

CHAPTER 7- Atomic Structure and Periodicity

7.1 Electromagnetic radiation- energy that exhibits wavelike behavior and travels at the speed of light in a vacuum.

Wavelength- (λ) distance between crests or troughs

Frequency- (ν) number of wave cycles per second that pass a given point in space

-there is an inverse relationship between wavelength and frequency

- $\lambda\nu=c$ [λ is in meters, ν is cycles per second, and c is the speed of light (2.9979×10^8 meters per second)]

SAMPLE 7.1

The brilliant red colors seen in fireworks are due to the emission of light with wavelengths around 650 nm when strontium salts such as $\text{Sr}(\text{NO}_3)_2$ and SrCO_3 are heated. (This can be easily demonstrated in the lab by dissolving one of these salts in methanol that contains a little water and igniting the mixture in an evaporating dish.) Calculate the frequency of red light of wavelength 6.50×10^2 nm

Answer:

$$6.5 \times 10^2 \text{ nm} \times (1 \text{ m}) / 10^9 \text{ nm} = 6.50 \times 10^{-7} \text{ m}$$

$$\text{And } \nu = c / \lambda, \text{ so } (2.9979 \times 10^8 \text{ m/s}) / (6.50 \times 10^{-7} \text{ m}) = 4.61 \times 10^{14} \text{ Hz.}$$

7.2 The Nature of Matter

-energy can be gained or lost in integer multiples of $h\nu$.

$E = nh\nu$ [n = an integer, h = Planck's constant (6.626×10^{-34} J*s), and ν = frequency]

-Energy is quantized and can only occur in discrete units of size $h\nu$ ("packets" of energy called quanta)

SAMPLE 7.2

The blue color in fireworks is often achieved by heating copper(I) chloride (CuCl) to about 1200°C . Then the compound emits blue light having a wavelength of 450 nm. What is the increment of energy (the quantum) that is emitted at 4.50×10^2 nm by CuCl ?

Answer:

$$-\nu = c / \lambda, \text{ so } (2.9979 \times 10^8 \text{ m/s}) / (4.50 \times 10^{-7} \text{ m}) = 6.66 \times 10^{14} \text{ s}^{-1} .$$

$$\text{And } E = h\nu = (6.626 \times 10^{-34} \text{ J*s}) (6.66 \times 10^{14} \text{ s}^{-1}) \\ = 4.41 \times 10^{-19} \text{ J.}$$

-EMR can be viewed as a stream of particles called photons

$$E_{\text{photon}} = h\nu = (hc) / \lambda$$

-Einstein's special theory of relativity, $E = mc^2$

- $m = E/c^2$ and $E_{\text{photon}} = hc/\lambda$, so $m = (hc/\lambda)/c^2 = h/(c\lambda)$

-EMR seems to have characteristics that of waves and particles and this phenomenon is sometimes referred to as the dual nature of light

-matter can exhibit wavelike properties

de Broglie equation $m = h/c\lambda = h/v$, so $\lambda = h/mv$

SAMPLE 7.3

-diffraction- results when light is scattered from a regular array of points or lines

- constructive interference- (in phase) peaks of one wave match the peaks of another
- destructive interference- (out of phase)
- -extremely small masses(i.e. photons)- exhibit primarily wave properties
- -intermediate masses(i.e. electrons) – exhibit both particulate and wave properties
- large masses(i.e. macroscopic objects)- exhibit primarily particulate properties (small λ 's)

7.3 The atomic Spectrum of Hydrogen

-continuous spectrum- contains all wavelengths of visible light

-emission spectrum (line spectrum) – wavelengths of visible light emitted by an element

-only certain energies are allowed for the e^- ("E" of the e^- is quantized)

-there are discrete energy levels where the E can be calculated as the e^- moves from one "E" level to the next $E = hc/\lambda$

7.4 The Bohr Model

-Niels Bohr developed the quantum model for hydrogen and proposed that the e^- in a hydrogen atom moves around the nucleus in certain allowed orbits.

-“E” of the electron can be calculated based on what energy level it's in (H atom)

- $E = -2.178 \times 10^{-18} \text{ J}(Z^2 / n^2)$
- -[n is an integer (represents energy level), Z is the nuclear charge (# of protons),
- -large the negative value, stronger the bond between e^- and nucleus
- ground state- lowest possible energy state
- to calculate the change in energy as an electron changes orbits
- $E = E_{\text{final state}} - E_{\text{initial state}}$ (if answer is negative, the atom lost energy – photon is emitted)

SAMPLE 7.4

7.5 The Quantum Mechanical Model of the Atom

-electrons behave like standing waves

- nodes are areas of zero displacement
- a “wave” function, or orbital, is a function of the electrons position in three-dimensional space (not a Bohr orbit)
- Heisenburb uncertainty principle- there is a limitation t just how precisely we can know both the position and momentum of a particle at a given position
- Orbitals are drawn from electron density maps to show where the electron can be found 90% of the time

