Q9. Which is the correct electron configuration for Fe^{2+}?
MISSED THIS? *Read Section 9.7; Watch IWE 9.6*

a) $[Ar]4s^2 3d^6$ b) $[Ar]4s^2 3d^4$
c) $[Ar]4s^0 3d^6$ d) $[Ar]4s^2 3d^8$

Q10. Which species is diamagnetic?
MISSED THIS? *Read Section 9.7; Watch IWE 9.6*

a) Cr^{2+} b) Zn c) Mn d) C

Q11. Arrange these atoms and ions in order of increasing radius: Cs^+, Ba^{2+}, I^-. **MISSED THIS?** *Read Section 9.7*

a) $I^- < Ba^{2+} < Cs^+$ b) $Cs^+ < Ba^{2+} < I^-$
c) $Ba^{2+} < Cs^+ < I^-$ d) $I^- < Cs^+ < Ba^{2+}$

Q12. Arrange these elements in order of increasing first ionization energy: Cl, Sn, Si.
MISSED THIS? *Read Section 9.7; Watch IWE 9.8*

a) Cl < Si < Sn b) Sn < Si < Cl
c) Si < Cl < Sn d) Sn < Cl < Si

Q13. The ionization energies of an unknown third-period element are listed here. Identify the element. $IE_1 = 786$ kJ/mol; $IE_2 = 1580$ kJ/mol; $IE_3 = 3230$ kJ/mol; $IE_4 = 4360$ kJ/mol; $IE_5 = 16{,}100$ kJ/mol
MISSED THIS? *Read Section 9.7*

a) Mg b) Al c) Si d) P

Q14. Which statement is true about trends in metallic character?
MISSED THIS? *Read Section 9.8*

a) Metallic character *increases* as you move to the right across a row in the periodic table and *increases* as you move down a column.
b) Metallic character *decreases* as you move to the right across a row in the periodic table and *increases* as you move down a column.
c) Metallic character *decreases* as you move to the right across a row in the periodic table and *decreases* as you move down a column.
d) Metallic character *increases* as you move to the right across a row in the periodic table and *decreases* as you move down a column.

Q15. For which element is the gaining of an electron most exothermic? **MISSED THIS?** *Read Section 9.8*

a) Li b) N c) F d) B

Answers: **1.** (b) **2.** (c) **3.** (b) **4.** (d) **5.** (a) **6.** (d) **7.** (d) **8.** (b) **9.** (c) **10.** (b) **11.** (c) **12.** (b) **13.** (c) **14.** (b) **15.** (c)

CHAPTER 9 IN REVIEW

TERMS

Section 9.1
periodic property (352)

Section 9.3
electron configuration (353)
ground state (353)
orbital diagram (354)
Pauli exclusion principle (354)
degenerate (355)
Coulomb's law (355)
shielding (356)
effective nuclear charge (Z_{eff}) (356)
penetration (356)
aufbau principle (358)
Hund's rule (358)

Section 9.4
valence electrons (361)
core electrons (361)

Section 9.6
van der Waals radius (nonbonding atomic radius) (366)
covalent radius (bonding atomic radius) (366)
atomic radius (366)

Section 9.7
paramagnetic (371)
diamagnetic (372)
ionization energy (IE) (375)

Section 9.8
electron affinity (EA) (379)

CONCEPTS

Periodic Properties and the Development of the Periodic Table (9.1, 9.2)

- In the nineteenth century, Dmitri Mendeleev arranged the elements in an early version of the periodic table so that atomic mass increased from left to right in a row and elements with similar properties fell in the same columns.
- Periodic properties are predictable based on an element's position within the periodic table. Periodic properties include atomic radius, ionization energy, electron affinity, density, and metallic character.
- Quantum mechanics explains the periodic table by describing how electrons fill the quantum-mechanical orbitals within the atoms that compose the elements.

