Solution Manual for Introductory Chemistry An Atoms First Approach 1st Edition Burdge Driessen 0073402702 9780073402703

Chapter 2 Electrons and the Periodic Table

Practice Problems C

2.1 (a) ultraviolet, (b) infrared.

2.2 Red: 1*s*, blue: 2*p*, yellow: 3*d*, pink: 3*p*, green: 4*f*, purple: 4*s*.

2.3 (a) should be $[Ar]4s^23d^8$, Ni; (b) s/b $[Ar]4s^23d^{10}4p^5$, Br.

2.4 (a) Sn, (b) As.

2.5 No, because we can write configurations for transition elements with the *d* electrons last, even though they have a lower principal quantum number than the *s* electrons.

2.6 (a) S < P, but trends alone cannot determine the size of Br relative to P or to S. (b) I < Rb < Cs

2.7 No. The trends allow us to determine that F is more metallic than Ne and that Ar is more metallic than Ne, but do not allow us to rank metallic character of F relative to Ar.

2.8 No, because of competing trends. Difficulty of electron removal increases across a period—but decreases down a group.

2.9 Noble gases have completed subshells and it is especially difficult to remove electrons from them.

2.10 No, because isoelectronic ions of different elements have identical electron configurations.

2.11 (a) 1, (b) 2

Key Skills

2.1 b, 2.2 d, 2.3 c, 2.4 d, 2.5 b, 2.6 b.

Questions and Problems

2.1. They are inversely proportional—the longer the wavelength, the shorter the frequency.

2.2. They are directly proportional—the higher the frequency, the higher the energy. $E = h\nu$, where h = Planck's constant (6.626×10^{-34} Js) and ν is the frequency.

2.3. They are inversely proportional—as energy increases, wavelength decreases. $E = \frac{hc}{\lambda}$, where h = Planck's constant (6.626×10^{-34} Js), c = speed of light (3.00×10^{8} m/s), and λ = wavelength in meters.

2.4. Refer to the rainbow (the visible spectrum) to see the arrangement of visible light by color and wavelength. Some students use ROY G BIV (red, orange, yellow,…) to remember the order of light by color. Yellow has the longest wavelength of the three colors given.

2.5. Refer to the rainbow (the visible spectrum) to see the arrangement of visible light by color and wavelength. Some students use ROY G BIV (red, orange, yellow,…) to remember the order of light by color. Red has the longest wavelength of the three colors given.

2.6. Refer to the rainbow (the visible spectrum) to see the arrangement of visible light by color and wavelength. Then remember that wavelength and frequency are inversely proportional. Some students use ROY G BIV (red, orange, yellow,…) to remember the order of light by color and wavelength. Violet has the largest frequency (shortest wavelength) of the three colors given.

2.7. Refer to the rainbow (the visible spectrum) to see the arrangement of visible light by color and wavelength. Then remember that wavelength and frequency are inversely proportional. Some students use ROY G BIV (red, orange, yellow,…) to remember the order of light by color and wavelength. Blue has the largest frequency (shortest wavelength) of the three colors given.

2.8. Remember that frequency and wavelength are inversely proportional.
550 nm < 450 nm < 350 nm

2.9. Remember that frequency and wavelength are inversely proportional.
400 nm > 550 nm > 700 nm

2.10. radio < infrared < X rays

2.11. microwave < visible < gamma

2.12. red < yellow < violet

2.13. blue > green > orange

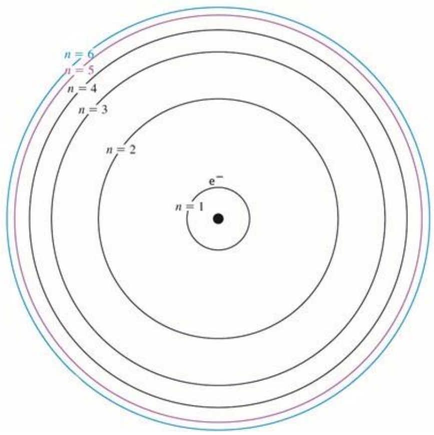
2.14. Only certain quantities are allowed.

2.15. A "packet" or particle of light energy.

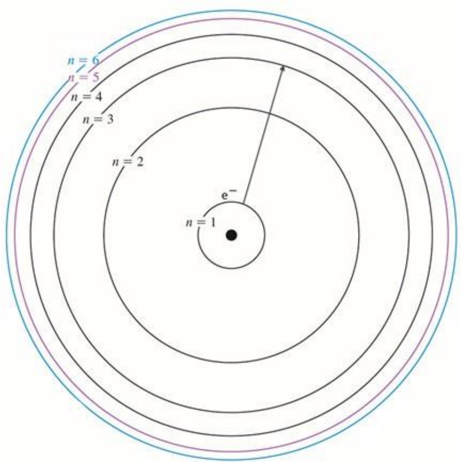
2.16. All of the electrons in the atom are in the lowest possible energy levels.

2.17. The atom has absorbed energy and at least one electron has moved to a higher energy level than in the ground state.

