Assess

In comparing equations (9.9) and (9.10), we find that the difference in the slope (coefficient for Z) is due to the difference in the principal quantum number of the orbital from which the ionization occurs. The difference in the intercepts is due both to the principal quantum number from which the ionization occurs and to the number of electrons screening the valence-shell electron.

Equation (9.9) for the second-row elements is in remarkable agreement with the empirically observed equation (9.6) at the beginning of this example, especially considering we started from an equation for the energy levels of a one-electron species.

Although the general form of equation (9.10) is correct for the third-row series, the agreement between the numerical constants is not as good. This is to be expected because we have assumed perfect screening by the core electrons, which completely ignores the different characteristics of the electrons composing the inner core. The intricate, correlated motions of the electrons in the core leads to a complicated combination of screening and penetration that cannot be accounted for with our simple model.

PRACTICE EXAMPLE A: Francium ($Z = 87$) is an extremely rare radioactive element formed when actinium ($Z = 89$) undergoes alpha-particle emission. Francium occurs in natural uranium minerals, but estimates are that little more than 15 g of francium exists in the top 1 km of Earth's crust. Few of francium's properties have been measured, but some can be inferred from its position in the periodic table. Estimate the melting point, density, and atomic (metallic) radius of francium. [*Hint:* Plot each property versus atomic number, Z, and extrapolate to $Z = 87$.]

PRACTICE EXAMPLE B: Discuss the likelihood that element 168, should it ever be synthesized in sufficient quantity, would be a "noble liquid" at 298 K and 1 bar. Could element 168 be a "noble solid" at 298 K and 1 bar? Use *spdf* notation to show the electron configuration you would expect for element 168. [*Hint:* Prepare graphs of boiling point versus atomic number and melting point versus atomic number. Extrapolate to $Z = 168$.]

Element	mp, K	bp, K
Argon	83.95	87.45
Helium		4.25
Krypton	116.5	120.9
Neon	24.48	27.3
Radon	202	211.4
Xenon	161.3	166.1

Exercises

The Periodic Law

1. Use data from Figure 9-1 and equation (9.1) to estimate the density of the recently discovered element 114.
2. Suppose that lanthanum ($Z = 57$) were a newly discovered element having a density of 6.145 g/cm^3. Estimate its molar mass.
3. The following densities, in grams per cubic centimeter, are for the listed elements in their standard states at 298 K. Show that density is a periodic property of these elements: Al, 2.699; Ar, 0.0018; As, 5.778; Br, 3.100; Ca, 1.550; Cl, 0.0032; Ga, 5.904; Ge, 5.323; Kr, 0.0037; Mg, 1.738; P, 1.823; K, 0.856; Se, 4.285; Si, 2.336; Na, 0.968; S, 2.069.
4. The following melting points are in degrees Celsius. Show that melting point is a periodic property of these elements: Al, 660; Ar, −189; Be, 1278; B, 2300; C, 3350; Cl, −101; F, −220; Li, 179; Mg, 651; Ne, −249; N, −210; O, −218; P, 590; Si, 1410; Na, 98; S, 119.

The Periodic Table

5. Mendeleev's periodic table did not preclude the possibility of a new group of elements that would fit within the existing table, as was the case with the noble gases. Moseley's work did preclude this possibility. Explain this difference.
6. Explain why the several periods in the periodic table do not all have the same number of members.
7. Assuming that the seventh period is 32 members long, what should be the atomic number of the noble gas following radon (Rn)? Of the alkali metal following francium (Fr)? What would you expect their approximate atomic masses to be?
8. Concerning the incomplete seventh period of the periodic table, what should be the atomic number of the element **(a)** for which the filling of the 6*d* subshell is completed; **(b)** that should most closely resemble bismuth; **(c)** that should be a noble gas?

