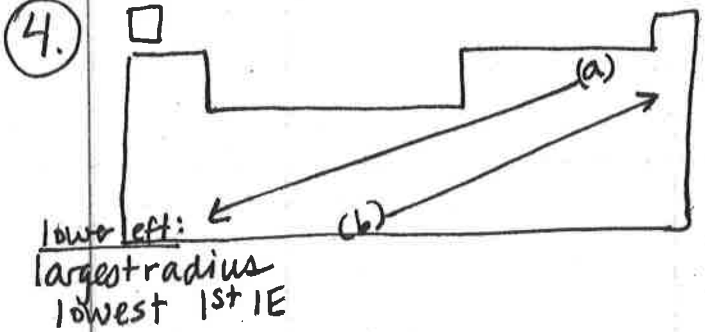
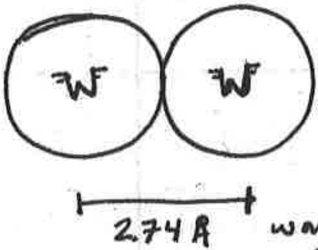
(4.)



smallest radius } upper right of
highest first IE } periodic table

(c) As radius increases, 1st ionization decreases, and vice versa.

(19.) Tungsten: Distance between Tungsten atoms in Tungsten metal is 2.74 Å



(a)

R = the "metallic radius"

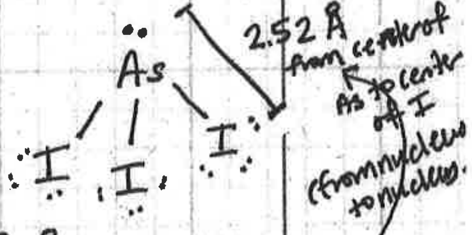
$2R = 2.74$ Å

$R = \frac{2.74 \text{ Å}}{2} = \boxed{1.37 \text{ Å}}$ or $\boxed{1.37 \times 10^{-10} m}$

2.74 Å would be the distance from the center of one Tungsten (W) to the other Tungsten, which is the same as two Tungsten radii (or one diameter...)

(b) Under high pressure, we would expect R to decrease slightly as the atoms move slightly closer together (Density increases slightly).

(21.) Estimate As-I bond in AsI_3, and compare to the experimental bond length of 2.55 Å acc to ~~table~~ figure 7.6 (page 255),



As has a bonding radius (covalent radius) of 1.19 Å

I has a bonding radius (") of 1.33 Å

So the bond length would be ≃ 1.19 Å + 1.33 Å = $\boxed{2.52 \text{ Å}}$ or $\boxed{2.52 \times 10^{-10} m}$

This is very close to the exp'tl value of 2.55 Å ☺

(24) b) Arrange Si, Al, Ge, and Ga in order of increasing atomic radius

13 Al	14 Si
31 Ga	32 Ge

$\boxed{Si < Al < Ge < Ga}$

Si: smallest Radius since upper right on p. table

Al < Ge: Al is further left, and Ge is lower... usually going up/down one has more of an effect than going left/right one, since going up/down involves changing n-level of valence e⁻

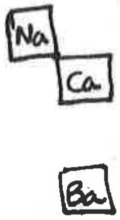
Ga: largest radius since lower left on periodic table

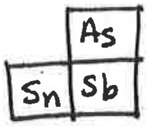
(26.) Arrange each set in order of increasing radius

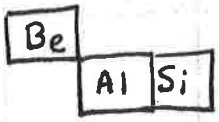
(a) Ba, Ca, Na

(b) Sn, Sb, As

(c) Al, Be, Si



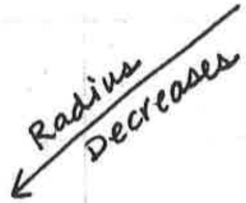




(a) $\boxed{R_{Na} < R_{Ca} < R_{Ba}}$

(b) $\boxed{R_{As} < R_{Sb} < R_{Sn}}$

(c) $\boxed{R_{Be} < R_{Si} < R_{Al}}$



← I used this trend to answer the above Questions, together with the fact that going up / down has more of an effect than going left / right, since going up/down involves an n-level change for the valence electrons. As you can see from figure 7.6 (p. 255) going down one element on the p. table increases the radius as much as going left several elements (at least 4) would increase the radius. I would avoid Questions like (c) on a test, FYI.