Electron Configurations (9.3)

- An electron configuration for an atom shows which quantum-mechanical orbitals the atom's electrons occupy. For example, the electron configuration of helium ($1s^2$) indicates that helium's two electrons occupy the 1*s* orbital.
- The order of filling quantum-mechanical orbitals in multielectron atoms is 1*s* 2*s* 2*p* 3*s* 3*p* 4*s* 3*d* 4*p* 5*s* 4*d* 5*p* 6*s*.
- According to the Pauli exclusion principle, each orbital can hold a maximum of two electrons and those electrons have opposing spins.
- According to Hund's rule, orbitals of the same energy first fill singly with electrons having parallel spins before pairing.

Electron Configurations and the Periodic Table (9.4, 9.5)

- Because quantum-mechanical orbitals fill sequentially with increasing atomic number, we can infer the electron configuration of an element from its position in the periodic table.
- The most stable electron configurations are those with completely full *s* and *p* sublevels. Therefore, the most stable and unreactive elements—those with the lowest energy electron configurations—are the noble gases.
- Elements with one or two valence electrons are among the most active metals, readily losing their valence electrons to attain noble gas configurations.
- Elements with six or seven valence electrons are among the most active nonmetals, readily gaining enough electrons to attain a noble gas configuration.

Effective Nuclear Charge and Periodic Trends in Atomic Size (9.6)

- The size of an atom is largely determined by its outermost electrons. As we move down a column in the periodic table, the principal quantum number (n) of the outermost electrons increases, resulting in successively larger orbitals and therefore larger atomic radii.
- As we move across a row in the periodic table, atomic radii decrease because the effective nuclear charge—the net or average charge experienced by the atom's outermost electrons—increases.
- The atomic radii of the transition elements stay roughly constant as we move across each row because electrons are added to the $n_{\text{highest}} - 1$ orbitals, while the number of highest n electrons stays roughly constant.

Ion Properties (9.7)

- We determine the electron configuration of an ion by adding or subtracting the corresponding number of electrons to the electron configuration of the neutral atom.
- For main-group ions, the order in which electrons are removed is the same as the order in which they are added in building up the electron configuration.
- For transition metal atoms, the *ns* electrons are removed before the $(n - 1)d$ electrons.
- The radius of a cation is much *smaller* than that of the corresponding atom, and the radius of an anion is much *larger* than that of the corresponding atom.
- The first ionization energy—the energy required to remove the first electron from an atom in the gaseous state—generally decreases as we move down a column in the periodic table and increases when we move to the right across a row.
- Successive ionization energies increase smoothly from one valence electron to the next, but the ionization energy increases dramatically for the first core electron.

Electron Affinities and Metallic Character (9.8)

- Electron affinity—the energy associated with an atom in its gaseous state gaining an electron—does not show a general trend as we move down a column in the periodic table, but it generally becomes more negative (more exothermic) to the right across a row.
- Metallic character—the tendency to lose electrons in a chemical reaction—generally increases down a column in the periodic table and decreases to the right across a row.

EQUATIONS AND RELATIONSHIPS

Order of Filling Quantum-Mechanical Orbitals (9.3)

$$1s\ 2s\ 2p\ 3s\ 3p\ 4s\ 3d\ 4p\ 5s\ 4d\ 5p\ 6s$$

LEARNING OUTCOMES

Chapter Objectives	Assessment
Write electron configurations for elements (9.3)	Example 9.1 **For Practice** 9.1 **Exercises** 39–40
Draw orbital diagrams for elements (9.3)	Example 9.2 **For Practice** 9.2 **Exercises** 41–42
Write electron configurations based on periodic table location (9.4)	Examples 9.3, 9.4 **For Practice** 9.3, 9.4 **For More Practice** 9.4 **Exercises** 43–52
Predict relative atomic sizes of elements (9.6)	Example 9.5 **For Practice** 9.5 **For More Practice** 9.5 **Exercises** 53–62
Analyze ions in terms of magnetic properties (9.7)	Example 9.6 **For Practice** 9.6 **Exercises** 63–66
Predict the relative size of ions from periodic trends (9.7)	Example 9.7 **For Practice** 9.7 **Exercises** 67–70
Predict relative ionization energies for atoms and ions based on periodic trends (9.7)	Example 9.8 **For Practice** 9.8 **For More Practice** 9.8 **Exercises** 71–76
Predict the metallic character of atoms based on periodic trends (9.8)	Example 9.9 **For Practice** 9.9 **For More Practice** 9.9 **Exercises** 77–82