2.18.



a.



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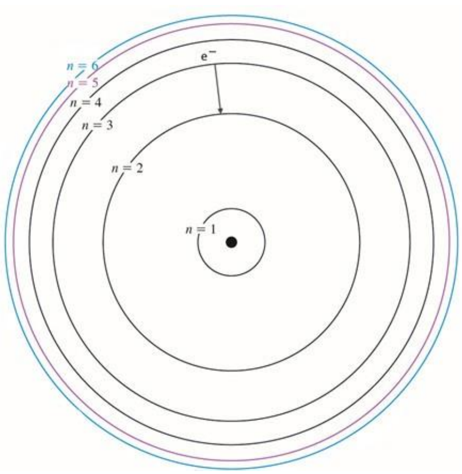


c.

2.19. (b). Remember that visible light emission from a hydrogen atom starts at either n = 6, n = 5, n = 4, or n = 3 and ends on n = 2.

2.20. (c). Remember that visible light emission from a hydrogen atom starts at either n = 6, n = 5, n = 4, or n = 3 and ends on n = 2.

2.21. (a). Remember that emission occurs when the starting n value is greater than the final n value.

2.22. (b) and (c). Remember that emission occurs when the starting n value is greater than the final n value.

2.23. (a) and (b). Remember that absorption occurs when the starting n value is less than the final n value.

2.24. (a) and (c). Remember that absorption occurs when the starting n value is less than the final n value.

2.25. The larger the difference between the n values, the greater the energy difference between them. Since wavelength is inversely proportional to energy ($E = hc/\lambda$), the smaller the wavelength, the larger the energy difference.
410 nm matches the n = 6 to n = 2 transition
434 nm matches the n = 5 to n = 2 transition
486 nm matches the n = 4 to n = 2 transition
657 nm matches the n = 3 to n = 2 transition

2.26. The number of photons emitted is equivalent to the number of atoms undergoing the transition.
a. one photon, b. one photon, c. twelve photons, d. fifty photons

VC 2.1 b

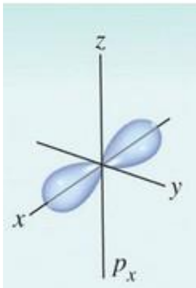
VC 2.2 a

VC 2.3 c

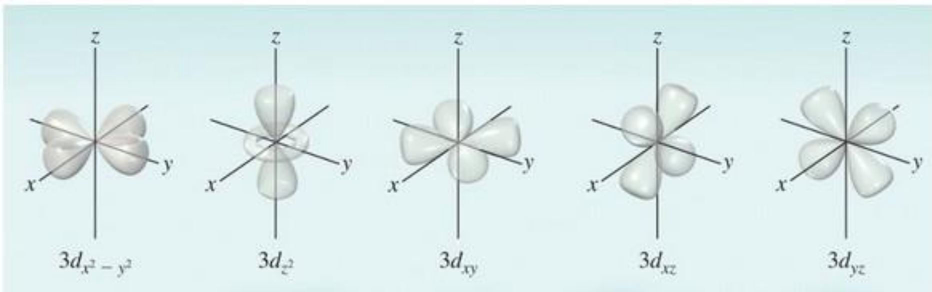
VC 2.4 c

2.27. The volume where an electron is most likely to be found.

2.28. Sublevels can contain one or more orbitals, depending on the type.

2.29. 3s =  , 3p = 

,

3d = 

2.30. The 2s orbital is the same shape as the 3s orbital, but smaller in size/volume.

2.31. The 4p orbitals are larger, but have the same shape.

2.32. It is larger, but has the same shape.

2.33. When comparing the same type of orbital, the one with the higher n value is larger.
a. 4s, b. they are equal in size, c. 4p

2.34. When comparing the same type of orbital, the one with the higher n value is larger.
a. 5s, b. $4p_y$, c. 4p

2.35. When comparing the same type of orbital, the one with the lower n value is smaller.
a. 3d, b. 1s, c. $2p_x$

2.36. When comparing the same type of orbital, the one with the lower n value is smaller.
a. 3s, b. 2p, c. $3d_y$

2.37. When comparing the same type of orbital, the one with the lower n value is smaller. This means the electrons in that orbital will be closer to the nucleus, on average.
a. 1s, b. 2p, c. 3d

2.38. When comparing orbitals, the one with the lower n value or the one that fills first is smaller. This means the electrons in that orbital will be closer to the nucleus, on average.
a. 2s, b. 3p, c. 4s

2.39. a. Yes, the fifth shell (level) contains p orbitals.
b. Yes, the fourth shell (level) contains an s orbital.
c. No, there are no f orbitals in the second shell (level).
d. No, there are no p orbitals in the first shell (level).

2.40. a. No, there are no p orbitals in the first shell (level).
b. Yes, there is an s orbital in the sixth shell (level).
c. No, there are no f orbitals in the third shell (level).
d. Yes, there are p orbitals in the fourth shell (level).

