

Solutions to Suggested Problems From Chapter 3

6. The frequency is 1 s^{-1} because one boxcar passes every second. We know that speed is the product of wavelength and frequency; thus

$$c = \lambda \times \nu = 45 \text{ ft} \times 1 \text{ s}^{-1} = 45 \text{ ft s}^{-1}$$

The speed in mph is

$$\frac{45 \text{ ft}}{\text{s}} \times \frac{60 \text{ s}}{\text{min}} \times \frac{60 \text{ min}}{\text{hr}} \times \frac{1 \text{ mile}}{5280 \text{ ft}} = 31 \text{ mph}$$

7. Because the product of λ and ν is a constant, a doubling of the frequency must cut the wavelength in half.
14. Given that there are 1,650,763.73 wavelengths per meter, the reciprocal of this gives the wavelength of the orange-red line in the ^{86}Kr spectrum as

$$\lambda = \frac{1 \text{ m}}{1,650,763.73} = 6.06 \times 10^{-7} \text{ m} = 606 \text{ nm}$$

Now that we know the wavelength, we can calculate the frequency using the equation

$$c = \lambda \times \nu$$

Thus

$$2.998 \times 10^8 \text{ m/s} = (6.06 \times 10^{-7} \text{ m})\nu$$

$$\nu = 4.95 \times 10^{14} \text{ s}^{-1}$$

The wavelength of 606 nm falls in the visible region of the electromagnetic spectrum (as if identifying the radiation as orange-red didn't give that away!).

15. Since we know the wavelength, we can calculate the frequency using the equation

$$c = \lambda \times \nu$$

Thus

$$2.998 \times 10^8 \text{ m/s} = (6 \times 10^{-9} \text{ m})\nu$$

$$\nu = 5 \times 10^{16} \text{ s}^{-1}$$

18. Since we know the wavelength, we can calculate the frequency using the equation

$$c = \lambda \times \nu$$

Thus

$$2.998 \times 10^8 \text{ m/s} = (588.9953 \times 10^{-9} \text{ m})\nu$$

$$\nu = 5.090 \times 10^{14} \text{ s}^{-1}$$

and

$$2.998 \times 10^8 \text{ m/s} = (589.5293 \times 10^{-9} \text{ m})\nu$$

$$\nu = 5.085 \times 10^{14} \text{ s}^{-1}$$

25. Here we need the energy for a single Cl_2 molecule, but we have been given the energy for a mole of Cl_2 molecules. To convert we use Avogadro's number

$$\frac{243.4 \text{ kJ}}{\text{mole}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23}} \times \frac{1000 \text{ J}}{\text{kJ}} = 4.042 \times 10^{-19} \text{ J}$$

Knowing the energy of the photon, we can now calculate the wavelength using the equation

$$E = \frac{hc}{\lambda}$$

Thus

$$4.042 \times 10^{-19} \text{ J} = \frac{(6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s})}{\lambda}$$

$$\lambda = 4.915 \times 10^{-7} \text{ m or } 491.5 \text{ nm}$$

This wavelength corresponds to visible light.

32. If the two particles are separated by an infinite distance then the denominator becomes infinitely large and the force approaches zero.

If the signs are identical, then the force between the particles will be positive.

If the signs are not identical, then the force must be negative.

Because a proton has $q = +1$ and an electron has $q = -1$, the force must be negative.

Because force is inversely proportional to distance the electron that is 500 pm from a +1 charge experiences the greater force.

Because force is directly proportional to charge, the electron that is 1000 pm from the +2 charge will experience the greater force.

34. The second one; attraction requires opposite charges and +2/-2 will have a greater attractive force than +1/-1

The first and third are equal (at $-1/r^2$) and greater than the second (which is $-1/4r^2$)

The middle has the smallest force, so it requires the least amount of energy to remove the negatively charged particle.