EXERCISES

Mastering Chemistry provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

REVIEW QUESTIONS

1. What are periodic properties?

2. Which periodic property is particularly important to nerve signal transmission? Why?

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

4. Who is credited with arranging the periodic table? How are the elements arranged in the modern periodic 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? Give 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 multielectron atoms but not for the hydrogen atom?

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

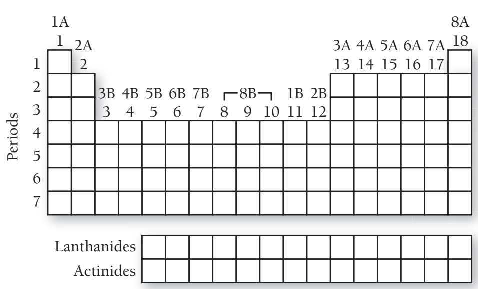
13. Why is electron spin important when writing electron configurations? Explain in terms of the Pauli exclusion principle.

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 multielectron 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. Why do the rows in the periodic table get progressively longer as you move down the table? For example, the first row contains 2 elements, the second and third rows each contain 8 elements, and the fourth and fifth rows each contain 18 elements. Explain.

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

21. Describe 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?

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

24. Describe 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, 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 you move:

a. across a period in the periodic table
b. down a column in the periodic table

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

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

29. For transition elements, describe and explain the observed trends in atomic radius as you move:

a. across a period in the periodic table
b. down a column in the periodic table

30. 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?

31. Describe 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?

32. Describe the relationship between

a. the radius of a cation and that of the atom from which forms
b. the radius of an anion and that of the atom from which it forms

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

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

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

36. 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 you might predict them.

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

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

PROBLEMS BY TOPIC

Electron Configurations

39. Write the full electron configuration for each element.
MISSED THIS? *Read Section 9.3; Watch KCV 9.3*
a. Si b. O c. K d. Ne

40. Write the full electron configuration for each element.
a. C b. P **c.** Ar d. Na

41. Write the full orbital diagram for each element.
MISSED THIS? *Read Section 9.3; Watch KCV 9.3, IWE 9.2*
a. N b. F c. Mg d. Al

42. Write the full orbital diagram for each element.
a. S b. Ca c. Ne d. He

43. Use the periodic table to write an electron configuration for each element. Represent core electrons with the symbol of the previous noble gas in brackets.
MISSED THIS? *Read Section 9.4; Watch KCV 9.4, IWE 9.4*
a. P b. Ge c. Zr d. I

44. Use the periodic table to determine the element corresponding to each electron configuration.
a. $[Ar]\,4s^23d^{10}4p^6$ b. $[Ar]\,4s^23d^2$
c. $[Kr]\,5s^24d^{10}5p^2$ d. $[Kr]\,5s^2$

45. Use the periodic table to determine each quantity.
MISSED THIS? *Read Section 9.4; Watch KCV 9.4, IWE 9.4*
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

46. Use the periodic table to determine each quantity.
a. the number of 3*s* electrons in Mg
b. the number of 3*d* electrons in Cr
c. the number of 4*d* electrons in Y
d. the number of 6*p* electrons in Pb

47. Name an element in the fourth period (row) of the periodic table with the following:
MISSED THIS? *Read Section 9.4; Watch KCV 9.4, IWE 9.4*
a. five valence electrons
b. four 4*p* electrons
c. three 3*d* electrons
d. full *s* and *p* sublevels

48. Name an element in the third period (row) of the periodic table with the following:
a. three valence electrons
b. four 3*p* electrons
c. six 3*p* electrons
d. two 3*s* electrons and no 3*p* electrons