2.41. The s subshell/sublevel contains only one orbital, while others contain several.
a. sublevel, b. orbital and sublevel, c. single orbital, d. sublevel

2.42. The s subshell/sublevel contains only one orbital, while others contain several.
a. sublevel, b. orbital and sublevel, c. single orbital, d. sublevel

2.43. 1s = spherical orbital in the first level
2p = dumbbell-shaped orbital in the second level
4s = spherical orbital in the fourth level
3d = cloverleaf-shaped orbital in the third level

2.44. The lower energy orbitals are the ones that fill first when writing electron configurations.
a. 3d > 3p > 3s
b. 3s > 2s > 1s
c. 3s > 2p > 2s

2.45. The lower energy orbitals are the ones that fill first when writing electron configurations.

a. 4d > 4p > 4s

b. 4p > 3p > 2p

c. 3d > 2p > 1s

2.46. The ground state means that the electrons are filled in the lowest energy orbitals/sublevels possible. This is the filling order followed for writing electron configurations.

2.47. 2 electrons

2.48. Any orbital can contain two electrons.

a. 2, b. 2, c. 2, d. 2

2.49. a. 1, b. 5, c. 3, d. 7

2.50. Remember how many orbitals each type of subshell contains. This number must be multiplied by two since each orbital can contain two electrons.

a. 2, b. 6, c. 10, d. 14

2.51. Remember how many orbitals each type of subshell contains. This number must be multiplied by two since each orbital can contain two electrons.

a. 6, b. 10, c. 2, d. 6

2.52. a. $1s^2 2s^2 2p^6 3s^2 3p^5$

b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

c. $1s^2 2s^2 2p^3$

2.53. a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

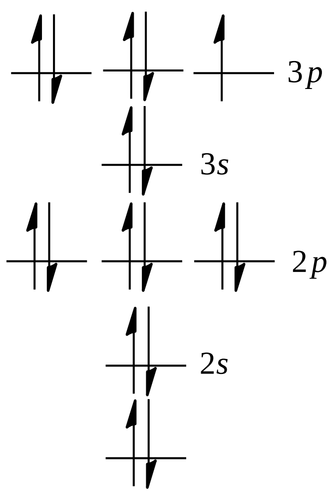
b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$

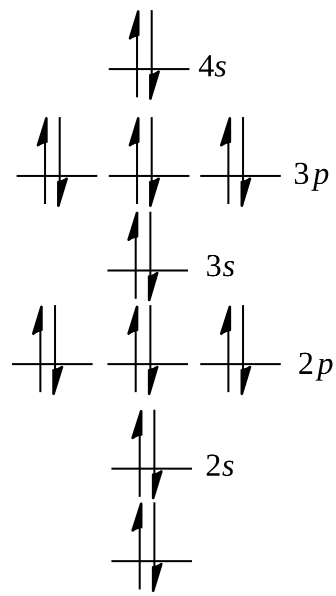
2.54. a. $1s^2 2s^2 2p^4$
b. $1s^2 2s^2 2p^6 3s^2 3p^4$
c. $1s^2 2s^2 2p^6 3s^1$

2.55. a. $1s^2 2s^1$
b. $1s^2 2s^2 2p^6 3s^2 3p^2$
c. $1s^2 2s^2 2p^6 3s^2$

2.56. Write out the electron configuration and use it to fill in the orbital diagram. Remember that within a sublevel, the electrons do not pair up until all of the orbitals have one electron.



a.