Atomic Radii and Ionic Radii

9. For each of the following pairs, indicate the atom that has the *larger* size: **(a)** Te or Br; **(b)** K or Ca; **(c)** Ca or Cs; **(d)** N or O; **(e)** O or P; **(f)** Al or Au.
10. Indicate the *smallest* and the *largest* species (atom or ion) in the following group: Al atom, F atom, As atom, Cs^+ ion, I^- ion, N atom.
11. Explain why the radii of atoms do not simply increase uniformly with increasing atomic number.
12. The masses of individual atoms can be determined with great precision, yet there is considerable uncertainty about the exact size of an atom. Explain why this is the case.
13. Which is **(a)** the smallest atom in group 13; **(b)** the smallest of the following atoms: Te, In, Sr, Po, Sb? Why?
14. How would you expect the sizes of the hydrogen ion, H^+, and the hydride ion, H^-, to compare with that of the H atom and the He atom? Explain.
15. Arrange the following in expected order of increasing radius: Br, Li^+, Se, I^-. Explain your answer.
16. Explain why the generalizations presented in Figure 9-19 cannot be used to answer the question, Which is larger, an Al atom or an I atom?
17. Among the following ions, several pairs are *isoelectronic*. Identify these pairs. Fe^{2+}, Sc^{3+}, Ca^{2+}, F^-, Co^{2+}, Co^{3+}, Sr^{2+}, Cu^+, Zn^{2+}, Al^{3+}.
18. The following species are isoelectronic with the noble gas krypton. Arrange them in order of increasing radius and comment on the principles involved in doing so: Rb^+, Y^{3+}, Br^-, Sr^{2+}, Se^{2-}.
19. All the isoelectronic species illustrated in the text had the electron configurations of noble gases. Can two ions be isoelectronic *without* having noble-gas electron configurations? Explain.
20. Is it possible for two different atoms to be isoelectronic? two different cations? two different anions? a cation and an anion? Explain.

Ionization Energies; Electron Affinities

21. Use principles established in this chapter to arrange the following atoms in order of *increasing* value of the first ionization energy: Sr, Cs, S, F, As.
22. Are there any atoms for which the second ionization energy is smaller than the first? Explain.
23. Some electron affinities are negative quantities, and some are zero or positive. Why is this not also the case with ionization energies?
24. How much energy, in joules, must be absorbed to convert to Na^+ all the atoms present in 1.00 mg of *gaseous* Na? The first ionization energy of Na is 495.8 kJ/mol.
25. How much energy, in kilojoules, is required to remove all the third-shell electrons in a mole of gaseous silicon atoms?
26. What is the maximum number of Cs^+ ions that can be produced per joule of energy absorbed by a sample of gaseous Cs atoms?
27. The production of gaseous bromide ions from bromine molecules can be considered a two-step process in which the first step is

$$Br_2(g) \longrightarrow 2\,Br(g) \qquad \Delta_r H = +193\ \text{kJ mol}^{-1}$$

Is the formation of $Br^-(g)$ from $Br_2(g)$ an endothermic or exothermic process?
28. Use ionization energies and electron affinities listed in the text to determine whether the following reaction is endothermic or exothermic.

$$Mg(g) + 2\,F(g) \longrightarrow Mg^{2+}(g) + 2\,F^-(g)$$

"ease of loss" is a misnomer; here it means: "energy cost of removing"

29. The Na^+ ion and the Ne atom are isoelectronic. The ease of loss of an electron by a gaseous Ne atom, first ionization energy, has a value of 2081 kJ/mol. The ease of loss of an electron from a gaseous Na^+ ion, second ionization energy, has a value of 4562 kJ/mol. Why are these values not the same?
30. From the data in Figure 9-12, the formation of a gaseous anion Li^- appears energetically favorable. That is, energy is given off when gaseous Li atoms accept electrons. Comment on the likelihood of forming a stable compound containing the Li^- ion, such as Li^+Li^- or Na^+Li^-.
31. Compare the elements Al, Si, S, and Cl.
(a) Place the elements in order of increasing ionization energy.
(b) Place the elements in order of increasing electron affinity.
(c) Place the elements in order of increasing polarizability.
32. Compare the elements Na, Mg, O, and P.
(a) Place the elements in order of increasing ionization energy.
(b) Place the elements in order of increasing electron affinity.
(c) Place the elements in order of increasing polarizability.