Valence Electrons and Simple Chemical Behavior from the Periodic Table

49. Determine the number of valence electrons in an atom of each element. **MISSED THIS?** *Read Section 9.4; Watch KCV 9.4*
a. Ba b. Cs c. Ni d. S

50. Determine the number of valence electrons in an atom of each element. Which elements do you expect to lose electrons in their chemical reactions? Which do you expect to gain electrons?
a. Al b. Sn c. Br d. Se

51. Which outer electron configuration would you expect to belong to a reactive metal? To a reactive nonmetal?
MISSED THIS? *Read Section 9.5*
a. ns^2 b. ns^2np^6 c. ns^2np^5 d. ns^2np^2

52. Which outer electron configurations would you expect to belong to a noble gas? To a metalloid?
a. ns^2 b. ns^2np^6 c. ns^2np^5 d. ns^2np^2

Coulomb's Law and Effective Nuclear Charge

53. According to Coulomb's law, which pair of charged particles has the lowest potential energy? **MISSED THIS?** *Read Section 9.3*
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

54. 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

55. Which experience a greater effective nuclear charge: the valence electrons in beryllium or the valence electrons in nitrogen? Why?
MISSED THIS? *Read Section 9.6; Watch KCV 9.6*

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

57. If core electrons completely shielded valence electrons from nuclear charge (i.e., if each core electron reduced nuclear charge by 1 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?
MISSED THIS? *Read Section 9.6; Watch KCV 9.6*
a. K b. Ca c. O d. C

58. In Section 9.6, we estimated the effective nuclear charge on beryllium's valence electrons to be slightly greater than 2+. What would a similar process 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

59. Choose the larger atom from each pair.
MISSED THIS? *Read Section 9.6; Watch KCV 9.6, IWE 9.5*
a. Al or In b. Si or N
c. P or Pb d. C or F

60. Choose the larger atom from each pair, if possible.
a. Sn or Si b. Br or Ga
c. Sn or Bi d. Se or Sn

61. Arrange these elements in order of increasing atomic radius: Ca, Rb, S, Si, Ge, F.
MISSED THIS? *Read Section 9.6; Watch KCV 9.6, IWE 9.5*

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

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

63. Write the electron configuration for each ion.
MISSED THIS? *Read Section 9.7; Watch IWE 9.6*
a. O^{2-} b. Br^- c. Sr^{2+} d. Co^{3+} e. Cu^{2+}

64. Write the electron configuration for each ion.
a. Cl^- b. P^{3-} c. K^+ d. Mo^{3+} e. V^{3+}

65. Write orbital diagrams for each ion and indicate whether the ion is diamagnetic or paramagnetic.
MISSED THIS? *Read Section 9.7; Watch IWE 9.6*
a. V^{5+} b. Cr^{3+} c. Ni^{2+} d. Fe^{3+}

66. Write orbital diagrams for each ion and indicate whether the ion is diamagnetic or paramagnetic.
a. Cd^{2+} b. Au^+ c. Mo^{3+} d. Zr^{2+}

67. Which is the larger species in each pair?
MISSED THIS? *Read Section 9.7*
a. Li or Li^+ b. I^- or Cs^+
c. Cr or Cr^{3+} d. O or O^{2-}

68. 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+}

69. Arrange this isoelectronic series in order of decreasing radius: F^-, O^{2-}, Mg^{2+}, Na^+. **MISSED THIS?** *Read Section 9.7*

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

71. Choose the element with the higher first ionization energy from each pair. **MISSED THIS?** *Read Section 9.7; Watch IWE 9.8*
a. Br or Bi b. Na or Rb
c. As or At d. P or Sn

72. Choose the element with the higher first ionization energy from each pair.
a. P or I b. Si or Cl
c. P or Sb d. Ga or Ge

73. Arrange these elements in order of increasing first ionization energy: Si, F, In, N. **MISSED THIS?** *Read Section 9.7; Watch IWE 9.8*

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

75. 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?) **MISSED THIS?** *Read Section 9.7*
a. Be b. N c. O d. Li