(27) True / False :

(a) Cations are larger than their corresponding neutral atoms.

FALSE. cation radius is smaller than the atom radius.

(b) Li^{+1} is smaller than Li. TRUE... (Li^{+1} radius is less than Li radius)

(c) Cl^{-1} is bigger than I^{-1}. FALSE

Atomic Radius increases as you go down a column, because the outer n-level occupied by the valence electrons increases. The same is true for ionic radius, so long as you are comparing ions with the same charge. Iodine is lower on the periodic table than chlorine, and they are both in the halogen column / group 7A.

(31) (a) An isoelectronic series is a set of ions (or it could include an atom) that have the same number of electrons. (iso = same)

(b)

Ga^{+3}	Zr^{+4}	Mn^{+7}	I^{-1}	Pb^{+2}	
31	40	25	53	82	← # protons in the ion
28	36	18	54	80	← # electrons in the ion
Ni	Kr	Ar	Xe	Hg	← Element that the ion is isoelectronic with

(c) ← this problem was cancelled ....

(32) ID at least two ions with the given ground state e- configurations.

(a) [Ar] : $\boxed{P^{-3}, S^{-2}, Cl^{-1}, K^{+1}, Ca^{+2}, Sc^{+3}, Ti^{+4}}$ (pick 2)

(all have 18 electrons)

↑ cheaters! This should be [Ne]$3s^2 3p^6$

(b) [Ar]$3d^5$: $\boxed{Mn^{+2}, Fe^{+3}}$ (I went with two that I knew of ... these are both on your hot pink ion sheet)

(23 electrons)

(c) [Kr]$5s^2 4d^{10}$ $\boxed{In^{+1}, Sn^{+2}, Sb^{+3}}$

(48 electrons)

↑ this can form a +1 ion, though the +3 is more common for Indium.

(36 ab) Consider S, Cl, K and their most common ions.

(a) put atoms in order by increasing size (radius)

S	Cl
K	

$$\boxed{R_{Cl} < R_S < R_K}$$

(b) put ions in order by increasing size

These 3 elements form these ions most commonly: S^{-2}, Cl^{-1}, K^{+1}
These three ions are isoelectronic: all have 18 electrons, like the noble gas Argon. So the ion with the highest nuclear charge (most protons) will have the smallest radius, since the electrons will be more strongly attracted to the nucleus and therefore pulled closer to the nucleus.

$$\boxed{R_{K^{+1}} < R_{Cl^{-1}} < R_{S^{-2}}}$$

19 protons 17 protons 16 protons

(37) Arrange in order of increasing size:

(a) Se^{-2}, Te^{-2}, Se ⟹ $\boxed{R_{Se} < R_{Se^{-2}} < R_{Te}}$

Se
Te

↓ radius increases and anions are larger than the corresponding atom.

(b) $Co^{+3}, Fe^{+2}, Fe^{+3}$ ⟹ $\boxed{R(Co^{+3}) < R(Fe^{+3}) < R(Fe^{+2})}$

these two are isoelectronic (both have 24 electrons) but Cobalt has more protons so will be smaller than Fe^{+2}.

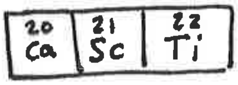
(c) and (d) on next page

Fe^{+3} will be smaller than Fe^{+2} since each electron lost will decrease the radius, since the electrons will experience less repulsion with other e-

Fe^{+3} will be smaller than Co^{+3} since Fe has a higher effective nuclear charge (similar to going right on p. table for elements,

(Cont'd)

(37) (c) Ca, Ti^{+4}, Sc^{+3}



20p 22p 21p
20e 18e 18e

These two are isoelectronic, but since Ti^{+4} has more protons, Ti^{+4} will be smaller / Sc^{+3} would be larger.

Ca^{+2} ion is isoelectronic with Ti^{+4}, Sc^{+3} (all three ions have 18 e-) so, Ca^{+2} would have a larger radius than Ti^{+4}, Sc^{+3} since it has fewer protons (20). And Ca atom should have an even larger radius than Ca^{+2}, since an atom is always larger than its corresponding Cation.

