Chemistry 2e
6: Electronic Structure and Periodic Properties of Elements
6.1: Electromagnetic Energy

1. The light produced by a red neon sign is due to the emission of light by excited neon atoms. Qualitatively describe the spectrum produced by passing light from a neon lamp through a prism.

Solution

The spectrum consists of colored lines, at least one of which (probably the brightest) is red.

3. FM–95, an FM radio station, broadcasts at a frequency of $9.51 \times 10^7\ s^{-1}$ (95.1 MHz). What is the wavelength of these radio waves in meters?

Solution

$$\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m s}^{-1}}{9.51 \times 10^7 \text{ s}^{-1}} = 3.15 \text{ m}$$

5. Light with a wavelength of 614.5 nm looks orange. What is the energy, in joules, per photon of this orange light? What is the energy in eV ($1 \text{ eV} = 1.602 \times 10^{-19}$ J)?

Solution

$c = \lambda\nu$

$$\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \cancel{\text{m}} \text{ s}^{-1}}{614.5 \times 10^{-9} \cancel{\text{m}}} = 4.879 \times 10^{14} \text{ s}^{-1}$$

$$E = h\nu = 6.626 \times 10^{-34} \text{ J} \cancel{\text{s}} \times 4.879 \times 10^{14} \cancel{\text{s}}^{-1} = 3.233 \times 10^{-19} \text{ J}$$

$$E = \frac{3.233 \times 10^{-19} \cancel{\text{J}}}{1.602 \times 10^{-19} \cancel{\text{J}} \text{ eV}^{-1}} = 2.018 \text{ eV}$$

7. A photon of light produced by a surgical laser has an energy of 3.027×10^{-19} J. Calculate the frequency and wavelength of the photon. What is the total energy in 1 mole of photons? What is the color of the emitted light?

Solution

$E = h\nu$

$$\nu = \frac{3.027 \times 10^{-19} \cancel{\text{J}}}{6.626 \times 10^{-34} \cancel{\text{J}} \text{ s}^{-1}} = 4.568 \times 10^{14} \text{ s}^{-1}$$

$$\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m } \cancel{\text{s}^{-1}}}{4.568 \times 10^{14} \cancel{\text{s}^{-1}}} = 6.563 \times 10^{-7} \text{ m} = 6563 \text{ Å} = 656.3 \text{ nm}$$

Energy $\text{mol}^{-1} = 3.027 \times 10^{-19} \text{ J} \times 6.022 \times 10^{23} \text{ mol}^{-1} = 1.823 \times 10^5 \text{ J mol}^{-1}$; red.

9. The emission spectrum of cesium contains two lines whose frequencies are (a) 3.45×10^{14} Hz and (b) 6.53×10^{14} Hz. What are the wavelengths and energies per photon of the two lines? What color are the lines?

Solution

From $c = \lambda\nu$

(a) $$\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m } \cancel{\text{s}^{-1}}}{3.45 \times 10^{14} \cancel{\text{s}^{-1}}} = 8.69 \times 10^{-7} \text{ m}$$

$E = h\nu = 6.626 \times 10^{-34} \text{ J} \cancel{\text{s}} \times 3.45 \times 10^{14} \cancel{\text{s}^{-1}} = 2.29 \times 10^{-19}$ J

(b) $$\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8 \text{ m } \cancel{\text{s}^{-1}}}{6.53 \times 10^{14} \cancel{\text{s}^{-1}}} = 4.59 \times 10^{-7} \text{ m}$$

$E = h\nu = 6.626 \times 10^{-34} \text{ J} \cancel{\text{s}} \times 6.53 \times 10^{14} \cancel{\text{s}^{-1}} = 4.33 \times 10^{-19}$ J

The color of (a) is red; (b) is blue.

11. One of the radiographic devices used in a dentist's office emits an X-ray of wavelength 2.090 $\times$ 10^{-11} m. What is the energy, in joules, and frequency of this X-ray?

Solution

$c = \lambda\nu$,

$$\nu = \frac{c}{\lambda} = \frac{2.998 \times 10^{8}\ \cancel{\text{m}}\ \text{s}^{-1}}{2.090 \times 10^{-11}\ \cancel{\text{m}}} = 1.434 \times 10^{19}\ \text{s}^{-1}$$

$E = h\nu = 6.626 \times 10^{-34}\ \text{J}\ \cancel{\text{s}} \times 1.434 \times 10^{19} \cancel{\text{s}^{-1}} = 9.502 \times 10^{-15}\ \text{J}$

13. RGB color television and computer displays use cathode ray tubes that produce colors by mixing red, green, and blue light. If we look at the screen with a magnifying glass, we can see individual dots turn on and off as the colors change. Using a spectrum of visible light, determine the approximate wavelength of each of these colors. What is the frequency and energy of a photon of each of these colors?