36. First, convert the energy to a wavelength; thus

$$E = \frac{hc}{\lambda}$$

$$2.178 \times 10^{-18} \text{ J} = \frac{(6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s})}{\lambda}$$

$$\lambda = 9.121 \times 10^{-8} \text{ m or } 91.21 \text{ nm}$$

Next, use the empirical equation relating the wavelength of an emission line to values of n_1 and n_2 ; thus

$$\frac{1}{91.21 \text{ nm}} = 1.09737 \times 10^{-2} \text{ nm} \left(\frac{1}{1^2} - \frac{1}{n_2^2} \right)$$

$$0.9988 = \left(1 - \frac{1}{n_2^2} \right)$$

$$1.248 \times 10^{-3} = \frac{1}{n_2^2}$$

$$n_2 = 28$$

Note – you may end up with a different answer, but don't worry as long as your approach is okay. The difficulty here is that this energy is nearly sufficient to remove

the electron from the atom. I get $n_2 = 34$ when working through the math without rounding off any intermediate answer.

37. Use the empirical equation relating the wavelength of an emission line to values of n_1 and n_2 . Thus, for $n_1 = 2$ and $n_2 = 3$ we have

$$\frac{1}{\lambda} = 1.09737 \times 10^{-2} \text{ nm} \left(\frac{1}{4} - \frac{1}{9} \right)$$

Solving for the wavelength gives

$$\lambda = 656.114 \text{ nm}$$

The energy, therefore, is

$$E = \frac{hc}{\lambda}$$

$$E = \frac{(6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s})}{656.114 \times 10^{-9} \text{ m}} = 3.028 \times 10^{-19} \text{ J}$$

For $n_1 = 2$ and $n_2 = 3$ we have

$$\frac{1}{\lambda} = 1.09737 \times 10^{-2} \text{ nm} \left(\frac{1}{4} - \frac{1}{16} \right)$$

Solving for the wavelength gives

$$\lambda = 486.010 \text{ nm}$$

The energy, therefore, is

$$E = \frac{hc}{\lambda}$$

$$E = \frac{(6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s})}{486.010 \times 10^{-9} \text{ m}} = 4.087 \times 10^{-19} \text{ J}$$

For $n_1 = 2$ and $n_2 = \infty$ we note that $1/\infty$ is essentially zero; thus

$$\frac{1}{\lambda} = 1.09737 \times 10^{-2} \text{ nm} \left(\frac{1}{4} - 0 \right)$$

Solving for the wavelength gives

$$\lambda = 364.508 \text{ nm}$$

The energy, therefore, is

$$E = \frac{hc}{\lambda}$$

$$E = \frac{(6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s})}{364.508 \times 10^{-9} \text{ m}} = 5.450 \times 10^{-19} \text{ J}$$

38. Use the empirical equation relating the wavelength of an emission line to values of n_1 and n_2 . Thus, for $n_1 = 2$ and $n_2 = 3$ we have

$$\frac{1}{\lambda} = 1.09737 \times 10^{-2} \text{ nm} \left(\frac{1}{4} - \frac{1}{9} \right)$$

Solving for the wavelength gives

$$\lambda = 656.114 \text{ nm}$$

The energy, therefore, is

$$E = \frac{hc}{\lambda}$$

$$E = \frac{(6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s})}{656.114 \times 10^{-9} \text{ m}} = 3.028 \times 10^{-19} \text{ J}$$

Note that this answer is the same as in problem 37a. In that case, however, we were finding the energy that must be absorbed to yield the transition from $n_1 = 2$ to $n_2 = 3$. The energy is the same for absorption and emission as required by the conservation of energy.

39. Removing an electron from an atom is the same as moving it to a point where n_2 is ∞ . It will be easier to do this from $n_1 = 2$ than for $n_1 = 1$. You can convince yourself of this by looking at the following equation

$$\frac{1}{\lambda} = 1.09737 \times 10^{-2} \text{ nm} \left(\frac{1}{n_1^2} - 0 \right)$$

The larger the value of n_1 , the smaller is the term in parentheses and the larger the wavelength. Because energy is inversely proportional to energy

$$E = \frac{hc}{\lambda}$$

a longer wavelength corresponds to a smaller energy. Obviously, the electron is most stable at $n_1 = 1$ since it requires more energy to remove it.