76. Consider this set of ionization energies.

$$IE_1 = 578 \text{ kJ/mol}$$
$$IE_2 = 1820 \text{ kJ/mol}$$
$$IE_3 = 2750 \text{ kJ/mol}$$
$$IE_4 = 11{,}600 \text{ kJ/mol}$$

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

Electron Affinities and Metallic Character

77. Choose the element with the more negative (more exothermic) electron affinity from each pair. **MISSED THIS?** *Read Section 9.8*
a. Na or Rb b. B or S
c. C or N d. Li or F

78. Choose the element with the more negative (more exothermic) electron affinity from each pair.
a. Mg or S b. K or Cs
c. Si or P d. Ga or Br

79. Choose the more metallic element from each pair.
MISSED THIS? *Read Section 9.8*
a. Sr or Sb b. As or Bi c. Cl or O d. S or As

80. Choose the more metallic element from each pair.
a. Sb or Pb b. K or Ge c. Ge or Sb d. As or Sn

81. Arrange these elements in order of increasing metallic character: Fr, Sb, In, S, Ba, Se. **MISSED THIS?** *Read Section 9.8*

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

CUMULATIVE PROBLEMS

83. Bromine is a highly reactive liquid while krypton is an inert gas. Explain this difference based on their electron configurations.

84. Potassium is a highly reactive metal while argon is an inert gas. Explain this difference based on their electron configurations.

85. Both vanadium and its 3+ ion are paramagnetic. Refer to their electron configurations to explain this statement.

86. Refer to their electron configurations to explain why copper is paramagnetic while its 1+ ion is not.

87. Suppose you were trying to find a substitute for K^+ in nerve signal transmission. Where would you begin your search? What ions would be most like K^+? For each ion you propose, explain the ways in which it would be similar to K^+ and the ways it would be different. Refer to periodic trends in your discussion.

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

89. 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?

90. 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

91. 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.

92. 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.

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

94. 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?

95. 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.

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

97. Write general outer electron configurations (ns^xnp^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 why this is so.

98. 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 observation.

99. 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 share these chemical properties.

100. Write the electron 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.

101. 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. The message 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

102. 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.

103. 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 has sufficient energy to ionize the atom?

104. 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 3*s* electron. How does this distance compare to the atomic radius of sodium? Explain the difference.

105. Consider the elements: B, C, N, O, F.

a. Which element has the highest first ionization energy?
b. Which element has the largest atomic radius?
c. Which element is most metallic?
d. Which element has three unpaired electrons?

106. Consider the elements: Na, Mg, Al, Si, P.

a. Which element has the highest second ionization energy?
b. Which element has the smallest atomic radius?
c. Which element is least metallic?
d. Which element is diamagnetic?

CHALLENGE PROBLEMS

107. 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. Estimate the density of the 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 present 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?

108. 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?

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

110. Imagine a universe in which the value of m_s can be $+\frac{1}{2}$, 0, and $-\frac{1}{2}$. 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

111. A carbon atom can absorb radiation of various wavelengths with resulting changes in its electron configuration. Write orbital diagrams for the electron configuration of carbon that results from absorption of the three longest wavelengths of radiation it can absorb.

112. 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.

113. 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.

114. 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.

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

116. Using the data in Figures 9.15 and 9.16, calculate ΔE for the reaction $Na(g) + Cl(g) \longrightarrow Na^{+}(g) + Cl^{-}(g)$.

117. 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.

118. In Section 2.7 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.

119. 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.

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

121. Table 9.2 does not include francium because none of francium's isotopes are stable. Predict the values of the entries for Fr in Table 9.2. Predict the nature of the products of the reaction of Fr with: (a) water, (b) oxygen, and (c) chlorine.

122. From its electronic configuration, predict which of the first 10 elements would be most similar in chemical behavior to the as yet undiscovered element 165.

CONCEPTUAL PROBLEMS

123. 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 each of the following:

a. a noble gas
b. a reactive nonmetal
c. a reactive metal

124. 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.

125. 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 penetrates into the region occupied by core electrons.
c. An electron in an orbital that penetrates closer to the nucleus always experiences 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 tends to experience a higher effective nuclear charge than an electron in an orbital that does not penetrate close to the nucleus.