Solution

See Figure 6.13. Red: 660 nm; 4.54×10^{14} Hz; 3.01×10^{-19} J. Green: 520 nm; 5.77×10^{14} Hz; 3.82×10^{-19} J. Blue: 440 nm; 6.81×10^{14} Hz; 4.51×10^{-19} J. Somewhat different numbers are also possible.

15. What is the threshold frequency for sodium metal if a photon with frequency 6.66×10^{14} s^{-1} ejects a photon with 7.74×10^{-20} J kinetic energy? Will the photoelectric effect be observed if sodium is exposed to orange light?

Solution

$E_{\text{photon}} = h\nu = (6.626 \times 10^{-34}\ \text{J s}) \times (6.66 \times 10^{14}\ \text{s}^{-1}) = 4.41 \times 10^{-19}\ \text{J}$

$E_{\text{photon}} = E_{\text{kinetic}} + E_{\text{threshold}}$,.which means that $4.41 \times 10^{-19}\ \text{J} = 7.74 \times 10^{-20}\ \text{J} + E_{\text{threshold}}$, thus $E_{\text{threshold}} = 3.64 \times 10^{-19}$ J, which corresponds to a threshold frequency of $5.49 \times 10^{14}\ \text{s}^{-1}$. Orange light is roughly 620 nm or $4.84 \times 10^{14}\ \text{s}^{-1}$, which is less than the threshold, so no electrons will be ejected.

Chemistry 2e
6: Electronic Structure and Periodic Properties of Elements
6.2: The Bohr Model

17. What does it mean to say that the energy of the electrons in an atom is quantized?

Solution

Quantized energy means that the electrons can possess only certain discrete energy values; values between those quantized values are not permitted.

19. The electron volt (eV) is a convenient unit of energy for expressing atomic-scale energies. It is the amount of energy that an electron gains when subjected to a potential of 1 volt; 1 eV = 1.602×10^{-19} J. Using the Bohr model, determine the energy, in electron volts, of the photon produced when an electron in a hydrogen atom moves from the orbit with $n = 5$ to the orbit with $n = 2$. Show your calculations.

Solution

$$E = E_2 - E_5 = 2.179 \times 10^{-18}\left(\frac{1}{n_2^2} - \frac{1}{n_5^2}\right)\text{ J}$$

$$= 2.179 \times 10^{-18}\left(\frac{1}{2^2} - \frac{1}{5^2}\right) = 4.576 \times 10^{-19}\text{ J}$$

$$= \frac{4.576 \times 10^{-19}\ \cancel{\text{J}}}{1.602 \times 10^{-19}\ \cancel{\text{J}}\ \text{eV}^{-1}} = 2.856\text{ eV}$$

21. Using the Bohr model, determine the lowest possible energy for the electron in the He^{+} ion.

Solution

$E = -2.179 \times 10^{-18}\text{ J}\left(Z^2\right)\left(\frac{1}{n^2}\right)$. Since Z is the number of protons in the nucleus, $Z = 2$ for He. $E = -2.179 \times 10^{-18}\text{ J}(4)\left(\frac{1}{1}\right) = -8.716 \times 10^{-18}\text{ J}$

23. Using the Bohr model, determine the energy of an electron with $n = 8$ in a hydrogen atom.

Solution

$E = -2.179 \times 10^{-18}\text{ J}\left(Z^2\right)\left(\frac{1}{n^2}\right)$. Since Z is the number of protons in the nucleus, $Z = 1$ for H.

$$E = -2.179 \times 10^{-18}\text{ J}(1)\left(\frac{1}{64}\right) = -3.405 \times 10^{-20}\text{ J}$$

25. What is the radius, in angstroms, of the orbital of an electron with $n = 8$ in a hydrogen atom?

Solution

$r = \frac{n^2 a_0}{Z}$. For hydrogen, Z = 1; $a_0 = 0.529$ Å; $n = 8$.

$$r = \frac{8^2(0.529\text{ Å})}{1} = 64(0.529) = 33.9\text{ Å}$$

27. Using the Bohr model, determine the energy in joules of the photon produced when an electron in a Li^{2+} ion moves from the orbit with $n = 2$ to the orbit with $n = 1$.