40. The wavelength corresponding to this energy is

$$E = \frac{hc}{\lambda}$$

$$2.091 \times 10^{-18} \text{ J} = \frac{(6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s})}{\lambda}$$

which gives

$$\lambda = 9.500 \times 10^{-8} \text{ m or } 95.00 \text{ nm}$$

The value of n_2 , therefore, is

$$\frac{1}{95.00 \text{ nm}} = 1.09737 \times 10^{-2} \text{ nm} \left(\frac{1}{1^2} - \frac{1}{n_2^2} \right)$$

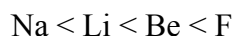
which gives

$$n_2 = 5$$

42. Looking at Table 3.4 and Figure 3.7 we see that IE_1 for F is 1681 kJ/mol and the IE_1 for Cl is 1251 kJ/mol. The IE_1 for Br will be smaller than that for Cl, but the change from Cl \rightarrow Br probably isn't as big as that for F \rightarrow Cl. A good guess, therefore, might be ~ 1000 -1100 kJ/mol.

43. It takes energy because of the coulombic attraction between the electron and the atom's nucleus.

44. If you examine the data in Table 3.4 and Figure 3.7 you will see that ionization energies *generally* increase across a row (period) and decrease down a group (column). Thus, the predicted order is



45. A table will suffice here

Element	Electrons in			
	First Shell	Second Shell	Third Shell	Fourth Shell
Na	2	8	1	0
Mg	2	8	2	0
Al	2	8	3	0
Si	2	8	4	0
P	2	8	5	0
S	2	8	6	0
Cl	2	8	7	0
Ar	2	8	8	0
K	2	8	8	1
Ca	2	8	8	2

46. Recall that

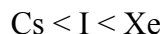
$$\text{core charge} = \text{charge on nucleus} - \text{number of core electrons}$$

Thus

Element	Charge on Nucleus	Number Core e ⁻	Core Charge
Na	11	10	1
Mg	12	10	2
Al	13	10	3
Si	14	10	4
P	15	10	5
S	16	10	6
Cl	17	10	7
Ar	18	10	8
K	19	18	1
Ca	20	18	2

49. The ionization energy for Ar will be greater than that for Cl⁻ because the core charge for Ar (+8) is greater than that for Cl⁻ (+7).

50. In general, ionization energies increase across a row and decrease down a column; thus we predict



You can verify this in Table B.5 in Appendix B.

51. The hardest electron to lose is always the one that is closest to the nucleus; that is, the one(s) in the innermost shell.

62. Ionization energies increase across a period because the charge on the nucleus increases.
63. Ionization energies decrease down a group because the distance between the electron and nucleus increases.
64. Nitrogen will have the largest first ionization energy because it is furthest to the right on the periodic table (larger charge on the nucleus) and nearer to the top of periodic table (closer to the nucleus) when compared to the other elements in the list.
65. Calcium will have the smallest first ionization energy because it is furthest to the left on the periodic table (smaller charge on the nucleus) and nearer to the bottom of the periodic table (further from the nucleus) when compared to the other elements in the list.
66. The number of valence electrons for each element is shown here

Element	Number Valence e ⁻
H	1
He	2
Li	1
Be	2
B	3
C	4
N	5
O	6
F	7
Ne	8

67. For the first three rows of the periodic table, the core charge, the number of valence electrons, and the group number (using the roman numeral and A/B format) are all the same. Things become less consistent beyond the third row, which is a weakness of this model.
68. Valence electrons are easier to remove because they are further from the nucleus and because they experience a core charge that is less than the charge on the nucleus.
69. No. The PES data show that each shell has one or more subshells. When there is more than one subshell, then the energy of an electron depends upon which subshell it is in.
70. Yes. This is just enough energy to remove the outermost electron; that is, the photon's energy just matches the first ionization energy.

72. Each PES experiment (that is, each incident in which a photon strikes an atom) results in the emission of a single electron.

