, He, Ar
(b) Sn, In, Sb
(c) Sr, Ca, Ba
(d) Kr, Br, K
(e) K, Ca, Rb
(f) Kr, Br, Rb

9. Using only a periodic table, identify the atom in each of the following pairs with the *lower* first ionization energy.

(a) B, O
(b) B, In
(c) I, F
(d) F, N
(e) Ca, K
(f) B, Tl

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Electron Affinity (EA)

Define Electron affinity: _____.

Generally, as you go down a group of elements E.A. _____ which means it is _____ to add an electron because…

Generally, as you go across a period of elements E.A. _____ which means it is _____ to add an electron because…

In which corner of the periodic table is it the highest? _____. Explain why this is so.

How can the type (metal/non-metal) of element tell you its E.A. value?

Electronegativity

Pauling Electronegativity Values

H 2.1																He
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.3											Al 1.5	Si 1.9	P 2.2	S 2.6	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 2.8	Kr
Rb 0.8	Sr 0.9	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 1.9	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.5	Xe

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Higher the value, higher the electronegativity, stronger the ability to pull the electrons toward itself when form a covalent bond.

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Section Wrap-up

Despite some irregularities and exceptions, the following periodic trends summarize the relationships among atomic size, ionization energy, and electron affinity:

- Trends for atomic size are the reverse of trends for ionization energy and electron affinity. Larger atoms tend to have lower ionization energies and lower electron affinities.
- Group 16 (VIA) and 17 (VIIA) elements attract electrons strongly. They do not give up electrons readily. In other words, they have a strong tendency to form negative ions. Thus, they have high ionization energies and high electron affinities.
- Group 1 (IA) and 2 (IIA) elements give up electrons readily. They have low or no attraction for electrons. In other words, they have a strong tendency to form positive ions. Thus, they have low ionization energies and low electron affinities.
- Group 18 (VIIIA) elements do not attract electrons and do not give up electrons. In other words, they do not naturally form ions. (They are very stable.) Thus, they have very high ionization energies and very low electron affinities.

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)

Figure 2.18 The units for electron affinity are the same as the units for ionization energy: kJ/mol. High negative numbers mean a high electron affinity. Low negative numbers and any positive numbers mean a low electron affinity.

Practice:

1 K/U How does your understanding of electron arrangement and forces in atoms help you explain the following periodic trends?
(a) atomic radius
(b) ionization energy
(c) electron affinity

2 K/U Using only their location in a periodic table, rank each of the following sets of elements in order of increasing atomic size. Explain your answer in each case.
(a) Mg, S, Cl
(b) Al, B, In
(c) Ne, Ar, Xe
(d) Rb, Xe, Te
(e) P, Na, F
(f) O, S, N

3 K/U Using only their location in a periodic table, rank each of the following sets of elements in order of decreasing ionization energy. Explain your answer in each case.
(a) Cl, Br, I
(b) Ga, Ge, Se
(c) K, Ca, Kr
(d) Na, Li, Cs
(e) S, Cl, Br
(f) Cl, Ar, K

4 K/U Which element in each of the following pairs will have the lower electron affinity? Explain your answer in each case.
(a) K or Ca
(b) O or Li
(c) S or Se
(d) Cs or F

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Overall periodic trend.