126. 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.

127. 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.

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

129. 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 9.5 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."

130. 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.

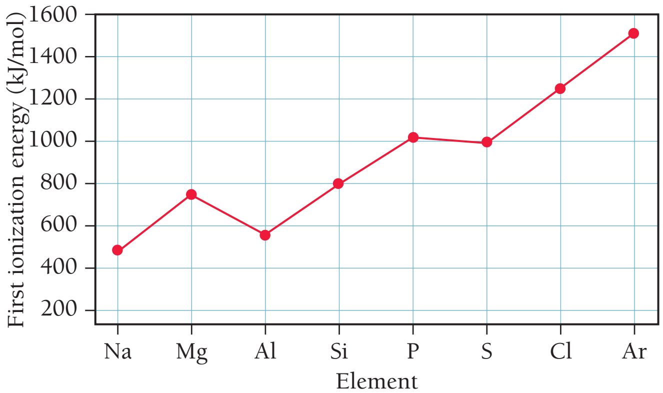
131. In complete sentences, explain: (a) why Se^{2-} and Br^{-} are about the same size; (b) why Br^{-} is slightly smaller than Se^{2-}; and (c) which singly charged cation you would expect to be approximately the same size as Se^{2-} and Br^{-} and why.

132. 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.

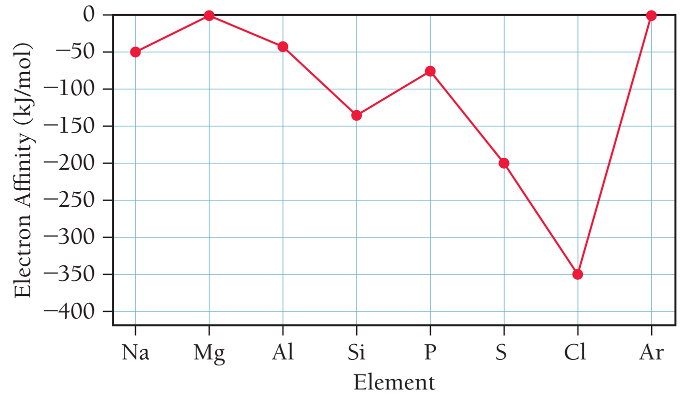
DATA INTERPRETATION AND ANALYSIS

Periodic Properties of Period 3 Elements

133. The accompanying 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 exceptions: one at Al and another at 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) that 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?

Cc ANSWERS TO CONCEPTUAL CONNECTIONS

Coulomb's Law

9.1 **(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 (becomes more negative).

Penetration and Shielding

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

Electron Configurations and Quantum Numbers

9.3 **(b)** $n = 4,\ l = 0,\ m_l = 0,\ m_s = +\frac{1}{2};\ n = 4,\ l = 0,\ m_1 = 0,\ m_s = -\frac{1}{2}$

Valence Electrons and Group Number

9.4 **(c)** Nitrogen has five valence electrons. Since nitrogen is a main-group element, it has the same number of valence electrons as its lettered group number in the periodic table.

Electron Configuration and Ion Charge

9.5 **(b)** Elements with electron configurations *close* to those of the noble gases gain or lose electrons to attain a noble gas configuration. The 2− charge implies that the element gained two electrons, which results in the configuration ns^2np^6, which is a noble gas configuration.

Effective Nuclear Charge

9.6 **(c)** Since 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.

Ions, Isotopes, and Atomic Size

9.7 **(b)** 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.

Successive Ionization Energies

9.8 **(c)** Since B has three valence electrons, it would have a huge jump between its third and fourth ionization energies. The third ionization energy corresponds to removing the third valence electron, while the fourth ionization energy corresponds to removing the first core electron.

Ionization Energies and Chemical Bonding

9.9 **(b)** 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, however, are held tightly and are not easily transferred or shared. Consequently, valence electrons are most important to chemical bonding.

Periodic Trends

9.10 **(c)** The 3s 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 (2p).