Lota_2

A1_Lotukerfið og uppbygging atómanna

The Electromagnetic Spectrum

6.8 Which of the waves depicted here has the highest frequency? Explain your answer.

ANS "c" has the highest frequency because that photon has the most waves per unit time.

6.10 Calculate the wavelengths, in meters, of radiation of the following frequencies: (a) 5.00×10^{15} s^{-1}, (b) 2.11×10^{14} s^{-1}, (c) 5.44×10^{12} s^{-1}

ANS (a) 6.00 x 10^{-8} m, (b) 1.42 x 10^{-6} m, (c) 5.51 x 10^{-5} m

6.12 Define the term *refraction*.

ANS The bending of light as it passes from one media to another.

6.14 What is the relationship between the number of photons and the intensity of light?

ANS The greater the number of photons, the brighter (more intense) the light.

6.16 Place these types of radiation in order of increasing energy per photon. (a) green light from a mercury lamp, (b) X-rays from a dental X-ray, (c) microwaves in a microwave oven, (d) an FM music station broadcasting at 89.1 MHz

ANS 89.1 MHz < microwaves < green light < dental X-rays

6.18 For photons with the following energies, calculate the wavelength and identify the region of the spectrum they are from (a). 6.0×10^{-19} J (b). 8.7×10^{-22} J (c). 3.2×10^{-24} J (d). 1.9×10^{-28} J

ANS Solve $E = hc/\lambda$ for wavelength

(a) 3.3×10^{-7} m = UV

(b) 2.3×10^{-4} m = microwave

(c) 6.2×10^{-2} m = radio (super high frequency, SHF)

(d) 1.0×10^{3} m = radio (low frequency, LF)

6.20 The laser in most supermarket barcode scanners operates at a wavelength of 632.8 nm. What is the energy of a single photon emitted by such a laser? What is the energy of one mole of these photons?

ANS $E = 3.141 \times 10^{-19}$ J/photon; 189.2 kJ/mole of photons

6.22 Fill in the blanks below to complete a description of the photoelectric effect experiment. (You should be able to do this with just one or two words in each blank.)

A beam of photons strikes a metal surface, causing electrons to be emitted.

6.24 The electron binding energy for copper metal is 7.18×10^{-19} J. Find the longest wavelength of light that could eject electrons from copper in a photoelectric effect experiment.

ANS The photon must have a minimum energy above the binding energy of the electron. Therefore, use the binding energy value to calculate the minimum energy needed and convert this energy into a wavelength as in 6.18.

$\lambda = 2.77 \times 10^{-7}$ m

Atomic Spectra

6.26 What was novel about Bohr's model of the atom?

ANS Electrons reside in stable specific orbits around a nucleus. To go from one orbit to another, specific amounts of energy are absorbed or released, i.e. quantized energy.

6.28 According to the Bohr model of the atom, what happens when an atom absorbs energy?

ANS An electron will be excited to a higher energy level.

6.30 A neon atom emits light at many wavelengths, two of which are at 616.4 nm and 638.3 nm. Both of these transitions are to the same final state. (a). What is the energy difference between the two states for each transition? (b). If a transition between the two higher energy states could be observed, what would be the frequency of the light?

<u>ANS</u> (a) 1.11×10^{-20} J, (b) 1.68×10^{13} Hz

6.32 Describe how excited states of atoms play a role in light bulb technology.

<u>ANS</u> Electrons are excited to higher energy levels and emit light (photons) as the electrons return to ground state.

The Quantum Mechanical Model of the Atom

6.34 How did the observation of electron diffraction affect the development of the quantum mechanical model of the atom?

<u>ANS</u> Diffraction is a phenomenon typically associated with waves. This introduced the idea that wave-based treatment plays a critical part in explaining electron behavior.

6.36 What are the mathematical origins of quantum numbers?

<u>ANS</u> Quantum numbers describe atomic orbitals which arose from the wave equation solutions of Schrodinger.

6.38 Which of the following represent valid sets of quantum numbers? For a set that is invalid, explain briefly why it is not correct. (a) $n = 3$, $\ell = 3$, $m_\ell = 0$, (b) $n = 2$, $\ell = 1$, $m_\ell = 0$, (c) $n = 6$, $\ell = 5$, $m_\ell = -1$, (d) $n = 4$, $\ell = 3$, $m_\ell = -4$

ANS (a) not allowed because l must be no higher than n-1

(b) allowed

(c) allowed

(d) not allowed because m_ℓ has to be between -ℓ and +ℓ

6.40 Why are there no *2d* orbitals?

ANS In referring to "2d", the n value of 2 gives rise to l values of 1 and 0. The 1 corresponds to the "p" orbital and the 0 refers to the "s" orbital. A potential l value of 2 would have to be permitted to have a "d" orbital. This is not possible; therefore, there is no "2d" orbital.