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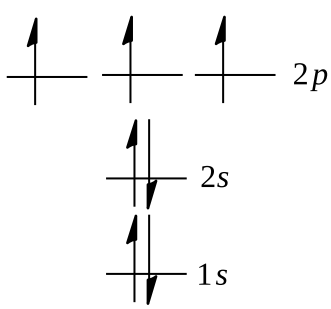
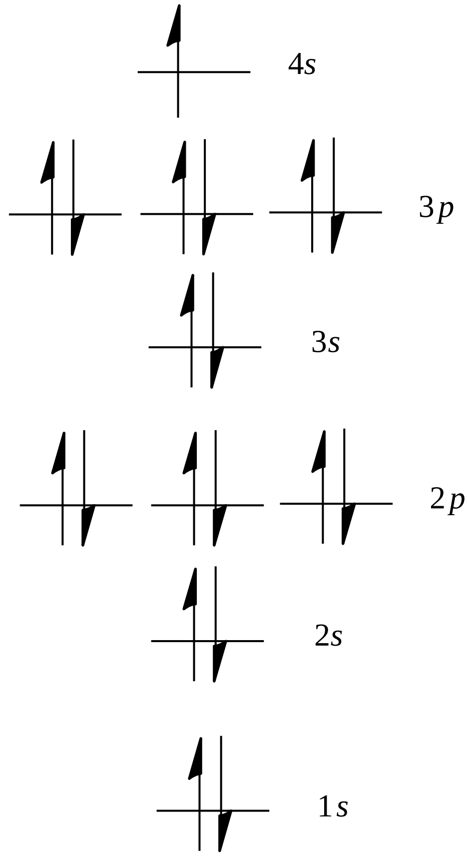
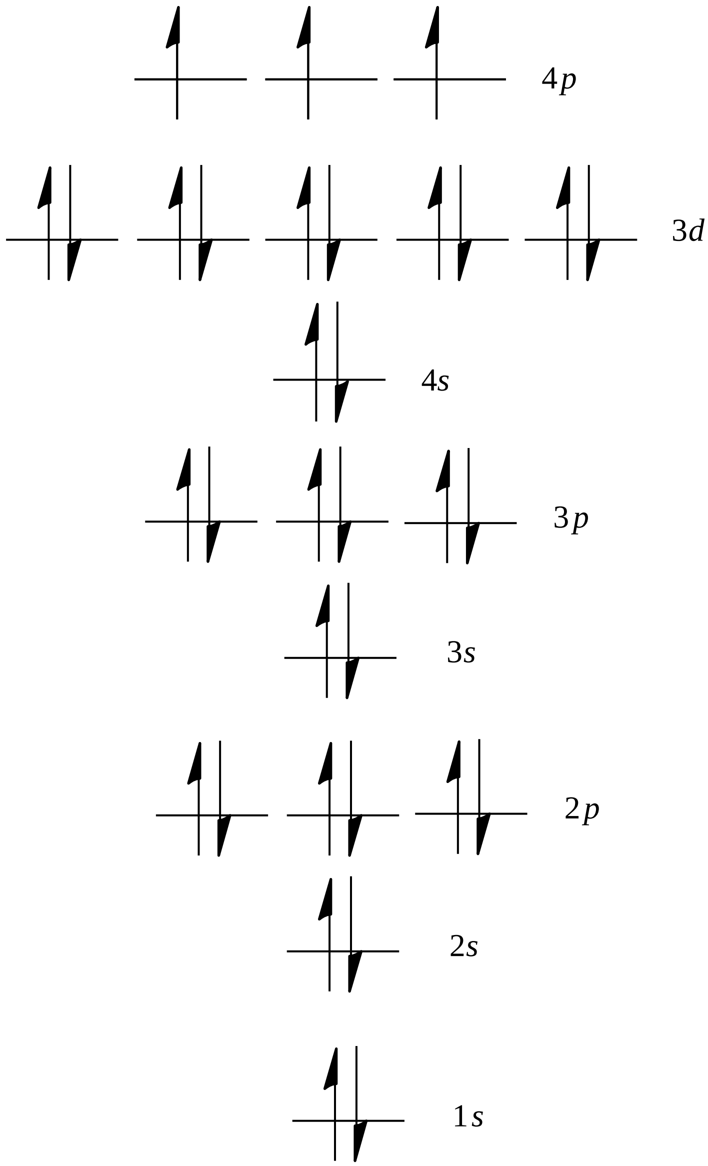


c.

2.57. Write out the electron configuration and use it to fill in the orbital diagram. Remember that within a sublevel, the electrons do not pair up until all of the orbitals have one electron.



a.



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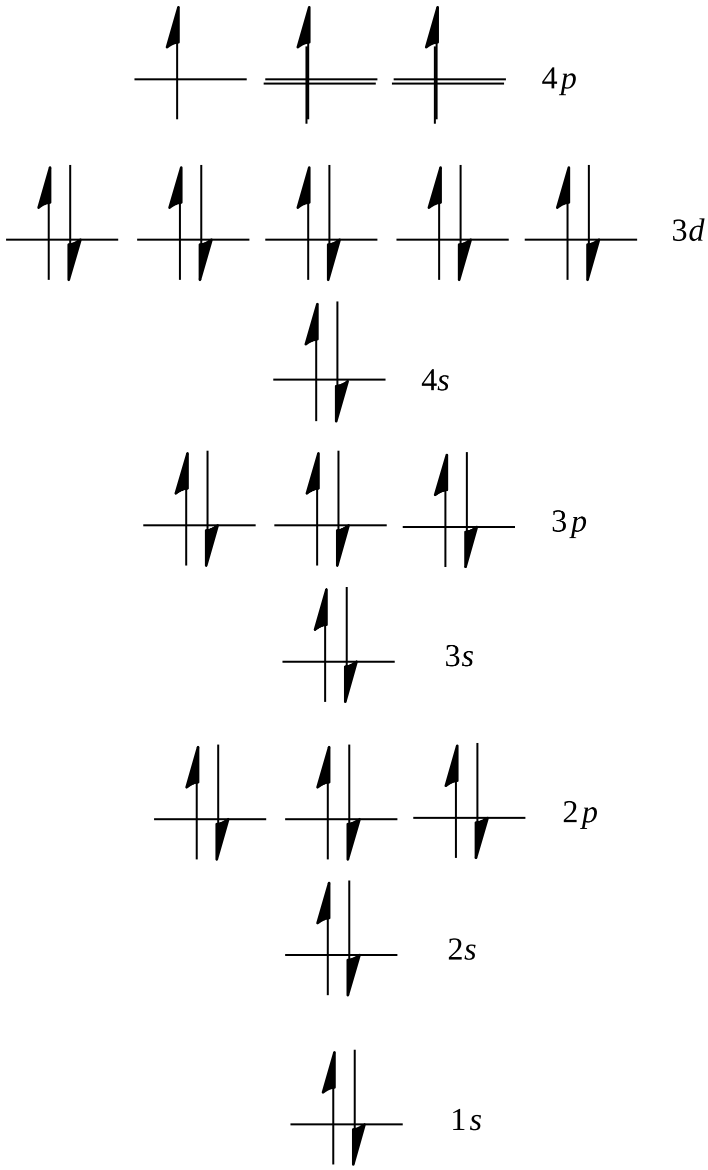
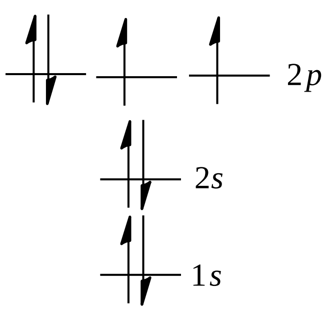
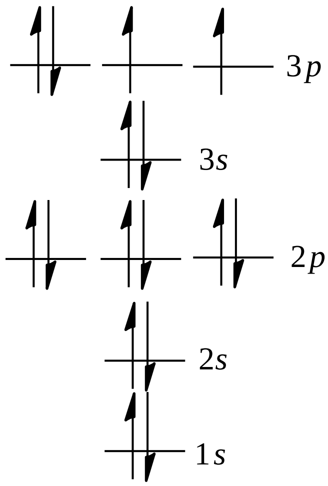