Solution

$$E = 2.179 \times 10^{-18} \text{ J}(Z^2)\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

$$= 2.179 \times 10^{-18} \text{ J}(3^2)\left(\frac{1}{1^2} - \frac{1}{2^2}\right)$$

$$= 1.471 \times 10^{-17} \text{ J}$$

29. How are the Bohr model and the Rutherford model of the atom similar? How are they different?

Solution

Both involve a relatively heavy nucleus with electrons moving around it, although strictly speaking, the Bohr model works only for one-electron atoms or ions. According to classical mechanics, the Rutherford model predicts a miniature “solar system” with electrons moving about the nucleus in circular or elliptical orbits that are confined to planes. If the requirements of classical electromagnetic theory that electrons in such orbits would emit electromagnetic radiation are ignored, such atoms would be stable, having constant energy and angular momentum, but would not emit any visible light (contrary to observation). If classical electromagnetic theory is applied, then the Rutherford atom would emit electromagnetic radiation of continually increasing frequency (contrary to the observed discrete spectra), thereby losing energy until the atom collapsed in an absurdly short time (contrary to the observed long-term stability of atoms). The Bohr model retains the classical mechanics view of circular orbits confined to planes having constant energy and angular momentum, but restricts these to quantized values dependent on a single quantum number, *n*. The orbiting electron in Bohr’s model is assumed not to emit any electromagnetic radiation while moving about the nucleus in its stationary orbits, but the atom can emit or absorb electromagnetic radiation when the electron changes from one orbit to another. Because of the quantized orbits, such “quantum jumps” will produce discrete spectra, in agreement with observations.

Chemistry 2e
6: Electronic Structure and Periodic Properties of Elements
6.3: Development of Quantum Theory

31. How are the Bohr model and the quantum mechanical model of the hydrogen atom similar? How are they different?

Solution

Both models have a central positively charged nucleus with electrons moving about the nucleus in accordance with the Coulomb electrostatic potential. The Bohr model *assumes* that the electrons move in circular orbits that have quantized energies, angular momentum, and radii that are specified by a single quantum number, n = 1,2,3,…, but this quantization is an ad hoc assumption made by Bohr to incorporate quantization into an essentially classical mechanics description of the atom. Bohr also assumed that electrons orbiting the nucleus normally do not emit or absorb electromagnetic radiation, but do so when the electron switches to a different orbit. In the quantum mechanical model, the electrons do not move in precise orbits (such orbits violate the Heisenberg uncertainty principle) and, instead, a probabilistic interpretation of the electron's position at any given instant is used, with a mathematical function ψ called a wave function that can be used to determine the electron's spatial probability distribution. These wave functions, or orbitals, are three-dimensional stationary waves that can be specified by three quantum numbers that arise naturally from their underlying mathematics (no ad hoc assumptions required): the principal quantum number, n (the same one used by Bohr), which specifies shells such that orbitals having the same n all have the same energy and approximately the same spatial extent; the angular momentum quantum number l, which is a measure of the orbital's angular momentum and corresponds to the orbitals' general shapes, as well as specifying subshells such that orbitals having the same l (and n) all have the same energy; and the orientation quantum number m, which is a measure of the z component of the angular momentum and corresponds to the orientations of the orbitals. The Bohr model gives the same expression for the energy as the quantum mechanical expression and, hence, both properly account for hydrogen's discrete spectrum (an example of getting the right answers for the wrong reasons, something that many chemistry students can sympathize with), but gives the wrong expression for the angular momentum (Bohr orbits necessarily all have non-zero angular momentum, but some quantum orbitals [*s* orbitals] can have zero angular momentum).

33. Describe the properties of an electron associated with each of the following four quantum numbers: n, l, m_l, and m_s.

Solution

n determines the general range for the value of energy and the probable distances that the electron can be from the nucleus. l determines the shape of the orbital. m_1 determines the orientation of the orbitals of the same l value with respect to one another. m_s determines the spin of an electron.

