Exercises

Review Questions

1. What are periodic properties?

2. Use aluminum as an example to explain how density is a periodic property.

3. Explain the contributions of Döbereiner and Newlands to the organization of elements according to their properties.

4. Who is credited with arranging the periodic table? How are elements arranged in this table?

5. Explain the contributions of Meyer and Moseley to the periodic table.

6. The periodic table is a result of the periodic law. What observations led to the periodic law? What theory explains the underlying reasons for the periodic law?

7. What is an electron configuration? Provide an example.

8. What is Coulomb's law? Explain how the potential energy of two charged particles depends on the distance between the charged particles and on the magnitude and sign of their charges.

9. What is shielding? In an atom, which electrons tend to do the most shielding (core electrons or valence electrons)?

10. What is penetration? How does the penetration of an orbital into the region occupied by core electrons affect the energy of an electron in that orbital?

11. Why are the sublevels within a principal level split into different energies for multi-electron atoms but not for the hydrogen atom?

12. What is an orbital diagram? Provide an example.

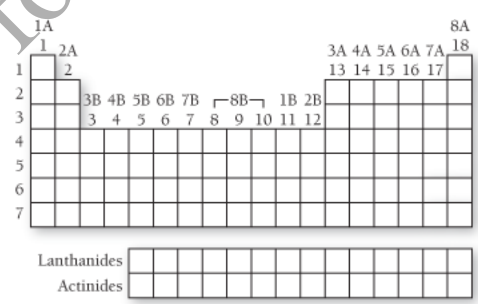
13. Why is electron spin important when writing electron configurations? Explain in terms of the Pauli exclusion principle.

14. What are degenerate orbitals? According to Hund's rule, how are degenerate orbitals occupied?

15. List all orbitals from 1*s* through 5*s* according to increasing energy for multi-electron atoms.

16. What are valence electrons? Why are they important?

17. Copy this blank periodic table onto a sheet of paper and label each of the blocks within the table: *s* block, *p* block, *d* block, and *f* block.



18. Explain why the *s* block in the periodic table has only two columns while the *p* block has six.

19. Explain why the rows in the periodic table become progressively longer as we move down the table. For example, the first row contains two elements, the second and third rows each contain eight elements, and the fourth and fifth rows each contain 18 elements.

20. Explain the relationship between a main-group element's lettered group number (the number of the element's column) and its valence electrons.

21. Explain the relationship between an element's row number in the periodic table and the highest principal quantum number in the element's electron configuration. How does this relationship differ for main-group elements, transition elements, and inner transition elements?

22. Which of the transition elements in the first transition series have anomalous electron configurations?

22. Which of the transition elements in the first transition series have anomalous electron configurations?

23. Explain how to write the electron configuration for an element based on its position in the periodic table.

24. Explain the relationship between the properties of an element and the number of valence electrons that it contains.

25. List the number of valence electrons for each family in the periodic table, and explain the relationship between the number of valence electrons and the resulting chemistry of the elements in the family.

- **a.** alkali metals
- **b.** alkaline earth metals
- **c.** halogens
- **d.** oxygen family

26. Define atomic radius. For main-group elements, describe the observed trends in atomic radius as we:

- **a.** move across a period in the periodic table
- **b.** move down a column in the periodic table

27. What is effective nuclear charge? What is shielding?

28. When an alkali metal forms an ion, what is the charge of the ion? What is the charge of an alkaline earth metal ion?

29. When a halogen forms an ion, what is the charge of the ion? When the nonmetals in the oxygen family form ions, what is the charge of the ions? What is the charge of the ions formed by N and Al?

30. Use the concepts of effective nuclear charge, shielding, and *n* value of the valence orbital to explain the trend in atomic radius as we move across a period in the periodic table.

31. For transition elements, describe the trends in atomic radius as we:

- **a.** move across a period in the periodic table
- **b.** move down a column in the periodic table

Explain the reasons for the trends described in parts a and b.