c.

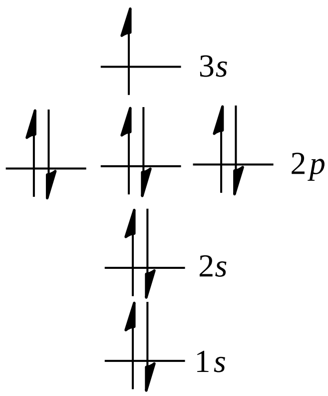
2.58. Write out the electron configuration and use it to fill in the orbital diagram. Remember that within a sublevel, the electrons do not pair up until all of the orbitals have one electron.



a.

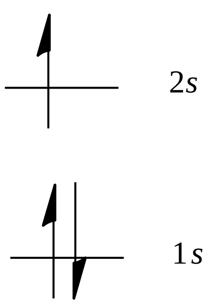


b.

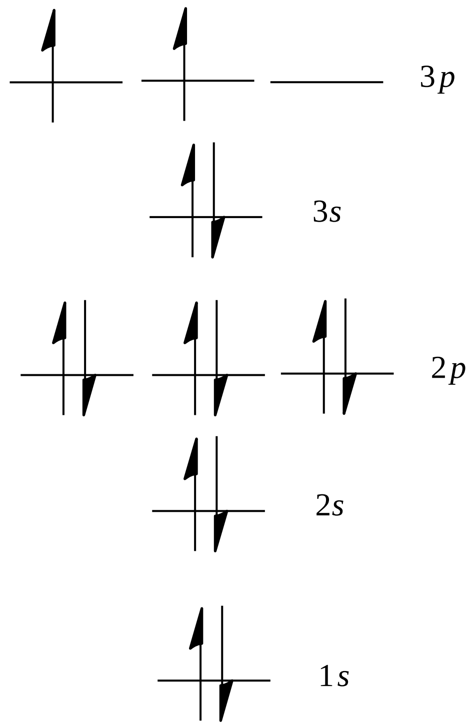


c.

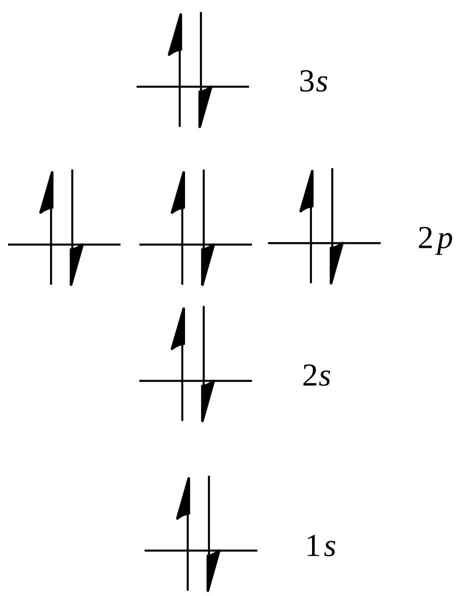
2.59. Write out the electron configuration and use it to fill in the orbital diagram. Remember that within a sublevel, the electrons do not pair up until all of the orbitals have one electron.



a.



b.



c.