Magnetic Properties

33. Unpaired electrons are found in only one of the following species. Indicate which one, and explain why: F^-, Ca^{2+}, Fe^{2+}, S^{2-}.
34. Which of the following species has the greatest number of unpaired electrons **(a)** Ge; **(b)** Cl; **(c)** Cr^{3+}; **(d)** Br^-?
35. Which of the following species [in the gas-phase] would you expect to be diamagnetic and which paramagnetic? **(a)** K^+; **(b)** Cr^{3+}; **(c)** Zn^{2+}; **(d)** Cd; **(e)** Co^{3+}; **(f)** Sn^{2+}; **(g)** Br.
36. Write electron configurations consistent with the following data on numbers of unpaired electrons: Ni^{2+}, 2; Cu^{2+}, 1; Cr^{3+}, 3.

37. Must all atoms with an odd atomic number be paramagnetic? Must all atoms with an even atomic number be diamagnetic? Explain.

38. Neither Co^{2+} nor Co^{3+} has 4*s* electrons in its electron configuration. How many unpaired electrons would you expect to find in each of these ions? Explain.

Predictions Based on the Periodic Table

39. Use ideas presented in this chapter to indicate **(a)** three metals that you would expect to exhibit the photoelectric effect with visible light and three that you would not; **(b)** the noble gas element that should have the highest density in the liquid state; **(c)** the approximate first ionization energy E_i of fermium ($Z = 100$); **(d)** the approximate density of solid radium ($Z = 88$).

40. Arrange the following atoms in order of increasing polarizability: F, Na, P, As, Br.

41. Arrange the following species in order of increasing polarizability: N, S, Be, K, O.

42. For the following groups of elements, select the one that has the property noted:
(a) the largest atom: Mg, Mn, Mo, Ba, Bi, Br
(b) the lowest first ionization energy: B, Sr, Al, Br, Mg, Pb
(c) the most negative electron affinity: As, B, Cl, K, Mg, S
(d) the largest number of unpaired electrons: F, N, S^{2-}, Mg^{2+}, Sc^{3+}, Ti^{3+}

43. Of the species Cl, Cl^+, and Cl^-, which has the highest polarizability? Which has lowest polarizability?

44. Of the species Na^+, Na, F, and F^-, which has the highest polarizability? Which has lowest polarizability?

45. Match each of the lettered items on the left with an appropriate numbered item on the right. All the numbered items should be used, and some more than once.

(a) $Z = 32$	**1.** two unpaired *p* electrons
(b) $Z = 8$	**2.** diamagnetic
(c) $Z = 53$	**3.** more negative electron affinity than elements on either side of it in the same period
(d) $Z = 38$	**4.** first ionization energy lower than that of Ca but greater than that of Cs
(e) $Z = 48$	
(f) $Z = 20$	

46. Match each of the lettered items in the column on the left with the most appropriate numbered item(s) in the column on the right. Some of the numbered items may be used more than once and some not at all.

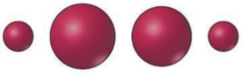
(a) Tl	**1.** an alkaline earth metal
(b) $Z = 70$	**2.** element in period 5 and group 15
(c) Ni	**3.** largest atomic radius of all the elements
(d) $[Ar]4s^2$	**4.** an element in period 4 and group 16
(e) a metalloid	**5.** $3d^8$
(f) a nonmetal	**6.** one *p* electron in the shell of highest *n*
	7. lowest ionization energy of all the elements
	8. an *f*-block element

47. Which of the following ions are unlikely to be found in chemical compounds: K^+, Ga^{4+}, Fe^{6+}, S^{2-}, Ge^{5+}, or Br^-? Explain briefly.

48. Which of the following ions are likely to be found in chemical compounds: Na^{2+}, Li^+, Al^{4+}, F^{2-}, or Te^{2-}? Explain briefly.