32. How is the electron configuration of an anion different from that of the corresponding neutral atom? How is the electron configuration of a cation different?

33. Explain how to write an electron configuration for a transition metal cation. Is the order of electron removal upon ionization simply the reverse of electron addition upon filling? Why or why not?

34. Describe the relationship between:

- **a.** the radius of a cation and the radius of the atom from which it is formed
- **b.** the radius of an anion and the radius of the atom from which it is formed

35. What is ionization energy? What is the difference between first ionization energy and second ionization energy?

36. What is the general trend in first ionization energy as we move down a column in the periodic table? As we move across a row?

37. What are the exceptions to the periodic trends in first ionization energy? Why do they occur?

38. Examination of the first few successive ionization energies for a given element usually reveals a large jump between two ionization energies. For example, the successive ionization energies of magnesium show a large jump between IE_2 and IE_3. The successive ionization energies of aluminum show a large jump between IE_3 and IE_4. Explain why these jumps occur and how we might predict them.

39. What is electron affinity? What are the observed periodic trends in electron affinity?

40. What is metallic character? What are the observed periodic trends in metallic character?

Problems by Topic

Note: Answers to all odd-numbered Problems can be found in Appendix III. Exercises in the Problems by Topic section are paired, with each odd-numbered problem followed by a similar even-numbered problem. Exercises in the Cumulative Problems section are also paired but more loosely. Because of their nature, Challenge Problems and Conceptual Problems are unpaired.

The Periodic Table

41. Write the name of each element and classify it as a metal, nonmetal, or metalloid.

- **a.** K
- **b.** Ba
- **c.** I
- **d.** O

e. Sb

42. Write the symbol for each element and classify it as a metal, nonmetal, or metalloid.

a. gold

b. fluorine

c. sodium

d. tin

e. argon

43. Determine whether each element is a main-group element.

a. tellurium

b. potassium

c. vanadium

d. manganese

44. Determine whether each element is a transition element.

a. Cr

b. Br

c. Mo

d. Cs

Electron Configurations

45. Write the full electron configuration for each element.

a. Si

b. O

c. K

d. Ne

46. Write the full electron configuration for each element.

a. C

b. P

c. Ar

d. Na

47. Write the full orbital diagram for each element.

a. N

b. F

c. Mg

d. Al

48. Write the full orbital diagram for each element.

a. S

b. Ca

c. Ne

d. He

49. Use the periodic table to write the electron configuration for each element. Represent core electrons with the symbol of the previous noble gas in brackets.

a. P

b. Ge

c. Zr

d. I

50. Use the periodic table to determine the element corresponding to each electron configuration.

a. $[Ar]\ 4s^2\ 3d^{10}\ 4p^6$

b. $[Ar]\ 4s^2\ 3d^2$

c. $[Kr]\,5s^2\ 4d^{10}\ 5p^2$

d. $[Kr]\,5s^2$

51. Use the periodic table to determine each quantity.

a. the number of 2*s* electrons in Li

b. the number of 3*d* electrons in Cu

c. the number of 4*p* electrons in Br

d. the number of 4*d* electrons in Zr

52. Use the periodic table to determine each quantity.

52. Use the periodic table to determine each quantity.

a. the number of $3s$ electrons in Mg

b. the number of $3d$ electrons in Cr

c. the number of $4d$ electrons in Y

d. the number of $6p$ electrons in Pb

53. Name an element in the fourth period (row) of the periodic table with:

a. five valence electrons

b. four $4p$ electrons

c. three $3d$ electrons

d. a complete outer shell

54. Name an element in the third period (row) of the periodic table with:

a. three valence electrons

b. four $3p$ electrons

c. six $3p$ electrons

d. two $3s$ electrons and zero $3p$ electrons

Valence Electrons and Simple Chemical Behavior from the Periodic Table

55. Determine the number of valence electrons in each element.

a. Ba

b. Cs

c. Ni

d. S

56. Determine the number of valence electrons in each element. Which elements do you expect to lose electrons in chemical reactions? Which do you expect to gain electrons?