So, $\boxed{R_{Ti^{+4}} < R_{Sc^{+3}} < R_{Ca}}$

(d) Be^{+2}, Na^{+1}, Ne $\Rightarrow \boxed{R_{Be^{+2}} < R_{Na^{+1}} < R_{Ne}}$

$R_{Na^{+1}} < R_{Ne}$ since these two are isoelectronic (both have 10 e-) but Na^{+1} has 1 more proton.

$R_{Be^{+2}} < R_{Na^{+1}}$ because Be^{+2} is isoelectronic with Li^{+1}, so Be^{+2} is smaller than Li^{+1} since Be^{+2} has 1 more proton than Li^{+1}, and Li^{+1} must be smaller than Na^{+1} since Li is 1 element above Na on the periodic table, and we are comparing two ions w/ same charge.

(41.) True or False (Correct/rewrite the false statement(s))

(a) Ionization energies are always negative Quantities: False.

Corrected: Ionization energy [of an element] is always positive. (endothermic)

(b) Oxygen has a larger IE, than Fluorine.

FALSE. Oxygen has a smaller IE, than Fluorine.

Fluorine is to the right of oxygen, so should have a higher IE_1. (There are exceptions to this rule, but this is not one of them)

(c) The second IE of an element is always greater than the first IE of that element. TRUE.

㊺ Which atom will have a ~~larger~~ smaller first ionization energy?

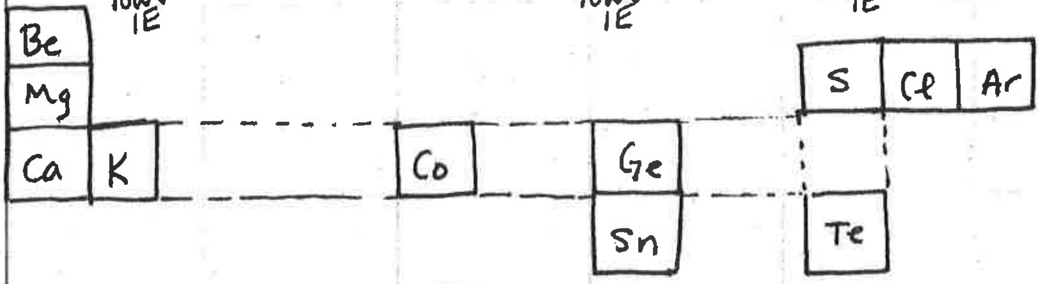
(a) Cl or Ar (b) Be or Ca (c) K or Co (d) S or Ge

(a) Cl ↑ lower IE

(b) Ca ↑ lower IE

(c) K ↑ lower IE

(d) Ge ↑ lower IE



(e) Sn or Te

Sn ↑ lower IE

1st Ionization energy increases as you go up or right on the periodic table.

(As you go up, the n-level decreases, so the radius decreases while the effective nuclear charge stays relatively constant, so the outer e- are more strongly attracted to the nucleus).

(As you go right, the effective (and actual) nuclear charge increases, so e- are more strongly attracted to the nucleus)

㊼ Write e- configurations for these ions:

(a) Fe^{+2}

Fe atom: $[Ar]\ 4s^2 3d^6$

$\boxed{Fe^{+2}\text{ ion: } [Ar]\ 3d^6}$ *

(b) Hg^{+2}

Hg atom: $[Xe]\ 6s^2 4f^{14} 5d^{10}$

$\boxed{Hg^{+2}\text{ ion: } [Xe]\ 4f^{14} 5d^{10}}$ *

(c) Mn^{+2}

Mn atom: $[Ar]\ 4s^2 3d^5$

$\boxed{Mn^{+2}\text{ ion } [Ar]\ 3d^5}$ *

(d) Pt^{+2}

Pt atom: $[Xe]\ 6s^1 4f^{14} 5d^9$ (weird....)