6.42 How many orbitals correspond to each of the following designations? (a) 3*p*, (b) 4*p*, (c) $4p_x$, (d) 6*d*, (e) 5*d*, (f) 5*f*, (g) $n = 5$, (h) 7*s*

ANS (a) 3 (b) 3 (c) 1 (d) 5 (e) 5 (f) 7 (g) 25 (h) 1

6.44 Define the term *nodal plane* (or *node*).

ANS A plane that separates the lobes of an orbital. There is 0 probability of finding an electron within this plane.

Pauli Exclusion Principle and Electron Configurations

6.46 Does the fact that the fourth quantum number is called the spin quantum number prove that electrons are spinning? What observations gave rise to this name?

ANS No, while it was first thought that the electrons were literally spinning, this is no longer taken in the literal sense. The observation of a magnetic field first led scientists to the conclusion that the electron was spinning in one of two directions in order to generate these observed fields.

6.48 On what does the Pauli Exclusion Principle place a limit?

ANS Pauli Exclusion places a two-electron limit on all sub-orbitals. No two electrons may have the same four quantum numbers; therefore, there is no means to have more than two electrons in each sub-orbital.

6.50 Why does the size of an orbital have an effect on its energy?

ANS As the orbital is further removed from the nucleus (diameter increases) the forces of attraction between electrons and the nucleus decrease.

6.52 Depict two ways to place electrons in the *2p* orbitals for a nitrogen atom. Which depiction is correct according to Hund's Rule?

ANS ↿⇂ ↿_ __ or ↿_ ↿_ ↿_; the depiction on the right is correct

6.54 Write the ground state electron configuration for (a) B, (b) Ba, (c) Be, (d) Bi, (e) Br.

ANS (a) B = $1s^22s^22p^1$

(b) Ba = $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^2$

(c) Be = $1s^22s^2$

(d) Bi = $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^3$

(e) Br = $1s^22s^22p^63s^23p^64s^23d^{10}4p^5$

6.56 From the list of atoms and ions given, identify any pairs that have the same electron configurations and write that configuration. Na^+, S^{2-}, Ne, Ca^{2+}, Fe^{2+}, Kr, I^-

ANS Na^+, Ne = $1s^22s^22p^6$

Ca^{2+}, S^{2-} = $1s^22s^22p^63s^23p^6$

Periodic Table and Electron Configurations

6.58 Distinguish between the terms *core electrons* and *valence electrons*.

ANS Valence electrons are those electrons in the outermost shell; core electrons are in the filled shells (orbitals) and are seldom involved in reactions.

6.60 Why is there no element to the immediate right of magnesium in the periodic table?

ANS There is no means to have a "2d" orbital.

6.62 Which blocks of the periodic table comprise the main group elements?

ANS The "s" block and "p" block

6.64 Look at the table of electron configurations in Appendix C. Which elements have configurations that are exceptions to the aufbau principle? Propose a reason why these elements have these exceptions.

ANS Several have electron configurations that are exceptions to Hund's Rule. Most of these exceptions occur in the "d" and "f" orbitals. The energy levels of the "d" and "f" orbitals lies very close to that of the "s" orbitals within a given period. Because of these energetics, electrons can move to what, at first, appear to be a higher orbital. This phenomenon tends to occur in situations that lead to a closed or half-closed (maximized) "d" orbital.

Periodic Trends of Atomic Properties

6.66 Use the electronic configuration of the alkali metals and your knowledge of orbitals and quantum numbers to explain the trend in atomic size of alkali metal atoms.

ANS As the "n" value increases down a family (or group), the size of the atom increases.

6.68 Define the term *ionization energy*. What is the difference between the first ionization energy and the second ionization energy?

ANS Ionization energy refers to the amount of energy required to remove an electron from an atom or ion. Second ionization energy refers to the energy required to remove a second electron from the ion.

6.70 Arrange the following atoms in order of increasing ionization energy: Li, K, C, and N.

ANS K < Li < C < N

6.72 Which element would you expect to have the largest second ionization energy, Na, C, or F? Why?

ANS Na because the removal of the first electron creates a Na^{1+} ion that would be isoelectric with Ne. Removing a second electron from Na^{1+} would be at least as difficult as removing the first electron from neon.

6.74 Indicate which species in each pair has the more favorable (more negative) electron affinity. Explain your answers. (a) Cl or S, (b) S or P, (c) Br or As

ANS (a) Cl, (b) S, (c) Br

6.76 Rank the following in order of decreasing ionization energy. Cl, F, Ne^+, S, S^-

ANS Ne^+, F, Cl, S, S^-

Modern Light Sources: LEDs and Lasers

6.78 The compact fluorescent bulb is another fairly new light source that is likely to have a large impact on home lighting. These bulbs combine the energy efficiency of traditional fluorescent lights with the smaller size, "warmer" light, and dimming ability of incandescent lights. Although compact fluorescent bulbs are expensive, manufacturers claim maintain that subsequent savings will more than offset the initial costs. Analyze the relative cost of incandescent versus compact fluorescent lighting, assuming the information in the table below is accurate. Are industry claims of cost savings justified?

ANS Yes, the industry claims are correct because the operation cost of the compact fluorescent is \$7.20 for 10,000 hours compared to \$30 for incandescent lighting.