a. Al

b. Sn

c. Br

d. Se

57. Which outer electron configuration would you expect to correspond to a reactive metal? To a reactive nonmetal?

a. ns^2

b. $ns^2\ np^6$

c. $ns^2\ np^5$

d. $ns^2\ np^2$

58. Which outer electron configuration would you expect to correspond to a noble gas? To a metalloid?

a. ns^2

b. $ns^2\ np^6$

c. $ns^2\ np^5$

d. $ns^2\ np^2$

59. List the number of valence electrons for each element and classify each element as an alkali metal, alkaline earth metal, halogen, or noble gas.

a. sodium

b. iodine

c. calcium

d. barium

e. krypton

60. List the number of valence electrons in each element and classify each element as an alkali metal, alkaline earth metal, halogen, or noble gas.

a. F

b. Sr

c. K

d. Ne

e. At

61. Which pair of elements do you expect to be most similar? Why?

a. N and Ni

b. Mo and Sn

c. Na and Mg

d. Cl and F

e. Si and P

62. Which pair of elements do you expect to be most similar? Why?

a. nitrogen and oxygen

b. titanium and gallium

c. lithium and sodium

d. germanium and arsenic

e. argon and bromine

63. Predict the charge of the ion formed by each element and write the electron configuration of the ion.

a. O

b. K

c. Al

d. Rb

64. Predict the charge of the ion formed by each element and write the electron configuration of the ion.

a. Mg

b. N

c. F

d. Na

Coulomb's Law and Effective Nuclear Charge

65. According to Coulomb's law, which pair of charged particles has the lowest potential energy?

a. a particle with a 1− charge separated by 150 pm from a particle with a 2+ charge

b. a particle with a 1− charge separated by 150 pm from a particle with a 1+ charge

c. a particle with a 1− charge separated by 100 pm from a particle with a 3+ charge

66. According to Coulomb's law, rank the interactions between charged particles from lowest potential energy to highest potential energy.

a. a 1+ charge and a 1− charge separated by 100 pm

b. a 2+ charge and a 1− charge separated by 100 pm

c. a 1+ charge and a 1+ charge separated by 100 pm

d. a 1+ charge and a 1− charge separated by 200 pm

67. Which electrons experience a greater effective nuclear charge: the valence electrons in beryllium or the valence electrons in nitrogen? Why?

68. Arrange the atoms according to decreasing effective nuclear charge experienced by their valence electrons: S, Mg, Al, Si.

69. If core electrons completely shielded valence electrons from nuclear charge (i.e., if each core electron reduced nuclear charge by one unit) and if valence electrons did not shield one another from nuclear charge at all, what would be the effective nuclear charge experienced by the valence electrons of each atom?

a. K

b. Ca

c. O

d. C

70. In Section 3.6, we estimated the effective nuclear charge on beryllium's valence electrons to be slightly greater than 2+. What would a similar treatment predict for the effective nuclear charge on boron's valence electrons? Would you expect the effective nuclear charge to be different for boron's 2*s* electrons compared to its 2*p* electron? In what way? (*Hint:* Consider the shape of the 2*p* orbital compared to that of the 2*s* orbital.)