2.60. Use the noble gas core notation (the noble gas symbol in square brackets) to represent the core electrons.

a. $[Ar]4s^2 3d^{10} 4p^5$

b. $[Kr]5s^2 4d^{10} 5p^4$

c. $[Xe]6s^1$

2.61. Use the noble gas core notation (the noble gas symbol in square brackets) to represent the core electrons.

a. $[Ar]4s^2 3d^6$

b. $[Ar]4s^2 3d^{10}$

c. $[Ar]4s^2 3d^8$

2.62. Use the noble gas core notation (the noble gas symbol in square brackets) to represent the core electrons.

a. $[Kr]5s^2 4d^2$

b. $[Ar]4s^2 3d^7$

c. $[Ar]4s^2 4d^5$

2.63. Use the noble gas core notation (the noble gas symbol in square brackets) to represent the core electrons.

a. $[Kr]5s^2 4d^{10}$

b. $[Kr]5s^2 4d^8$

c. $[Ar]4s^2 3d^3$

2.64. Valence electrons are those in the highest n level or shell. Core electrons are all of the other electrons in filled shells of the atom.

2.65. Remember that valence electrons are those in the highest n level or shell. Core electrons are all of the other electrons in filled shells of the atom.

a. 10 core electrons, 5 valence electrons

b. 46 core electrons, 7 valence electrons

c. 18 core electrons, 2 valence electrons

d. 18 core electrons, 1 valence electron

2.66. Remember that valence electrons are those in the highest n level or shell. Core electrons are all of the other electrons in filled shells of the atom.

a. 10 core electrons, 7 valence electrons

b. 2 core electrons, 5 valence electrons

c. 54 core electrons, 1 valence electron

d. 28 core electrons, 5 valence electrons

2.67. I = s-block, II = p-block, III = d-block

2.68. Remember that valence electrons are those in the highest n level or shell. Core electrons are all of the other electrons in filled shells of the atom.

a. $[Kr]5s^2$, 2 valence electrons

b. $[Ar]4s^2 3d^{10} 4p^5$, 7 valence electrons

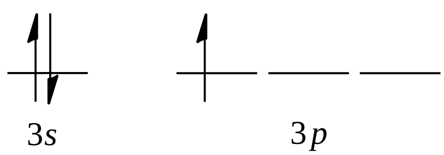
c. $[Kr]5s^2 4d^{10} 5p^6$, 8 valence electrons

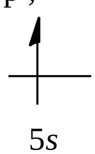
2.69. Remember that valence electrons are those in the highest n level or shell. Core electrons are all of the other electrons in filled shells of the atom.

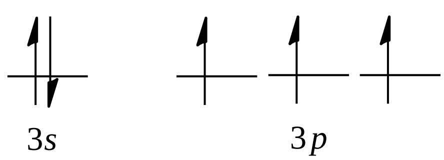
a. $[Kr]5s^2 4d^{10} 5p^3$, 5 valence electrons

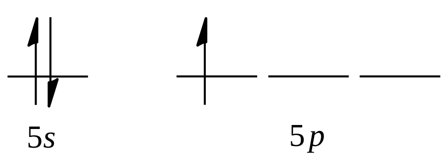
b. $[Xe]6s^2$, 2 valence electrons

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

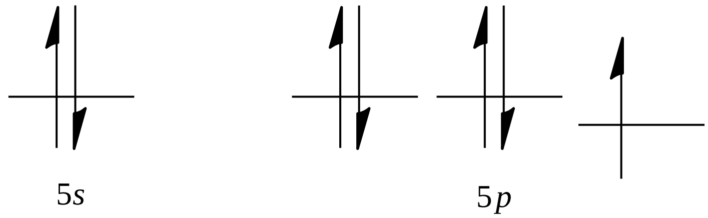
2.70. a. $[Ne]3s^2 3p^1$, 

b. $[Kr]5s^1$, 

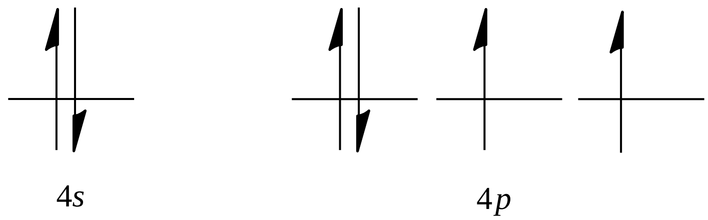
c. $[Ne]3s^2 3p^3$, 

d. $[Kr]5s^2 4d^{10} 5p^1$, 

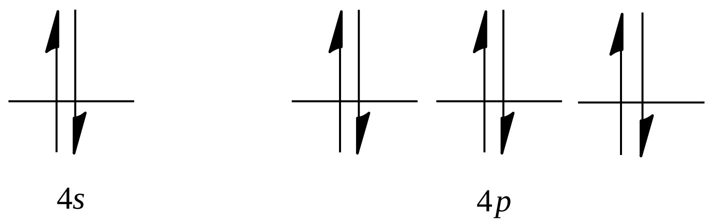
2.71. a. [Kr]$5s^2 4d^{10} 5p^5$,



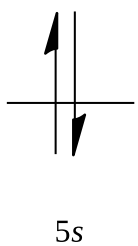
b. [Ar]$4s^2 3d^{10} 4p^4$,



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



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



2.72. a. 1 , b. 1, c. 3, d. 1

2.73. a. 1, b. 2, c. 0, d. 0

2.74. a. [He]$2s^1$, 1 valence electron
b. [Ne]$3s^1$, 1 valence electron
c. [Ar]$4s^1$, 1 valence electron
Each of these elements would form a 1+ ion when they lose their valence electron. They are all located in group 1A (1), so the charge can be predicted from their location on the periodic table.