Integrative and Advanced Exercises

49. Four atoms and/or ions are sketched below in accordance with their relative atomic and/or ionic radii.



Which of the following sets of species are compatible with the sketch? Explain. **(a)** C, Ca^{2+}, Cl^-, Br^-; **(b)** Sr, Cl, Br^-, Na^+; **(c)** Y, K, Ca, Na^+; **(d)** Al, Ra^{2+}, Zr^{2+}, Mg^{2+}; **(e)** Fe, Rb, Co, Cs.

50. Sketch a periodic table that would include *all* the elements in the main body of the table. How many "numbers" wide would the table be?

51. In Mendeleev's time, indium oxide, which is 82.5% In by mass, was thought to be InO. If this were the case, in which group of Mendeleev's table (page 378) should indium be placed?

52. Instead of accepting the atomic mass of indium implied by the data in Exercise 51, Mendeleev proposed that the formula of indium oxide is In_2O_3. Show that this assumption places indium in the proper group of Mendeleev's periodic table on page 378.

53. Refer to Figure 9-11 and explain why the difference between the ionic radii of the −1 and −2 anions does not remain constant from top to bottom of the periodic table.

54. Explain why the third ionization energy of Li(g) is an easier quantity to calculate than either the first or second ionization energies. Calculate the third ionization energy for Li, and express the result in kJ/mol.

55. Two elements, A and B, have the electron configurations shown.

$$A = [Ar]4s^1 \quad B = [Ar]3d^{10}4s^24p^3$$

(a) Which element is a metal?
(b) Which element has the greater ionization energy?
(c) Which element has the larger atomic radius?
(d) Which element has the greater electron affinity?

56. Two elements, A and B, have the electron configurations shown.

$$A = [Kr]4s^2 \quad B = [Ar]3d^{10}4s^24p^5$$

(a) Which element is a metal?
(b) Which element has the greater ionization energy?
(c) Which element has the larger atomic radius?
(d) Which element has the greater electron affinity?

57. Studies done in 1880 showed that a chloride of uranium had 37.34% Cl by mass and an approximate formula mass of 382 u. Other data indicated the specific heat of uranium to be 0.0276 cal $g^{-1}\,°C^{-1}$. Are these data in agreement with the atomic mass of uranium assigned by Mendeleev, 240 u? [*Hint:* Refer to Feature Problem 124 of Chapter 7.]

58. Assume that atoms are hard spheres, and use the metallic radius of 186 pm for Na to estimate the volumes of one Na atom and of one mole of Na atoms. How does your result compare with the atomic volume found in Figure 9-1? Why is there so much disagreement between the two values?

59. When sodium chloride is strongly heated in a flame, the flame takes on the yellow color associated with the emission spectrum of sodium atoms. The reaction that occurs in the *gaseous* state is

$$Na^+(g) + Cl^-(g) \longrightarrow Na(g) + Cl(g).$$

Calculate $\Delta_r H$ for this reaction.

60. Use information from Chapters 8 and 9 to calculate the *second* ionization energy for the He atom. Compare your result with the tabulated value of 5251 kJ mol^{-1}.

61. Refer only to the periodic table on the inside front cover, and arrange the following ionization energies in order of increasing value: the first ionization energy of F; the second ionization energy of Ba; the third ionization energy of Sc; the second ionization energy of Na; the third ionization energy of Mg. Explain the basis of any uncertainties.

62. Refer to the footnote on page 393. Then use values of basic physical constants and other data from the appendices to show that 1 eV/atom = 96.49 kJ mol^{-1}.

63. The ionization energies of Li, Be^+, B^{2+}, and C^{3+} are, respectively, 520, 1757, 3659, and 6221 kJ mol^{-1}. The ionization energies Na, Mg^+, Al^{2+}, and Si^{3+} are (from Table 9.4) 495.8, 1451, 2745, and 4356 kJ mol^{-1}. Plot a graph of the square roots of the ionization energies versus the nuclear charge for these two series. Explain the observed relationship with the aid of Bohr's expression for the binding energy of an electron in a one-electron atom.

64. Elements 114–116 have recently been reported to be synthesized. Using data given below and the periodic law, fill in the missing data for these elements.