$\boxed{Pt^{+2}\text{ ion: } [Xe]\ 4f^{14} 5d^8}$ *

(e) P^{-3}

P atom: $[Ne]\ 3s^2 3p^3$

$\boxed{P^{-3}\text{ ion: } [Ne]\ 3s^2 3p^6}$

* Transition metals tend to lose their outer s electrons first when they form cations. If they need to lose electrons beyond that to form the ion, they will lose them from their outer d electrons. Inner transition metals (Lanthanides & actinides) are similar, but they'll tend to lose outer s electrons first, then f electrons.

Chapte
Seven

(55.) (a) write equations for the first ionization $\overset{\text{energy}}{\vee}$ of neon, and for the electron affinity of fluorine. Include e- configurations.

First ionization energy of Ne:
this should show a $Ne_{(g)}$ atom losing an electron:

$$\underset{1s^22s^22p^6}{Ne_{(g)}} \longrightarrow \underset{1s^22s^22p^5}{Ne^{+1}_{(g)}} + 1e^-$$

Electron affinity of F:
this should show $F_{(g)}$ gaining an electron.

$$\underset{1s^22s^22p^5}{F_{(g)}} + 1e^- \longrightarrow \underset{1s^22s^22p^6}{F^{-1}_{(g)}}$$

(b) The first ionization energy of Ne will be positive
(ionization is always endothermic for an element;

$$Ne_{(g)} + \underset{(IE_1)}{\text{energy}} \longrightarrow Ne^{+1}_{(g)} + e^-$$

The electron affinity of F will be negative
(electron affinity is negative for most elements)

$$F_{(g)} + 1e^- \longrightarrow F^{-1}_{(g)} + \underset{\uparrow}{\text{energy}}$$

|Electron Affinity|

(c) The magnitudes will not be equal. I would expect IE (Ne) to be greater than |EA| of fluorine.
The change in electron configuration is the same:

$$\ldots 2p^5 \leftrightarrow \ldots 2p^6$$

but Neon has 1 more proton than Fluorine (10 vs 9), so the 2p electrons will be more strongly attracted to the nucleus in Ne than in F/F^-, so more energy will

be absorbed by Ne when it loses the ~~2p~~ 2p6 electron than released by F as it gains the 2p6 electron.

(if you look them up in your book: IE_1 for Ne = 2081 KJ/mole
EA for F = -328 KJ/mole)

(#57) higher Ionization energy corresponds to low metallic character.
lower Ionization energy corresponds to high metallic character.
Metals tend to lose electrons when they react. The lower the IE, the less energy is needed for a metal to lose e-, so the more reactive* the metal will be to form a cation / bond ionically. So the most metallic part of the periodic table is the lower left section, where IE is lowest.

Quantum part 2, (Chapter 7)

(107) Photoelectron spectroscopy (or ultraviolet photoelectron spectroscopy) with mercury vapor, using $\lambda = 58.4$ nm

(a) $E_{photon} = \frac{hc}{\lambda} = \frac{(6.63\times10^{-34} J\cdot s)(3.00\times10^{8} \frac{m}{s})}{(58.4\times10^{-9} m)}$

$E_{photon} = 3.4058\times10^{-18} J$

$(3.4058\times10^{-18} J)\left(\frac{1\ eV}{1.602\times10^{-19} J}\right) = 21.2598\ eV \rightarrow \boxed{21.3\ eV}$

(b) First ionization of Hg

$$\boxed{Hg_{(g)} + IE_1 \longrightarrow Hg^{+1}_{(g)} + e^-}$$

(c) KE of the "ejected" e- is 10.75 eV. Find IE_1 of Hg.

(the KE they give must be the KE of the fastest e- ejected; the e- with the most KE)

$IE = E_{photon} - KE = 21.2598\ eV - 10.75\ eV = 10.5098\ eV$

$IE = (10.5098\ eV)\left(\frac{1.602\times10^{-19} J}{eV}\right)\left(\frac{1\ KJ}{1000\ J}\right)\left(\frac{6.02\times10^{23}}{mole}\right)$

$IE = 1013.57 \rightarrow \boxed{1010\ KJ/mole}$

(d) Which halogen has the most similar IE value?
acc to Figure 9 (page 261), the halogen IE_1 values are:

F : 1681 KJ/mole
Cl : 1251 "
Br : 1140 "
I : 1008 "

so Hg's IE_1 is most similar to that of $\boxed{\text{Iodine}}$