35. Identify the subshell in which electrons with the following quantum numbers are found:
(a) $n = 2$, $l = 1$
(b) $n = 4$, $l = 2$
(c) $n = 6$, $l = 0$

Solution

(a) 2*p*, (b) 4*d*, (c) 6*s*

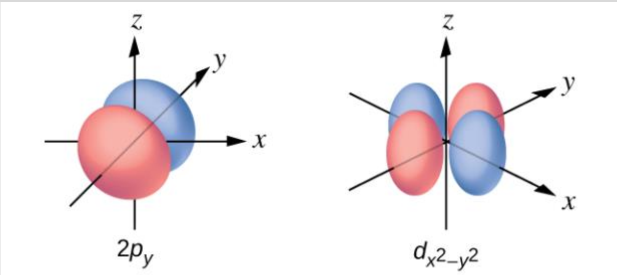
37. Identify the subshell in which electrons with the following quantum numbers are found:
(a) $n = 3$, $l = 2$

(b) $n = 1$, $l = 0$
(c) $n = 4$, $l = 3$

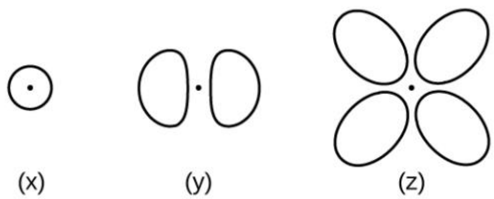
Solution
(a) 3*d;* (b) 1*s;* (c) 4*f*

39. Sketch the boundary surface of a $d_{x^2-y^2}$ and a p_y orbital. Be sure to show and label the axes.

Solution



41. Consider the orbitals shown here in outline.



(a) What is the maximum number of electrons contained in an orbital of type (x)? Of type (y)? Of type (z)?
(b) How many orbitals of type (x) are found in a shell with $n = 2$? How many of type (y)? How many of type (z)?
(c) Write a set of quantum numbers for an electron in an orbital of type (x) in a shell with $n = 4$. Of an orbital of type (y) in a shell with $n = 2$. Of an orbital of type (z) in a shell with $n = 3$.
(d) What is the smallest possible *n* value for an orbital of type (x)? Of type (y)? Of type (z)?
(e) What are the possible *l* and m_l values for an orbital of type (x)? Of type (y)? Of type (z)?

Solution
(a) x. 2, y. 2, z. 2; (b) x. 1, y. 3, z. 0; (c) x. 4 0 0 $\frac{1}{2}$, y. 2 1 0 $\frac{1}{2}$, z. 3 2 0 $\frac{1}{2}$; (d) x. 1, y. 2, z. 3; (e) x. $l = 0$, $m_l = 0$, y. $l = 1$, $m_l = –1$ 0 or +1, z. $l = 2$, $m_l = –2$ –1 0 +1 +2

43. How many electrons could be held in the second shell of an atom if the spin quantum number m_s could have three values instead of just two? (Hint: Consider the Pauli exclusion principle.)

Solution
In the second subshell ($n = 2$), there are two possible angular momentum (l) values (0 and 1). When $l = 0$, the magnetic quantum number (m_l) can only be 0. When $l = 1$, the magnetic quantum number (m_l) can be equal to –1, 0, or 1. Therefore, in the second shell, there are four orbitals (the number of orbitals in a shell is equal to n^2). If there were three possible values for the spin quantum number, each orbital could hold three different electrons and still obey the Pauli exclusion principle. In other words, each orbital could hold three electrons and none of those three electrons would have the exact same four quantum numbers. If each orbital could hold three electrons, the second shell, which has four orbitals, could hold a total of 12 electrons.

45. Write a set of quantum numbers for each of the electrons with an *n* of 4 in a Se atom.

Solution

n	l	m_l	s
4	0	0	$+\frac{1}{2}$
4	0	0	$-\frac{1}{2}$
4	1	−1	$+\frac{1}{2}$
4	1	0	$+\frac{1}{2}$
4	1	+1	$+\frac{1}{2}$
4	1	−1	$-\frac{1}{2}$

Chemistry 2e
6: Electronic Structure and Periodic Properties of Elements
6.4: Electronic Structure of Atoms (Electron Configurations)

47. Read the labels of several commercial products and identify monatomic ions of at least six main group elements contained in the products. Write the complete electron configurations of these cations and anions.

Solution

For example, Na^{+}: $1s^2 2s^2 2p^6$; Ca^{2+}: $1s^2 2s^2 2p^6 3s^2 3p^6$; Sn^{2+}: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2$; $F^{–}$: $1s^2 2s^2 2p^6$; $O^{2–}$: $1s^2 2s^2 2p^6$; $Cl^{–}$: $1s^2 2s^2 2p^6 3s^2 3p^6$.