Atomic Radius

71. Choose the larger atom in each pair.

a. Al or In

b. Si or N

c. P or Pb

d. C or F

72. Choose the larger atom in each pair.

a. Sn or Si

b. Br or Ga

c. Sn or Bi

d. Se or Sn

73. Arrange these elements in order of increasing atomic radius: Ca, Rb, S, Si, Ge, F.

74. Arrange these elements in order of decreasing atomic radius: Cs, Sb, S, Pb, Se.

Ionic Electron Configurations, Ionic Radii, Magnetic Properties, and Ionization Energy

75. Write the electron configuration for each ion.

a. O^{2-}

b. Br^{-}

c. Sr^{2+}

d. Co^{3+}

e. Cu^{2+}

76. Write the electron configuration for each ion.

a. Cl^{-}

b. P^{3-}

c. K^{-}

d. Mo^{3+}

e. V^{3+}

77. Write orbital diagrams for each ion and determine if the ion is diamagnetic or paramagnetic.

a. V^{5+}

b. Cr^{3+}

c. Ni^{2+}

d. Fe^{3+}

78. Write orbital diagrams for each ion and determine if the ion is diamagnetic or paramagnetic.

a. Cd^{2+}

b. Au^{+}

c. Mo^{3+}

d. Zr^{2+}

79. Which is the larger species in each pair?

a. Li or Li^{+}

b. I^{-} or Cs^{+}

c. Cr or Cr^{3+}

d. O or O^{2-}

80. Which is the larger species in each pair?

a. Sr or Sr^{2+}

b. N or N^{3-}

c. Ni or Ni^{2+}

d. S^{2-} or Ca^{2+}

81. Arrange this isoelectronic series in order of decreasing radius: $F^{-}, O^{2-}, Mg^{2+}, Na^{+}$.

82. Arrange this isoelectronic series in order of increasing atomic radius: $Se^{2-}, Sr^{2+}, Rb^{+}, Br^{-}$.

83. Choose the element with the higher first ionization energy in each pair.

a. Br or Bi

b. Na or Rb

c. As or At

d. P or Sn

84. Choose the element with the higher first ionization energy in each pair.

a. P or I

b. Si or Cl

c. P or Sb

d. Ga or Ge

85. Arrange these elements in order of increasing first ionization energy: Si, F, In, N.

86. Arrange these elements in order of decreasing first ionization energy: Cl, S, Sn, Pb.

87. For each element, predict where the "jump" occurs for successive ionization energies. (For example, does the jump occur between the first and second ionization energies, the second and third, or the third and fourth?)

a. Be

b. N

c. O

d. Li

88. Consider this set of successive ionization energies:

$IE_1 = 578\ kJ/mol \quad IE_2 = 1820\ kJ/mol$

$IE_3 = 2750\ kJ/mol \quad IE_4 = 11,600\ kJ/mol$

To which third-period element do these ionization values belong?

Electron Affinities and Metallic Character

89. Choose the element with the more negative (more exothermic) electron affinity in each pair.

a. Na or Rb

b. B or S

c. C or N

d. Li or F

90. Choose the element with the more negative (more exothermic) electron affinity in each pair.

a. Mg or S

b. K or Cs

c. Si or P

d. Ga or Br

91. Choose the more metallic element in each pair.

a. Sr or Sb

b. As or Bi

c. Cl or O

d. S or As

92. Choose the more metallic element in each pair.

a. Sb or Pb

b. K or Ge

c. Ge or Sb

d. As or Sn

93. Arrange these elements in order of increasing metallic character: Fr, Sb, In, S, Ba, Se.

94. Arrange these elements in order of decreasing metallic character: Sr, N, Si, P, Ga, Al.

Cumulative Problems

95. Bromine is a highly reactive liquid, whereas krypton is an inert gas. Explain the difference based on their electron configurations.

96. Potassium is a highly reactive metal, whereas argon is an inert gas. Explain the difference based on their electron configurations.

97. Both vanadium and its 3+ ion are paramagnetic. Use electron configurations to explain this statement.

98. Use electron configurations to explain why copper is paramagnetic while its 1+ ion is not.

99. Suppose you were trying to find a substitute for K^+ for some application. Where would you begin your search? Which ions are most like K^+? For each ion you propose, explain the ways in which it is similar to K^+ and the ways it is different. Refer to periodic trends in your discussion.

100. Suppose you were trying to find a substitute for Na^+ for some application. Where would you begin your search? What ions are most like Na^+? For each ion you propose, explain the ways in which it is similar to Na^+ and the ways it is different. Use periodic trends in your discussion.