73. Yes. Every electron has an equal probability of being ejected.

74. The key equation is

$$E_{\text{photon}} = \text{IE} + \text{KE}$$

$$920.6 \text{ kJ/mol} = \text{IE} + 120.0 \text{ kJ/mol}$$

$$\text{IE} = 700.6 \text{ kJ/mol}$$

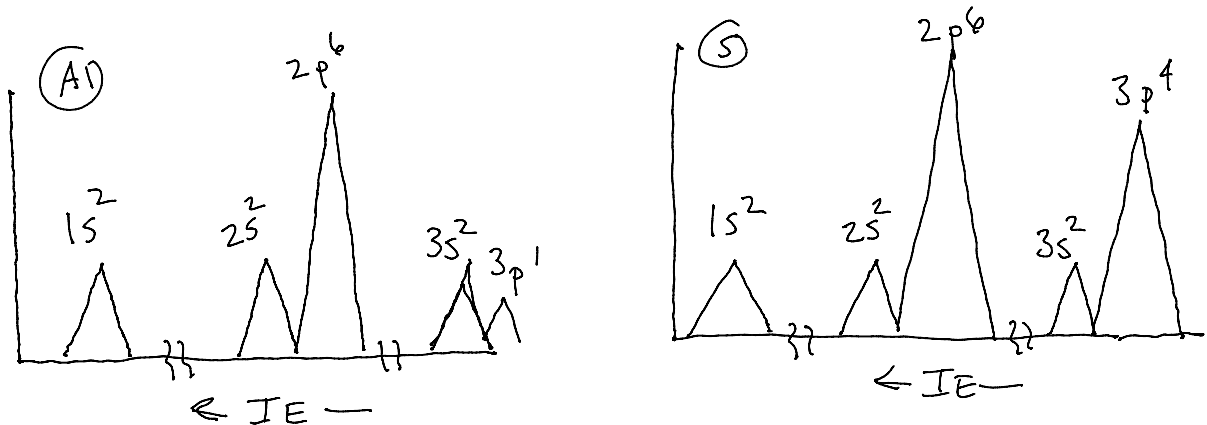
76. Hydrogen and helium each have only one peak in their PES spectra because each has only 1s electrons.

77. Lithium has electrons in both the first and second shells. The first shell has two electrons ($1s^2$) and the second shell has one electron ($2s^1$). This difference in the number of electrons explains the difference in the heights of the two peaks.

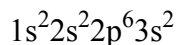
78. The principal reason for this is that B has a nucleus with a charge of +5, while the nucleus for H has a charge of +1. The larger the charge on the nucleus, the greater the ionization energy. Note that for the innermost electron (the one with the largest IE, the charge on the nucleus is the same as the core charge).

82. Two peaks are assigned to the $n = 2$ shell rather than to the $n = 1$ shell because the first shell has only one subshell and each peak represents a subshell. The rationale for assigning the peak at 84.0 MJ/mol to the first shell is that this is the larger of the two ionization energies, which is consistent with electrons that lie closer to the nucleus.

83. The PES spectra are shown below. The squiggly lines on the x-axis indicate a change of scale. Each peak is labeled showing the shell's number, the subshell's letter, and the number of electrons in the subshell.



84. The presence of four peaks suggests that we have the 1s, 2s, 2p, and 3s subshells. The relative heights suggests an electron configuration of



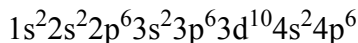
or magnesium, Mg. Note that we know that the first subshell (1s) has two electrons as we can't populate the next subshell without filling the first one. This establishes the relative intensity for two electrons in a subshell.

85. We expect the ionization energies for the subshells to be slightly larger for K than they are for Ar due to the higher nuclear charge; thus

1s ~340 MJ/mol
2s ~40 MJ/mol
2p ~30 MJ/mol
3s ~4 MJ/mol
3p ~2.5 MJ/mol

In addition, K will have a 4s electron which should be similar to that for a 3s electron in Na or a 2s electron in Li, but slightly smaller; a guess of ~0.45 MJ/mol is reasonable.

86. The electron configuration appears to be



which is krypton, Kr.

87. Note that there is a typo in the figure. The first peak must have an ionization energy of 151 instead of 15.1 (otherwise there would be no need to show a change in the scale for the x-axis).