7.6 Quantum Numbers

- describe various properties of the orbital
- principal quantum number(n)- related to the size and energy of the orbital
- angular quantum number(l)- related to the shape of the orbital ($l = n-1$ and corresponds to s, p, d, f, and g...) –subshell
- magnetic quantum number(m_l)- orientation of the orbital in space
- relative to other orbitals (l and $-l$)- value of subshell

7.7 Orbital Shapes and Energies

- nodes- areas of zero probability of finding an e^- , increases with n (# of nodes = $n-1$)
- orbital surfaces increase with “n”
- all orbitals with the same value of “n” have the same energy

7.8 Electron Spin and the Pauli Exclusion Principle

- electron spin quantum number (m_s)= $+(1/2)$ or $-(1/2)$ (only 2 possible orientations)
- Pauli Exclusion Principle- no two electrons can have the same set of four quantum numbers (n, l, m_l , or m_s), thus only 2 electrons can be contained within an orbital

7.9 Polyelectronic Atoms

-electron correlation problem- electron propulsions cannot be calculated exactly due to the unknown electron pathways

-we treat this problem by looking at each electron individually

-polyelectronic atoms have “hydrogen-like” orbitals, but they differ in size and energy due to the difference in nuclear attraction and electron propulsion

-all orbitals with the same principle quantum numbers have the same energy in the H atom (degenerate) which is not the case in polyelectronic atoms

$-E_{ns} < E_{np} < E_{nd} < E_{nf}$ (transparency) has greater electron penetration

- (a) – more strongly bonded to nucleus
- (b) – able to penetrate shielding electrons easier

7.10 The History of the Periodic Table (skip)

7.11 The Aufbau Principle and the Periodic Table

- aufbau principle- as protons are added one by one to build up the nucleus, electrons are similarly added to the “hydrogen-like” orbitals
- Hund’s rule- the lowest energy configuration is the one having the maximum number of unpaired electrons in a particular set of degenerate orbitals (parallel spins)
- valence electrons- electrons in the outermost principal quantum level
- core electrons- inner electrons
 - elements with the same valence electron configuration show similar chemical behaviors
- expected electron configuration- follow rules exceptional
 - *(Cu and Cr break the rules)
- Peculiarity in Electron Filled-

- (1) $(n-1)$ orbitals always fill before nd orbitals (ex. 4s before 3d). The $(n+1)s$ is a lower energy orbital due to more penetration in the vicinity of the nucleus
- (2) – lanthanide series (4f orbitals) and actinide series (5f orbitals)-
 - some times electrons will fill the d orbitals instead of the f orbitals due to very similar energies
 - d electrons are not considered valence electrons because they fill one period late. Group #'s indicate number of valence e^- 's.

7.12 Periodic Trends in Atomic Properties

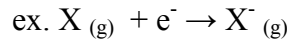
- ionization energy- energy required to remove an electron from a gaseous atom or ion where the atom or ion is assumed to be in their ground state
 - ex. $X_{(s)} \rightarrow X^+_{(g)} + e^-$
- 1st ionization energy- energy required to remove the highest energy electron of an atom
 - ionization energies increase as more electrons are removed from an atom (electrons become more tightly bound due to the increase the atom (+) charge)
 - generally the first ionization energy increases as we go from left to right across a period (due to increased nuclear charge). First ionization energies decrease as we move down a group (e^- 's are further from the nucleus)

SAMPLE 7.9

$1s^2 2s^2 2p^6$ (neon). Because this element is found at the right end of Period 2. Since the 2p electrons do not shield each other very effectively, I_1 will be relatively large. The other configurations given include 3s electrons. These electrons are effectively shielded by the core electrons and are farther from the nucleus than the 2p electrons in neon. Thus I_1 for these atoms will be smaller than for neon.

The atom with the smallest value of I_2 is the one with the configuration $1s^2 2s^2 2p^6 3s^2$ (magnesium). For magnesium, both I_1 and I_2 involve valence electrons. For the atom with the configuration $1s^2 2s^2 2p^6 3s^1$ (sodium), the second electron lost (corresponding to I_2) is a core electron (from a 2p orbital).

-electron affinity- the energy change associated with the addition of an electron to a gaseous atom



-if adding an e- releases energy than the sign of the electron affinity will be (-)
-generally, electron affinities become more negative as we go from left to right across a period (many exceptions due to electron repulsions as a function of electron configuration)

-electron affinities become more positive (less energy released) as we go down a group

-atomic radius- half the distance between the nuclei in a molecule (nonmetallic atoms) or half the distance between atoms in a solid crystal (metal atoms)

-atomic radii decreases as we go from left to right across a period and increases as we move down a group