2.75. a. [He]$2s^2 2p^5$, 7 valence electrons
b. [Ne]$3s^2 3p^5$, 7 valence electrons
c. [Ar]$4s^2 3d^{10} 4p^5$, 7 valence electrons

All three of these elements would be predicted to form 1– ions by gaining one electron to fill their valence shell. They are all found in group 7A (17), so the charge can be predicted from their location on the periodic table.

2.76. Count the total number of electrons present and use this to determine the atomic number and identity of the element from the periodic table.

a. nickel, Ni

b. phosphorus, P

c. selenium, Se

2.77. Count the total number of electrons present and use this to determine the atomic number and identity of the element from the periodic table.

a. tin, Sn

b. cesium, Cs

c. bromine, Br

2.78. Count the number of electrons and use this to determine the atomic number and element identity using the periodic table.

a. carbon, C

b. sodium, Na

c. phosphorus, P

2.79. a. This element contains 14 electrons (add up the superscripts), so it is Si. $1s^22s^22p^63s^23p^2$

b. This element contains 6 electrons and can be identified as C. $1s^22s^22p^2$

c. This element contains 10 electrons and can be identified as Ne. $1s^22s^22p^6$

2.80. a. This element contains 20 electrons (add up the superscripts), so it is Ca. $1s^22s^22p^63s^23p^64s^2$

b. This element contains 19 electrons and can be identified as K. $1s^22s^22p^63s^23p^64s^1$

c. This element contains 26 electrons and can be identified as Fe. $1s^22s^22p^63s^23p^64s^23d^6$

2.81. Br, because it has the same number of valence electrons and is located in the same group on the periodic table.

2.82. Se, because it has the same number of valence electrons and is located in the same group on the periodic table.

2.83. a. Mg

b. P

c. F

d. Ar

2.84. a. C

b. Cl

c. Ne

d. K

2.85. a. N

b. Br

c. Ca

d. Li

2.86. a. O

b. S

c. Se

The all have the same number of valence electrons (dots) and are located in the same family/group on the periodic table.

2.87. a. N

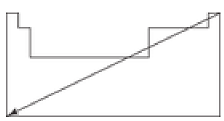
b. P

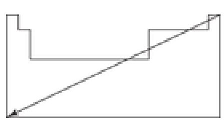
c. As

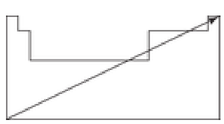
They have the same number of valence electrons (dots). They are in the same group on the periodic table.

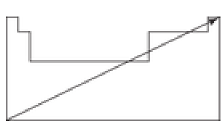
2.88. It costs less energy when an electron is easier to remove than if it is more difficult. Atoms lose electrons more easily when the outer electrons are further from the nucleus, as they are when atoms are larger.

2.89. Nonmetals gain electrons most easily.

2.90. 

2.91. 

2.92. 

2.93. 

2.94. Atomic size increases from the upper right corner of the periodic table to the lower left.

a. P > S > Cl

b. Se > S > O

c. K > Ca > Kr

2.95. Atomic size increases from the upper right corner of the periodic table to the lower left.

a. S < Sr < Rb

b. Li < Mg < Ca

c. Br < Ca < K

2.96. Atomic size increases from the upper right corner of the periodic table to the lower left, which would make these elements all appear to be about the same size. Without more information, it is not possible to put these elements in order of increasing size.

2.97. Electrons are most difficult to remove from the nonmetals in the upper right corner of the periodic table and get easier to remove as you move toward the lower left corner.

a. S

b. Si

c. K

2.98. Electrons are most difficult to remove from the nonmetals in the upper right corner of the periodic table and get easier to remove as you move toward the lower left corner.

a. Se

b. S

c. O

2.99. Metallic character increases from the upper right of the periodic table to the lower left corner.

a. S

b. F

c. P

2.100. Metallic character increases from the upper right of the periodic table to the lower left corner.

a. Cs

b. Na

c. K

2.101. An ion is an atom that has lost or gained one or more electrons, leaving it with either a positive or negative charge. Atoms are neutral and become ions when they gain or lose electron(s).