(a) We see evidence for the following subshells: 1s, 2s, 2p, 3s, 3p given the five peaks. There appear to be a total of three electrons in the 3s and 3p subshells meaning that there are three valence electrons which is consistent with group 3.

(b) The maximum positive charge occurs when we remove all valence electrons; thus it is +3.

(c) The core charge is

$$\text{core charge} = \text{charge on nucleus} - \text{core electrons}$$

The charge on the nucleus, since this atom is neutral, must equal the total number of electrons, which is 13 ($1s^2 2s^2 2p^6 3s^2 3p^1$); thus

$$\text{core charge} = 13 - 10 = +3$$

(d) The element is aluminum, Al. This choice is consistent with the group number and with the total number of electrons.

(e) To determine this you must start by converting the wavelength of light to an energy; thus

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s})}{206 \times 10^{-9} \text{ m}} = 9.64 \times 10^{-19} \text{ J}$$

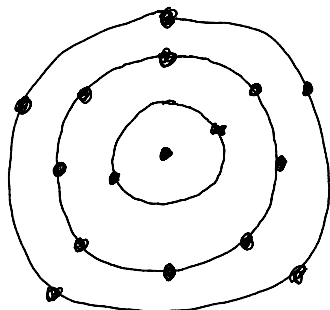
Next, we need to convert this energy for a single photon to an energy per mole; thus

$$9.64 \times 10^{-19} \text{ J} \times \frac{1 \text{ MJ}}{1 \times 10^6 \text{ J}} \times \frac{6.022 \times 10^{23}}{\text{mol}} = 0.581 \text{ MJ/mol}$$

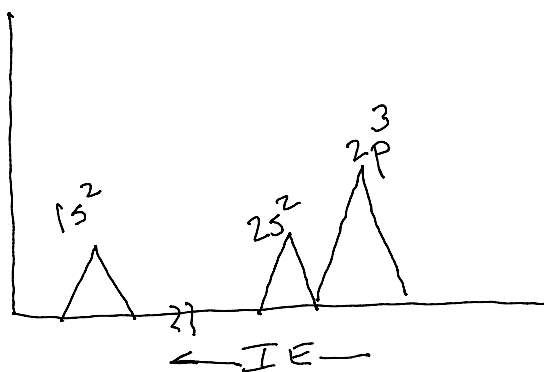
The smallest ionization energy in the PES spectrum is 0.58 MJ/mol so this light has just enough energy to remove the 3p electron.

88. (a) We see evidence for the following subshells: 1s, 2s, 2p, 3s, 3p given the presence of five peaks in the PES spectrum. There appear to be a total of five electrons in the 3s and 3p subshells, which corresponds to phosphorous, P.

(b)



(c)



(d) To determine this you must start by converting the wavelength of light to an energy; thus

$$E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ Js})(2.998 \times 10^8 \text{ m/s})}{1.20 \times 10^{-8} \text{ m}} = 1.665 \times 10^{-17} \text{ J}$$

Next, we need to convert this energy for a single photon to an energy per mole; thus

$$1.665 \times 10^{-17} \text{ J} \times \frac{1 \text{ MJ}}{1 \times 10^6 \text{ J}} \times \frac{6.022 \times 10^{23}}{\text{mol}} = 9.97 \text{ MJ/mol}$$

This energy is sufficient to eject either a 3p electron (IE = 1.01 MJ/mol) or a 3s electron (IE = 1.95 MJ/mol). All other electrons have ionization energies greater than the energy of the photon.