101. Life on Earth evolved based on the element carbon. Based on periodic properties, what two or three elements would you expect to be most like carbon?

102. Which pair of elements would you expect to have the most similar atomic radii, and why?

a. Si and Ga

b. Si and Ge

c. Si and As

103. Consider these elements: N, Mg, O, F, Al.

a. Write the electron configuration for each element.

b. Arrange the elements in order of decreasing atomic radius.

c. Arrange the elements in order of increasing ionization energy.

d. Use the electron configurations in part a to explain the differences between your answers to parts b and c.

104. Consider these elements: P, Ca, Si, S, Ga.

a. Write the electron configuration for each element.

b. Arrange the elements in order of decreasing atomic radius.

c. Arrange the elements in order of increasing ionization energy.

d. Use the electron configurations in part a to explain the differences between your answers to parts b and c.

105. Explain why atomic radius decreases as we move to the right across a period for main-group elements but not for transition elements.

106. Explain why vanadium (radius $= 134$ pm) and copper (radius $= 128$ pm) have nearly identical atomic radii, even though the atomic number of copper is about 25% higher than that of vanadium. What would you predict about the relative densities of these two metals? Look up the densities in a reference book, periodic table, or on the Internet. Are your predictions correct?

107. The lightest noble gases, such as helium and neon, are completely inert—they do not form any chemical compounds whatsoever. The heavier noble gases, in contrast, do form a limited number of compounds. Explain this difference in terms of trends in fundamental periodic properties.

108. The lightest halogen is also the most chemically reactive, and reactivity generally decreases as we move down the column of halogens in the periodic table. Explain this trend in terms of periodic properties.

109. Write general outer electron configurations ($ns^x\ np^y$) for groups 6A and 7A in the periodic table. The electron affinity of each group 7A element is more negative than that of each corresponding group 6A element. Use the electron configurations to explain this observation.

110. The electron affinity of each group 5A element is more positive than that of each corresponding group 4A element. Use the outer electron configurations for these columns to suggest a reason for this behavior.

111. The elements with atomic numbers 35 and 53 have similar chemical properties. Based on their electronic configurations predict the atomic number of a heavier element that also should have these chemical properties.

112. Write the electronic configurations of the six cations that form from sulfur by the loss of one to six electrons. For those cations that have unpaired electrons, write orbital diagrams.

113. You have cracked a secret code that uses elemental symbols to spell words. The code uses numbers to designate the elemental symbols. Each number is the sum of the atomic number and the highest principal quantum number of the highest occupied orbital of the element whose symbol is to be used. Messages may be written forward or backward. Decode the following messages:

a. 10, 12, 58, 11, 7, 44, 63, 66

b. 9, 99, 30, 95, 19, 47, 79

114. The electron affinity of sodium is lower than that of lithium, while the electron affinity of chlorine is higher than that of fluorine. Suggest an explanation for this observation.

115. Use Coulomb's law to calculate the ionization energy in kJ/mol of an atom composed of a proton and an electron separated by 100.00 pm. What wavelength of light would have sufficient energy to ionize the atom?

116. The first ionization energy of sodium is 496 kJ/mol. Use Coulomb's law to estimate the average distance between the sodium nucleus and the 3s electron. How does this distance compare to the atomic radius of sodium? Explain the difference.

Challenge Problems

117. Consider the densities and atomic radii of the noble gases at 25 °C:

Element	Atomic Radius (pm)	Density (g/L)
He	32	0.18
Ne	70	0.90
Ar	98	–
Kr	112	3.75
Xe	130	–

Rn	–	9.73

a. Estimate the densities of argon and xenon by interpolation from the data.

b. Provide an estimate of the density of the yet undiscovered element with atomic number 118 by extrapolation from the data.

c. Use the molar mass of neon to estimate the mass of a neon atom. Then use the atomic radius of neon to calculate the average density of a neon atom. How does this density compare to the density of neon gas? What does this comparison suggest about the nature of neon gas?

d. Use the densities and molar masses of krypton and neon to calculate the number of atoms of each element found in a volume of 1.0 L. Use these values to estimate the number of atoms that occur in 1.0 L of Ar. Now use the molar mass of argon to estimate the density of Ar. How does this estimate compare to that in part a?