2.102. A cation is an atom that has lost one or more electrons. It has a net positive charge.

2.103. An anion is an atom that has gained one or more electrons. It has a net negative charge.

2.104. Use the periodic table to determine the number of protons in the element and then use the number of electrons given to find the difference between the two values. If there are more electrons than protons, the ion will have a negative charge. If the reverse is true, the ion will have a positive charge.

a. 1–

b. 3+

c. 2+

d. 3–

2.105. Use the location of the element on the periodic table to determine if it will lose or gain electrons, and how many, to have the same number of electrons as the nearest noble gas. If it gains electrons, it will be negative; if it loses electrons, it will be positive.

a. Mg^{2+}

b. K^{+}

c. P^{3-}

d. O^{2-}

e. I^{-}

2.106. Use the location of the element on the periodic table to determine if it will lose or gain electrons, and how many, to have the same number of electrons as the nearest noble gas. If it gains electrons it will be negative, if it loses electrons it will be positive.

a. Ba^{2+}

b. N^{3-}

c. S^{2-}

d. Al^{3+}

e. Te^{2-}

2.107. Determine the electron configuration for the atom first. Then decide if it gains or loses electrons when it forms its ion. If it gains electrons, add them to the electron configuration and vice versa.

a. $[He]2s^2 2p^6$

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

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

2.108. Determine the electron configuration for the atom first. Then decide if it gains or loses electrons when it forms its ion. If it gains electrons, add them to the electron configuration and vice versa.

a. $[Ne]3s^2 3p^6$

b. $[Ne]3s^2 3p^6$

c. $[Ne]3s^2 3p^6$

2.109. Determine the electron configuration for the atom first. Then decide if it gains or loses electrons when it forms its ion. If it gains electrons, add them to the electron configuration and vice versa.

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

b. $[He]2s^2 2p^6$

c. $[Ne]3s^2 3p^6$

2.110.

a. $[He]2s^2 2p^6$

b. $[He]2s^2 2p^6$

c. $[He]2s^2 2p^6$

The ions all have the same electron configuration, but they contain different numbers of protons, which in turn results in different ionic charges: O^{2-}, F^-, and Na^+.

2.111. a. [Ne]$3s^2 3p^6$

b. [Ne]$3s^2 3p^6$

c. [Ne]$3s^2 3p^6$

The ions all have the same electron configuration, but they contain different numbers of protons, which in turn results in different ionic charges: S^{2-}, K^+, and Ca^{2+}.

2.112. a. $1s^2$

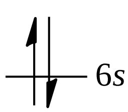
b. [Ar]$4s^2 3d^{10} 4p^6$

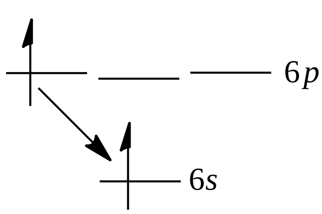
c. [Kr]$5s^2 4d^{10} 5p^6$

2.113. a. Ca^{2+} b. K^+ c. $:\underset{\cdot\cdot}{\overset{\cdot\cdot}{F}}:^{-}$ d. $:\underset{\cdot\cdot}{\overset{\cdot\cdot}{O}}:^{2-}$ e. $:\underset{\cdot\cdot}{\overset{\cdot\cdot}{N}}:^{3-}$

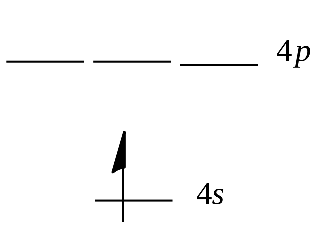
2.114. a. $:\underset{\cdot\cdot}{\overset{\cdot\cdot}{As}}:^{3-}$ b. $:\underset{\cdot\cdot}{\overset{\cdot\cdot}{S}}:^{2-}$ c. $:\underset{\cdot\cdot}{\overset{\cdot\cdot}{Br}}:^{-}$ d. Li^+ e. Cs^+

2.115. K^+ $:\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}:^{-}$

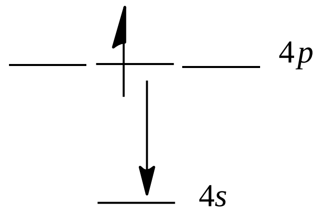
2.116. [Xe]$6s^2$. 



2.117. [Ar]$4s^2 3d^9$ (Please note that some elements defy our predicted electron configurations. Copper actually has a configuration of [Ar]$4s^1 3d^{10}$, but you are not expected to know this at this stage.)

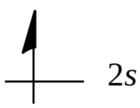


ground state



excited state
(atom absorbed energy and the 4*s* electron moved to an empty 4*p* orbital)

2.118. $E = \frac{hc}{\lambda} = 4.32 \times 10^{-19}\,\mathrm{J} =$ [illegible]
$10^{-?} \times 4.60 = \lambda$ m

2.119. [He]$2s^1$, 

$$E = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34}\,\mathrm{Js} \times 3.00 \times 10^{8}\,\mathrm{m/s}}{6.70 \times 10^{-7}\,\mathrm{m}} = 2.97 \times 10^{-19}\,\mathrm{J}$$

2.120. [Xe]$6s^1 6p^1$

2.121. [He]$2p^1$