90. If l is 1, then the allowed values for m_l are 1, 0, and -1. If l is 2, then the allowed values for m_l are 2, 1, 0, -1, and -2.
93. When n is 4, the allowed values for l are 0, 1, 2, and 3 (or s, p, d, and f orbitals). The difference between these orbitals is in their shape, although all are of the same size in that the electron, on average, as at the same distance from the nucleus.
94. When the principle quantum number, n , is 5, the maximum value for the angular quantum number, l , is 4.
95. For $n = 1$, the possible quantum numbers are
- $$n = 1, l = 0, m_l = 0$$
- For $n = 2$, the possible quantum numbers are
- $$n = 2, l = 0, m_l = 0$$
- $$n = 2, l = 1, m_l = -1$$
- $$n = 2, l = 1, m_l = 0$$
- $$n = 2, l = 1, m_l = +1$$
96. For $l = 0$ the symbol is s, for $l = 1$ the symbol is p, for $l = 2$ the symbol is d, for $l = 3$ the symbol is f.
97. For $n = 3$ we can have s, p, and d orbitals. There is one s orbital, three p orbitals, and five d orbitals for a total of 9 orbitals.
- For $n = 4$ we can have s, p, d, and f orbitals. There is one s orbital, three p orbitals, five d orbitals, and 7 f orbitals for a total of 16 orbitals.

For $n = 5$ we can have s, p, d, f, and g orbitals. There is one s orbital, three p orbitals, five d orbitals, 7 f orbitals, and 9 g orbitals for a total of 25 orbitals.

98. A 2d orbital is impossible because the smallest n for which we can have $l = 2$ (a d-orbital) is $n = 3$.

A 3f orbital is impossible because the smallest n for which we can have $l = 3$ (an f-orbital) is $n = 4$.

99. Paired electrons are two electrons that occupy the same orbital and have opposite spins.

102. The SG experiments will split a beam of atoms into two beams if the atom has one or more unpaired electrons; otherwise, the beam remains unsplit.

Element	Electron Configuration	No. Unpaired e^-	Result
H	$1s^1$	1	split
He	$1s^2$	0	unsplit
Be	$1s^2 2s^2$	0	unsplit
F	$1s^1 2s^2 2p^5$	1	split

104. Set (a) is not legitimate because l cannot be negative. Set (b) is not legitimate because m_l is limited to the range ± 2 since l is 2. Set (c) is not legitimate because m_s can only be $\pm 1/2$. Set (d) is legitimate.

111. To be a 2p electron it must have $n = 2$, $l = 1$, $m_l = +1, 0, \text{ or } -1$, and $m_s = \pm 1/2$. Of the choices, only (a) satisfies these conditions.

112. For $n = 4$ and $l = 3$, the possible values of m_l are +3, +2, +1, 0, -1, -2, and -3. Each of these orbitals can have $m_s = +1/2$ or $-1/2$; thus there are 14 possible electrons.

113. For $n = 4$ we can have one 4s orbital, three 4p orbitals, five 4d orbitals, and seven 4f orbitals. Each of these can have an electron with $m_s = +1/2$ for a total of 16 electrons.

114. The allowed values of m_s are always $+1/2$ or $-1/2$ regardless of the value of n , l , or m_l .

125. Of the choices provided, (e) is incorrect because 5f comes after 6s. Note, the problem is not asking if these are the exact order in which the orbitals are filled, but whether the relative order is correct.

126. The 6p orbitals are filled after the 4f orbitals.

129. The first element with a 3d electron is Sc, which is in the 4th row and the 3rd column.
131. All that matters in identifying the group is to look at the outermost, or valence electrons. In this case it is 5p³, which is the nitrogen group (VA or 15). The element is in the fourth row and is antimony, Sb.
132. Counting along the periodic table, element 119 will be an alkali metal located just below Francium, Fr.
137. Si is the second element in its period of the p-block; thus $x = 2$.
138. P is $1s^2 2s^2 2p^6 3s^2 3p^3$ or $[\text{Ne}] 3s^2 3p^3$; thus P^{3-} will be $[\text{Ne}] 3s^2 3p^6$.
141. Ti is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$. Losing two electrons to give Ti^{2+} gives an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2$ because the 4s electrons are the lowest in energy. The number of remaining s-electrons is six.
145. Only (b) satisfies Hund's rule. For (a) and (c) the spins must be unpaired. For (d) the two electrons should not be in the same orbital.
147. (a) This is incorrect because we must put one electron into each p-orbital before pairing up electrons.
- (b) This is incorrect because the s-orbital should be filled before the p-orbital takes additional electrons.
- (c) This is correct.
- (d) This is incorrect because the s-orbital must be filled before the p-orbital takes additional electrons.
148. Na is $1s^2 2s^2 2p^6 3s^1$ for a total of one unpaired electron in the 3s orbital.
- Al is $1s^2 2s^2 2p^6 3s^2 3p^1$ for a total of one unpaired electron in the 3p orbital.
- Si is $1s^2 2s^2 2p^6 3s^2 3p^2$ for a total of two unpaired electrons in the 3p orbital.
- P is $1s^2 2s^2 2p^6 3s^2 3p^3$ for a total of three unpaired electrons in the 3p orbital.
- S is $1s^2 2s^2 2p^6 3s^2 3p^4$ for a total of two unpaired electrons in the 3p orbital.
- Thus, phosphorous has the most unpaired electrons.
149. Ti^{4+} is $1s^2 2s^2 2p^6 3s^2 3p^6$ for a total of no unpaired electrons.