Additional Problems

6.80 The photoelectric effect can be used to measure the value of Planck's constant. Suppose that a photoelectric effect experiment was carried out using light with $\nu = 7.50 \times 10^{14}\ s^{-1}$ and ejected electrons were detected with a kinetic energy of 2.50×10^{-11} J. The experiment was then repeated using light with $\nu = 1.00 \times 10^{15}\ s^{-1}$ and the *same metal target,* and electrons were ejected with kinetic energy of 5.00×10^{-11} J. Use these data to find a value for Planck's constant. **HINTS:** These data are fictional, and will give a result that is quite different from the real value of Planck's constant. Be sure that you do *not* use the real value of Planck's constant in any calculations here. It may help to start by thinking about how you would calculate the metal's binding energy if you already knew Planck's constant.

ANS Recall that $KE_{e-} = h\nu - BE_{e-}$; BE and h are both unknowns, but two equations can be formed with the KE formula. Solving these equations leads to a value of $h = 1.00 \times 10^{25}$ Js.

6.82 When a photoelectric effect experiment was carried out using a metal 'M' and light at wavelength λ_1, electrons with a kinetic energy of 1.6×10^{-19} J were emitted. The wavelength was reduced to one-half of its original value and the experiment was repeated (still using the same metal target). This time electrons with a kinetic energy of 6.4×10^{-19} J were emitted. Find the electron binding energy for metal M. (The actual wavelengths used were not recorded, but it is still possible to find the binding energy.)

ANS Recall that $KE_{e-} = h\nu - BE_{e-}$; Assume the original frequency to be "ν". If the wavelength is halved, frequency must double. Use a value of 2ν in the second application of the equation. Solve for BE_{e-}. The binding energy is 3.2×10^{-19} J.

6.84 When a helium atom absorbs light at 58.44 nm, an electron is promoted from the 1s orbital to a *2p*-orbital. Given that the ionization energy of (ground state) helium is 2,372 kJ/mol, find the longest wavelength of light that could eject an electron from the excited state helium atom.

ANS The energy to ionize a ground state helium atom is 3.939×10^{-18} J. The helium electron absorbed 3.401×10^{-18} J to become excited. The difference in the ionization energy and the energy absorbed to excite the helium is the energy required to ionize the excited atom. This is 5.38×10^{-19} J. The wavelength needed to do this is 3.69×10^{-7} m or 369 nm.

Focus on Problem Solving

6.86 Arrange the following sets of anions in order of increasing ionic radii:

(a) Cl^-, S^{2-}, P^{3-}, (b) O^{2-}, S^{2-}, Se^{2-}, (c) N^{3-}, S^{2-}, Br^-, P^{3-}, (d) Cl^-, Br^-, I^-

ANS (a) $Cl^- < S^{2-} < P^{3-}$ (b) $O^{2-} < S^{2-} < Se^{2-}$ (c) $N^{3-} < S^{2-} < P^{3-} < Br^-$ (d) $Cl^- < Br^- < I^-$

6.88 Some spacecraft use ion propulsion engines. These engines create thrust by ionizing atoms and then accelerating and expelling them. According to Newton's laws, this leads to thrust in the opposite direction. Minimizing mass is always important for space applications. Discuss what periodic trends you would need to consider for choosing a material to ionize for engineering designs with this application.

ANS The considerations here would be ionization potential, mass, and cost. Alkali metals have low first ionization potentials, but the lightest (Li) has a slightly higher ionization potential than the heaviest of the group rubidium. Availability and cost also factor into the process. Sodium and potassium are potentially less expensive than the other members of Group I. The main concern of this approach is the reactivity of alkali metals, but these factors can be accounted for in the design.

6.90 Ionization gauges for pressure measurement (see Section 5.7) can be manipulated by changing the voltage of the electron collection grid (see Figure 5.11). Referring to Table 6.4, if the voltage is set to yield electron energies of 1200 kJ/mol, determine what elements the gauge would NOT be able to detect. Given the ability to vary this energy, describe how you would design an apparatus that could distinguish between sodium and magnesium atoms.

ANS The gauge would not be able to detect H, He, N, O, F, Ne, Cl, or Ar. To distinguish between Na and Mg atoms, set the threshold to yield 600 kJ/mol. This should detect Na but not Mg.

Cumulative Problems

6.92 The red color seen in fireworks is the result of having strontium-containing salts in the fireworks bomb. Similarly the green/blue colors sometimes seen in fireworks arise from copper containing salts. Based on your understanding of atomic spectra and the colors in fireworks, describe which atom, copper or strontium, has more widely separated energy levels.

ANS Greater distances between energy levels would lead to greater energies of emitted light. Because green/blue is higher in energy than red, the copper containing salts mix has the more widely separated energy levels.

6.94 When Bohr devised his model for the atom was he using deductive or inductive reasoning? Explain your answer.

ANS He was using primarily inductive reasoning in devising the model, but one could argue for both types of reasoning playing a role. There were observations that led to the creation of a model to explain the observations and experiments designed to test the hypothesis (model) that was created.