Sn	50	Sb	51	Te	52
$5s^25p^2$	118.7	$5s^25p^3$	121.8	$5s^25p^4$	127.6
2	145	3	145	1	140
16.29	107.3	18.19	103.2	20.46	190.2
708.6	7.31	834	6.69	869.3	6.24
Pb	82	Bi	83	Po	84
$6s^26p^2$	207.2	$6s^26p^3$	208.9	$6s^26p^4$	209
2	180	3	160	1	190
18.26	35.1	21.31	91.2	22.97	183.3
715.6	11.35	703	9.75	812.1	9.3
Fl	114	Unp	115	Lv	116
?	?	?	?	?	?
?	?	?	?	?	?
?	?	?	?	?	?
?	?	?	?	?	?

The entries for each element are organized as follows:

Atomic symbol	Z
Valence configuration	Atomic mass
No. of unpaired electrons	Atomic radius (pm)
Molar volume (cm^3)	Electron affinity (kJ mol^{-1})
First ionization energy (kJ mol^{-1})	Density (g cm^{-3})

Feature Problems

65. The work functions for a number of metals are given in the following table. How do the work functions vary
(a) down a group?
(b) across a period?
(c) Estimate the work function for potassium and compare it with a published value.
(d) What periodic property is the work function most like?

Metal	Work Function, J × 10^{19}
Al	6.86
Cs	3.45
Li	4.6
Mg	5.86
Na	4.40
Rb	3.46

66. The following are a few elements and their characteristic X-ray wavelengths:

Element	X-ray Wavelength, pm
Mg	987
S	536
Ca	333
Cr	229
Zn	143
Rb	93

Use these data to determine the constants A and b in Moseley's relationship (page 379). Compare your value of A with the value obtained from Bohr's theory for the frequencies emitted by one-electron atoms. Suggest a reasonable interpretation of the quantity b.

67. Gaseous sodium atoms absorb quanta with the energies shown in the table below.

Energy of Quanta, kJ mol^{-1}	Electron Configuration
0	$[Ne]3s^1$
203	$[Ne]3p^1$
308	$[Ne]4s^1$
349	$[Ne]3d^1$
362	$[Ne]4p^1$

(a) The ionization energy of the ground state is 496 kJ mol^{-1}. Calculate the ionization energies for each of the states given in the table.
(b) Calculate Z_{eff} for each state.
(c) Calculate $\bar{r}_{n\ell}$ for each state.
(d) Interpret the results obtained from parts (b) and (c) in terms of penetration and screening.

68. A method for estimating electron affinities is to extrapolate Z_{eff} values for atoms and ions that contain the same number of electrons as the negative ion of interest. Use the data in the table to answer the questions that follow.

Atom or Ion: E_i(kJ mol^{-1})	Atom or Ion: E_i(kJ mol^{-1})	Atom or Ion: E_i(kJ mol^{-1})
Ne: 2080	F: 1681	O: 1314
Na^+: 4565	Ne^+: 3963	F^+: 3375
Mg^{2+}: 7732	Na^{2+}: 6912	Ne^{2+}: 6276
Al^{3+}: 11,577	Mg^{3+}: 10,548	Na^{3+}: 9540

(a) Estimate the electron affinity of F, and compare it with the experimental value.
(b) Estimate the electron affinities of O and N.
(c) Examine your results in terms of penetration and screening.

69. We have seen that the wave functions of hydrogen-like atoms contain the nuclear charge Z for hydrogen-like atoms and ions, but modified through equation (9.3) to account for the phenomenon of shielding or screening. In 1930, John C. Slater devised the following set of empirical rules to calculate a shielding constant for a designated electron in the orbital ns or np:

(i) Write the electron configuration of the element, and group the subshells as follows: $(1s)$, $(2s, 2p)$, $(3s, 3p)$, $(3d)$, $(4s, 4p)$, $(4d)$, $(4f)$, $(5s, 5p)$, etc.
(ii) Electrons in groups to the right of the (ns, np) group contribute nothing to the shielding constant for the designated electron.
(iii) All the other electrons in the (ns, np) group shield the designated electron to the extent of 0.35 each.
(iv) All electrons in the $n - 1$ shell shield to the extent of 0.85 each.
(v) All electrons in the $n - 2$ shell, or lower, shield completely—their contributions to the shielding constant are 1.00 each.