118. As you have seen, the periodic table is a result of empirical observation (i.e., the periodic law), but quantum-mechanical theory explains why the table is so arranged. Suppose that, in another universe, quantum theory was such that there were one *s* orbital but only two *p* orbitals (instead of three) and only three *d* orbitals (instead of five). Draw out the first four periods of the periodic table in this alternative universe. Which elements would be the equivalent of the noble gases? Halogens? Alkali metals?

119. Consider the metals in the first transition series. Use periodic trends to predict a trend in density as we move to the right across the series.

120. Imagine a universe in which the value of m_s can be +½, 0, and −½. Assuming that all the other quantum numbers can take only the values possible in our world and that the Pauli exclusion principle applies, determine:

a. the new electronic configuration of neon

b. the atomic number of the element with a completed $n = 2$ shell

c. the number of unpaired electrons in fluorine

121. A carbon atom can absorb radiation of various wavelengths with resulting changes in its electronic configuration. Write orbital diagrams for the electronic configurations of carbon that result from absorption of the three longest wavelengths of radiation that change its electronic configuration.

122. Only trace amounts of the synthetic element darmstadtium, atomic number 110, have been obtained. The element is so highly unstable that no observations of its properties have been possible. Based on its position in the periodic table, propose three different reasonable valence electron configurations for this element.

123. What is the atomic number of the as yet undiscovered element in which the 8*s* and 8*p* electron energy levels fill? Predict the chemical behavior of this element.

124. The trend in second ionization energy for the elements from lithium to fluorine is not a regular one. Predict which of these elements has the highest second ionization energy and which has the lowest and explain. Of the elements N, O, and F, O has the highest and N the lowest second ionization energy. Explain.

125. Unlike the elements in groups 1A and 2A, those in group 3A do not show a regular decrease in first ionization energy in going down the column. Explain the irregularities.

126. Using the data in Figures 3.19 and 3.20, calculate ΔE (the change in energy) for the reaction

$$\mathrm{Na}\,(g) + \mathrm{Cl}\,(g) \rightarrow \mathrm{Na}^{+}\,(g) + \mathrm{Cl}^{-}\,(g)$$

127. Even though adding two electrons to O or S forms an ion with a noble gas electron configuration, the second electron affinity of both of these elements is positive. Explain.

128. In Section 3.5 we discussed the metalloids, which form a diagonal band separating the metals from the nonmetals. There are other instances in which elements such as lithium and magnesium that are diagonal to each other have comparable metallic character. Suggest an explanation for this observation.

129. The heaviest known alkaline earth metal is radium, atomic number 88. Find the atomic numbers of the as yet undiscovered next two members of the series.

130. Predict the electronic configurations of the first two excited states (next higher energy states beyond the ground state) of Pd.

Conceptual Problems

131. Imagine that in another universe, atoms and elements are identical to ours, except that atoms with six valence electrons have particular stability (in contrast to our universe where atoms with eight valence electrons have particular stability). Give an example of an element in the alternative universe that corresponds to:

a. a noble gas

b. a reactive nonmetal

c. a reactive metal

132. The outermost valence electron in atom A experiences an effective nuclear charge of 2+ and is on average 225 pm from the nucleus. The outermost valence electron in atom B experiences an effective nuclear charge of 1+ and is on average 175 pm from the nucleus. Which atom (A or B) has the higher first ionization energy? Explain.

