Chapter 3

Atomic Structure

1. Write valence electron configurations for each of the following:

a) carbon	$2s^2\,2p^2$	**b) cobalt**	$4s^2\,3d^7$
c) chlorine	$3s^2\,3p^5$	**d) magnesium**	$3s^2$

3. Which electron experiences the greater nuclear charge? Explain your reasoning in each case.

a) a 5p electron of In or a 5p electron of Sb? Explain.

Effective nuclear charge increases with atomic number within a period. Therefore, a 5**p** electron of Sb experiences the greater nuclear charge

b) a 5s electron of Sn or a 5p electron of Sn?

p orbitals have a nodal plane at the nucleus whereas **s** orbitals do not. Consequently, **p** electrons are shielded by **s** electrons which means that a 5**s** electron will experience the greater nuclear charge.

5. Use ionization energies to explain why +4 ions are very rare.

The ionization energy increases with each electron that is removed because the effective nuclear charge increases with each removed electron. Thus, removing an electron from a +3 ion to produce a +4 ion requires a substantial amount of energy.

7. Use only the Periodic Table to order the elements in each of the following groups by decreasing radius.

a) Na, K, Cl	K > Na > Cl	**b) Al, C, B**	Al > B > C
c) C, Ge, Sn	Sn > Ge > C	**d) Cs, Zn, O**	Cs > Zn > O

9. Order the elements in each group of Exercise 7 by decreasing first ionization energy.

a) Na, K, Cl	Cl > Na < K	**b) Al, C, B**	C > B > Al
c) C, Ge, Sn	C > Ge > Sn	**d) Cs, Zn, O**	O > Zn > Cs

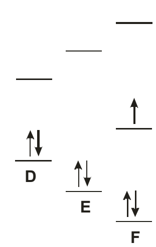
11. Order the elements in each group of Exercise 7 by decreasing electronegativity.

a) Na, K, Cl	Cl > Na > K	**b) Al, C, B**	Al < B < C
c) C, Ge, Sn	C > Ge > Sn	**d) Cs, Zn, O**	O > Zn > Cs

13. Which is the most electronegative atom in each of the following groups?

a) N, O, P O **b) O, S, Se** O **c) Si, P, S** S

15. Consider the diagram for atoms D, E, and F.



a) List the atoms in order of increasing ionization energy.

The reverse order of highst electron energy of: F < D < E

b) List the atoms in order of decreasing electronegativity.

The reverse order of lowest unfilled orbital energy: F > D > E

c) Identify each atom as either paramagnetic or diamagnetic.

Only F is paramagnetic.

17. The relative energies of the highest occupied orbitals of H, Li, and F are shown to the right.

a) Identify each as X, Y, or Z

Fluorine is most electronegative, so it has the lowest energy orbital. Lithium has the lowest ionization energy, so it has the highest energy orbital. X = F; Y = H; Z = Li

b) Where would the energies of the 3p orbital of Cl and the 4s orbital of K be placed? (That is below X, between X and Y, between Y and Z, or above Z)

Chlorine is less electronegative than fluorine but more electronegative than H. Chlorine's 3p orbital is between X and Y. Potassium has the lowest ionization energy, so its valence orbital above Li (Z).

19. Use Equation 3.2 to explain why first ionization energies decrease going down a group and why electronegativities increase going from left to right in a period.

Ionization energies decrease as E_n becomes less negative for occupied orbitals, so they decrease going down a group because n gets larger, which increases the denominator and makes E_n less negative. Electronegativities increase as E_n becomes more negative for unfilled orbitals, so they increase going across a period because Z_{eff} increases.

21. Use the following effective nuclear charges experienced by the outermost electrons to order their outermost orbitals from lowest to highest energy:

Element	**P**	**Al**	**Br**	**Pb**
$\mathbf{Z_{eff}}$	**4.89**	**4.07**	**9.03**	**12.39**

Use 3.2, $E_n \propto -\frac{Z_{eff}^2}{n^2}$ to get relative orbital energies. See above table for values of Z_{eff}. P and Al are each in the third period, so n = 3 for them. Br is in the fourth period, so n = 4. n = 6 for Pb.

Element	P	Al	Br	Pb
n	3	3	4	6
Z_{eff}	4.89	4.07	9.03	12.39
$-Z_{eff}^2/n^2$	-2.66	-1.84	-5.10	-4.26

Lowest energy orbital = Br < Pb < P < Al

The valence orbitals on Pb are predicted to be a lower energy than those on P, which is consistent with the fact that Pb is more electronegative than P.

23. Use Equation 3.2 to determine where in Figure 3.4 the energy of the valence orbitals of nitrogen (Z_{eff} = 3.9) would be found? Is the location consistent with its electronegativity of 3.04? … with its ionization energy of 1402 kJ/mol? If not, explain why.

$E_N \propto -\frac{Z_{eff}^2}{n^2} = -\frac{3.9^2}{2^2} = -3.8$, which places it between sulfur and chlorine in Figure 3.2. The electronegativities of S and Cl from the table on the last page of the text are 2.6 and 3.2. So the electronegativity of N is indeed between these values. The ionization energies of S and Cl from Figure 3.5 are about 1000 and 1300 kJ/mol, so the ionization energy of N is high. This is expected because the electron in nitrogen is removed from a half-filled valence shell.