When the designated electron being shielded is in an nd or nf group, rules (ii) and (iii) remain the same but rules (iv) and (v) are replaced by

(vi) Each electron in a group lying to the left of the nd or nf group contributes 1.00 to the shielding constant.

These rules are a simplified generalization based on the average behavior of different types of electrons. Use these rules to do the following:
(a) Calculate Z_{eff} for a valence electron of oxygen.
(b) Calculate Z_{eff} for the 4s electron in Cu.
(c) Calculate Z_{eff} for a 3d electron in Cu.
(d) Evaluate the Z_{eff} for the valence electrons in the group 1 elements (including H), and show that the ionization energies observed for this group are accounted for by using the Slater rules. [*Hint:* Do not overlook the effect of n on the orbital energy.]
(e) Evaluate Z_{eff} for a valence electron in the elements Li through Ne, and use the results to explain the observed trend in first ionization energies for these elements.
(f) Using the radial functions given in Table 8.2 and Z_{eff} estimated with the Slater rules, compare plots of the radial probability for the 3s, 3p, and 3d orbitals for the H atom and the Na atom. What do you observe from these plots regarding the effect of shielding on radial probability distributions?

Do as many as possible!

Self-Assessment Exercises

70. In your own words, define the following terms: **(a)** isoelectronic; **(b)** valence-shell electrons; **(c)** metal; **(d)** nonmetal; **(e)** metalloid.

71. Briefly describe each of the following ideas or phenomena: **(a)** the periodic law; **(b)** ionization energy; **(c)** electron affinity; **(d)** paramagnetism.

72. Explain the important distinctions between each pair of terms: **(a)** actinide and lanthanide element; **(b)** covalent and metallic radius; **(c)** atomic number and effective nuclear charge; **(d)** ionization energy and electron affinity; **(e)** paramagnetic and diamagnetic.

73. The element whose atoms have the electron configuration $[Kr]4d^{10}5s^25p^3$ **(a)** is in group 13 of the periodic table; **(b)** bears a similarity to the element Bi; **(c)** is similar to the element Te; **(d)** is a transition element.

74. The fourth-period element with the largest atom is **(a)** K; **(b)** Br; **(c)** Pb; **(d)** Kr.

75. Which of the following has the largest radius **(a)** an Ar atom; **(b)** a K^+ ion; **(c)** a Ca^{2+} ion; **(d)** a Cl^- ion?

76. The highest first ionization energy of the following is that of **(a)** Cs; **(b)** Cl; **(c)** I; **(d)** Li.

77. The most negative electron affinity of the following elements is that of **(a)** Br; **(b)** Sn; **(c)** Ba; **(d)** Li.

78. An ion that is isoelectronic with Se^{2-} is **(a)** S^{2-}; **(b)** I^-; **(c)** Xe; **(d)** Sr^{2+}.

79. Write electron configurations to show the first two ionizations for Cs. Explain why the second ionization energy is much greater than the first.

80. Explain why the first ionization energy of Mg is greater that of Na, whereas the second ionization of Na is greater than that of Mg.

81. Answer each of the following questions:
(a) Which of the elements P, As, and S has the largest atomic radius?
(b) Which of the following has the smallest radius: Xe, O^{2-}, N^{3-}, or F^-?
(c) Which should have the largest difference between the first and second ionization energy: Al, Si, P, or Cl?
(d) Which has the largest ionization energy: C, Si, or Sn?
(e) Which has the largest electron affinity: Na, B, Al, or C?

82. The first ionization energies of Si, P, S, and Cl are given in Table 9.4. Briefly provide an explanation for this trend.

83. Find three pairs of elements that are out of order in the periodic table in terms of their atomic masses. Why is it necessary to invert their order in the table?