133. Determine whether each statement regarding penetration and shielding is true or false. (Assume that all lower energy orbitals are fully occupied.)

a. An electron in a 3*s* orbital is more shielded than an electron in a 2*s* orbital.

b. An electron in a 3*s* orbital penetrates into the region occupied by core electrons more than electrons in a 3*p* orbital.

c. An electron in an orbital that penetrates closer to the nucleus will always experience more shielding than an electron in an orbital that does not penetrate as far.

d. An electron in an orbital that penetrates close to the nucleus will tend to experience a higher effective nuclear charge than one that does not.

134. Give a combination of four quantum numbers that could be assigned to an electron occupying a 5*p* orbital. Do the same for an electron occupying a 6*d* orbital.

135. Use the trends in ionization energy and electron affinity to explain why calcium fluoride has the formula CaF_2 and not Ca_2F or CaF.

Questions for Group Work

Active Classroom Learning

Discuss these questions with the group and record your consensus answer.

136. In a complete sentence, describe the relationship between shielding and penetration.

137. Play a game to memorize the order in which orbitals fill. Have each group member in turn state the name of the next orbital to fill and the maximum number of electrons it can hold (for example, "1*s* two," "2*s* two," "2*p* six"). If a member gets stuck, other group members can help, consulting Figure 3.8 and the accompanying text summary if necessary. However, when a member gets stuck, the next player starts back at "1*s* two." Keep going until each group member can list all the orbitals in order up to "6*s* two."

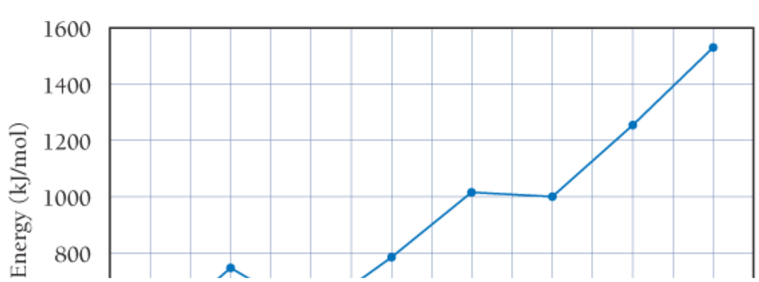
138. Sketch a periodic table (without element symbols). Include the correct number of rows and columns in the *s*, *p*, *d*, and *f* blocks. Shade in the squares for elements that have irregular electron configurations.

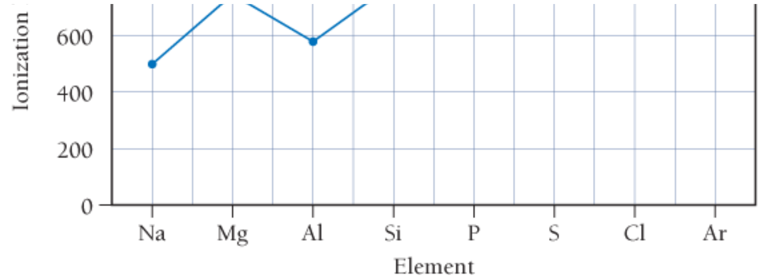
139. In complete sentences, explain: a) why Se^{2-} and Br^- are about the same size; b) why Br^- is slightly smaller than Se^{2-}; and c) which singly charged cation you would expect to be approximately the same size as Se^{2-} and Br^- and why.

140. Have each member of your group sketch a periodic table indicating a periodic trend (atomic size, first ionization energy, metallic character, etc.). Have each member present his or her table to the rest of the group and explain the trend based on concepts such as orbital size or effective nuclear charge.