Co^{2+} is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7$ for a total of three unpaired electrons.

V^{3+} is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2$ for a total of two unpaired electrons.

Fe^{2+} is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$ for a total of five unpaired electrons.

Zn^{2+} is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$ for a total of no unpaired electrons.

Thus, Fe^{2+} has the most unpaired electrons.

155. The smallest atomic radius is the element furthest to the right and closest to the top of the periodic table; thus, of the five choices, aluminum will have the smallest atomic radius.
157. The largest atomic radius is the element furthest to the left and closest to the bottom of the periodic table; thus, of the five choices, phosphorous will have the largest atomic radius.
159. The large increase in atomic radius from Xe to Cs is due to the change from the $n = 5$ shell to the $n = 6$ shell; increasing n leads to electrons that, on average, are significantly further from the nucleus.
161. With two fewer electrons in Pb^{4+} the nucleus will exert a stronger pull on the electrons, leading to a smaller ionic radius.
162. For the same reason given in problem 161, we expect that the ionic radius for Fe^{2+} will be larger than that for Fe^{3+} .
165. One should predict that the size of these anions should increase as you go down a column in the periodic table (note – when hydrogen has a -1 charge it can be considered as the first element in the halide group); thus, we predict



The actual ionic radii given in Appendix B.4 is consistent with this ordering, except that the ionic radius for H^- falls between that for F^- and Cl^- . The best explanation here is that H^- , with 2 electrons and 1 proton, has a much greater relative proportion of electrons to protons than for any other halide. The fluoride ion, for example has 10 electrons and 9 protons.

166. The Al^{3+} and Mg^{2+} ions are isoelectronic (that is, they have the same electron configuration). The Al^{3+} ion will be smaller because it has more protons in the nucleus and the greater core charge will pull the electrons in more closely.

168. The smallest will be Mg^{2+} . We know that Na^+ and Mg^{2+} will be smaller than their respective elemental form and that Mg^+ will be smaller than Na^+ for the same reasons outlined in the solution to problem 166. The oxide ion, O^{2-} , as an anion will be larger than Mg^{2+} .
171. The largest ion certainly won't be one of the cations as a cation is always smaller than its respective element. Of the two anions, P^{3-} and S^{2-} , the more highly charged anion, P^{3-} , will be the largest.
175. The largest ionization energy for this series of species will be for the one that has already lost the most electrons; thus, P^{4+} will have the largest ionization energy.
177. There is a big jump between the third and fourth ionization energies, suggesting that there are three valence electrons. Of the five choices, $[\text{Ne}]3s^23p^1$ is the only one with three valence electrons.
179. There are two considerations here. First, an element with only one valence electron will show a big difference between its first and second ionization energies. Second, ionization energies increase as we move toward the right and toward the top of the periodic table. The first effect, however, is more important than the second; thus, of the five choices Na will have the largest second ionization energy.
184. Beryllium has a smaller AVEE than oxygen and is more likely to lose electrons than oxygen.
193. As you move across a period two things happen: first, the AVEE increases and the difference between the IE of s-electrons and p-electrons becomes larger. Metals have smaller AVEEs and require a smaller difference between the energies of the s and p electrons.