84. For the atom $^{119}_{50}Sn$, indicate the number of **(a)** protons in the nucleus; **(b)** neutrons in the nucleus; **(c)** $4d$ electrons; **(d)** $3s$ electrons; **(e)** $5p$ electrons; **(f)** electrons in the valence shell.

85. Refer to the periodic table on the inside front cover and indicate **(a)** the most nonmetallic element; **(b)** the transition metal with lowest atomic number; **(c)** a metalloid whose atomic number is exactly midway between those of two noble gas elements.

86. Give the symbol of the element **(a)** in group 14 that has the smallest atoms; **(b)** in period 5 that has the largest atoms; **(c)** in group 17 that has the lowest first ionization energy.

87. Refer only to the periodic table on the inside front cover and indicate which of the atoms, Bi, S, Ba, As, and Ca, **(a)** is most metallic; **(b)** is most nonmetallic; **(c)** has the intermediate value when the five are arranged in order of increasing first ionization energy.

88. Arrange the following elements in order of decreasing metallic character: Sc, Fe, Rb, Br, O, Ca, F, Te.

89. In multielectron atoms many of the periodic trends can be explained in terms of Z_{eff}. Consider the following statements and discuss whether or not the statement is true or false.
(a) Electrons in a p orbital are more effective than electrons in the s orbitals in shielding other electrons from the nuclear charge.
(b) Z_{eff} for an electron in an s orbital is lower than that for an electron in a p orbital in the same shell.
(c) Z_{eff} is usually less than Z.
(d) Electrons in orbitals having $\ell = 1$ penetrate better than those with $\ell = 2$.
(e) Z_{eff} for the orbitals of the elements Na($3s$), Mg($3s$), Al($3p$), P($3p$), and S($3p$) are in the order $Z_{eff}(Na) < Z_{eff}(Mg) > Z_{eff}(Al) < Z_{eff}(P) > Z_{eff}(S)$.

90. Consider a nitrogen atom in the ground state and comment on whether the following statements are true or false.
(a) Z_{eff} for an electron in a $2s$ orbital is greater than that for the $1s$ orbital.
(b) The Z_{eff} for the $2p$ and $2s$ orbitals is the same.
(c) More energy is required to remove an electron from a $2s$ orbital than from the $2p$ orbital.
(d) The $2s$ electron is less shielded than the $2p$ electron.

91. Describe how the ionization energies of the ions $He^-, Li^-, Be^-, B^-, C^-, N^-, O^-$, and F^- vary with atomic number.

92. Describe how the ionization energies of the ions $Be^+, B^+, C^+, N^+, O^+, F^+, Ne^+$, and Na^+ vary with atomic number.

93. Which element Na or Mg is likely to have $\Delta_{ea}H$ greater than zero?

94. Why, in general, is the addition of an electron to an atom an exothermic process?

95. When compared to a nonmetal of the same period, a metal will have a larger **(a)** atomic radius; **(b)** ionization energy; **(c)** electron affinity; **(d)** atomic number; **(e)** none of these.

96. Which of the following is an example of a metalloid? **(a)** S; **(b)** Zn; **(c)** Ge; **(d)** Re; **(e)** none of these.

97. Which of the following has a smaller radius than a neon atom? **(a)** Mg^{2+}; **(b)** F^-; **(c)** O^{2-}; **(d)** K^+; **(e)** none of these.

98. Which electron is lost when an atom ionizes? **(a)** the electron with the highest principal quantum number; **(b)** the electron with lowest principal quantum number; **(c)** an outer-shell electron with the highest value of the orbital angular momentum quantum number; **(d)** the electron with highest orbital angular momentum quantum number; **(e)** none of these.

99. The electrons lost when Fe ionizes to Fe^{2+} are **(a)** $4f$; **(b)** $3d$; **(c)** $4s$; **(d)** $3p$; **(e)** none of these.

100. Construct a concept map (see Appendix E) connecting the ideas that govern the periodic law and the periodic variation of atomic properties.