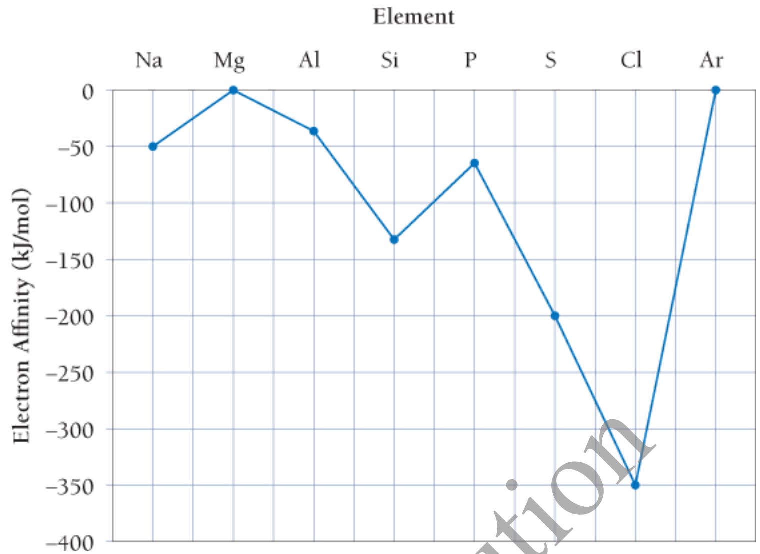
Data Interpretation and Analysis

141. The following graphs show the first ionization energies and electron affinities of the period 3 elements. Refer to the graphs to answer the questions that follow.





First Ionization Energies of Period 3 Elements



Electron Affinities of Period 3 Elements

a. Describe the general trend in period 3 first ionization energies as you move from left to right across the periodic table. Explain why this trend occurs.

b. The trend in first ionization energy has two execptions: one at Al and another S. Explain why the first ionization energy of Al is lower than that of Mg and why the first ionization of S is less than that of P.

c. Describe the general trend in period 3 electron affinities as you move from left to right across the periodic table. Explain why this trend occurs.

d. The trend in electron affinities has exceptions at Mg and P. Explain why the electron affinity of Mg is more positive (less exothermic) than that of Na and why the electron affinity of P is more positive (less exothermic) than that of Si.

e. Determine the overall energy change for removing one electron from Na and adding that electron to Cl. Is the exchange of the electron exothermic or endothermic?

Answers to Conceptual Connections

Cc 3.1 **(d)** Cr is in the transition elements section of the periodic table (see **Figure 3.4**).

Cc 3.2 **(a)** Since the charges are opposite, the potential energy of the interaction is negative. As the charges get closer together, r becomes smaller and the potential energy decreases (it becomes more negative).

Cc 3.3 **(c)** Penetration results in less shielding from nuclear charge and therefore lower energy.

Cc 3.4 $n = 4, l = 0, m_l = 0, m_s = +\frac{1}{2}; n = 4, l = 0, m_l = 0, m_s = -\frac{1}{2}$

Cc 3.5 **(c)** Because Z_{eff} increases from left to right across a row in the periodic table, the valence electrons in S experience a greater effective nuclear charge than the valence electrons in Al or in Mg.

Cc 3.6 The isotopes of an element all have the same radius for two reasons: (1) neutrons are negligibly small compared to the size of an atom, and therefore extra neutrons do not increase atomic size; and (2) neutrons have no charge and therefore do not attract electrons in the way that protons do.

Cc 3.7 As you can see from the successive ionization energies of any element, valence electrons are held most loosely and can therefore be transferred or shared most easily. Core electrons, in contrast, are held tightly and are not easily transferred or shared. Consequently, valence electrons play a central role in chemical bonding.

Cc 3.8 The 3*s* electron in sodium has a relatively low ionization energy (496 kJ/mol) because it is a valence electron. The energetic cost for sodium to lose a second electron is extraordinarily high (4560 kJ/mol) because the next electron to be lost is a core electron (2*p*). Similarly, the electron affinity of chlorine to gain one electron (–349 kJ/mol) is highly exothermic since the added electron completes chlorine's valence shell. The gain of a second electron by the negatively charged chlorine anion is not so favorable. Therefore, we expect sodium and chlorine to combine in a 1:1 ratio.