49. Using complete subshell notation ($1s^2 2s^2 2p^6$, and so forth), predict the electron configuration of each of the following atoms:

(a) N
(b) Si
(c) Fe
(d) Te
(e) Tb

Solution

(a) $1s^2 2s^2 2p^3$; (b) $1s^2 2s^2 2p^6 3s^2 3p^2$; (c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$; (d) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^4$; (e) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^9$

51. What additional information do we need to answer the question "Which ion has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6$"?

Solution

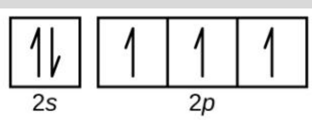
The charge on the ion.

53. Use an orbital diagram to describe the electron configuration of the valence shell of each of the following atoms:

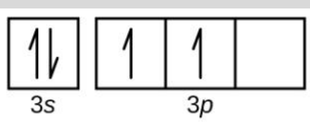
(a) N
(b) Si
(c) Fe
(d) Te
(e) Mo

Solution

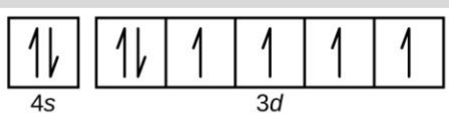
(a)



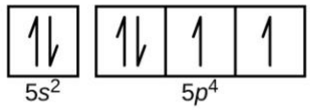
(b)



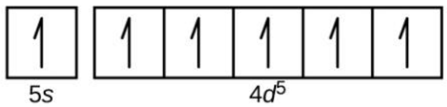
(c)



(d)



(e)



55. Which atom has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^2$?

Solution

Zr

57. Which ion with a +1 charge has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$? Which ion with a –2 charge has this configuration?

Solution

Rb^+, Se^{2-}

59. Which of the following has two unpaired electrons?

(a) Mg
(b) Si
(c) S
(d) Both Mg and S
(e) Both Si and S.

Solution

Although both (b) and (c) are correct, (e) encompasses both and is the best answer.

61. Which atom would be expected to have a half-filled 4*s* subshell?

Solution

K

63. Thallium was used as a poison in the Agatha Christie mystery story "The Pale Horse." Thallium has two possible cationic forms, +1 and +3. The +1 compounds are the more stable. Write the electron structure of the +1 cation of thallium.

Solution

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10}$

65. Cobalt–60 and iodine–131 are radioactive isotopes commonly used in nuclear medicine. How many protons, neutrons, and electrons are in atoms of these isotopes? Write the complete electron configuration for each isotope.

Solution

Co has 27 protons, 27 electrons, and 33 neutrons: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$.
I has 53 protons, 53 electrons, and 78 neutrons: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^5$.

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6: Electronic Structure and Periodic Properties of Elements
6.5: Periodic Variations in Element Properties

67. Based on their positions in the periodic table, predict which has the smallest atomic radius: Mg, Sr, Si, Cl, I.

Solution
Cl

69. Based on their positions in the periodic table, predict which has the largest first ionization energy: Mg, Ba, B, O, Te.

Solution
O

71. Based on their positions in the periodic table, rank the following atoms in order of increasing first ionization energy: F, Li, N, Rb

Solution
Rb < Li < N < F

73. Atoms of which group in the periodic table have a valence shell electron configuration of ns^2np^3?

Solution
15 (5A)

75. Based on their positions in the periodic table, list the following atoms in order of increasing radius: Mg, Ca, Rb, Cs.

Solution
Mg < Ca < Rb < Cs

77. Based on their positions in the periodic table, list the following ions in order of increasing radius: K^+, Ca^{2+}, Al^{3+}, Si^{4+}.

Solution
$Si^{4+} < Al^{3+} < Ca^{2+} < K^+$

79. Which atom and/or ion is (are) isoelectronic with Br^+: Se^{2+}, Se, As^-, Kr, Ga^{3+}, Cl^-?

Solution
Se, As^-

81. Compare both the numbers of protons and electrons present in each to rank the following ions in order of increasing radius: As^{3-}, Br^-, K^+, Mg^{2+}.

Solution
$Mg^{2+} < K^+ < Br^- < As^{3-}$

83. Of the five elements Sn, Si, Sb, O, Te, which has the most endothermic reaction? (E represents an atom.) What name is given to the energy for the reaction?

$$\text{E}(g) \longrightarrow \text{E}^+(g) + \text{e}^-$$

Solution
O, IE_1

85. Which main group atom would be expected to have the lowest second ionization energy?

Solution
Ra