Series N° 2

Exercise1 :

1. What are the frequency and wavelength of a photon emitted during a transition from $n = 5$ state to the $n = 2$ state in the hydrogen atom?
2. Calculate the energy associated with the first orbit of He^+. What is the radius of this orbit?

Exercise 2 :

1. A photon is emitted as an atom makes a transition from n = 4 to n = 2 level. What is the frequency, wavelength and energy of the emitted photon?
2. A Hydrogen atom initially in its ground state i.e., n = 1 level, absorbs a photon and ends up in third excited state. What must have been the frequency of the photon? Now the electron makes spontaneous emission and comes back to the ground state. What are the possible frequencies of the photons emitted during this process.

Exercice 3 :

1. For the Balmer series i.e., the atomic transitions where final state of the electron is n = 2, what is the longest and shortest wavelength possible.
2. The wave number of first line of Balmer series of hydrogen atom is 15200 cm^{-1} . What is the wave number of first line of Balmer series of Li^{2+} ion.

Exercise 4 :

1. Using *s*, *p*, *d*, f notations, describe the orbital with the following quantum numbers (a) $n = 2$, $l = 1$, (b) $n = 4$, $l = 0$, (c) $n = 5$, $l = 3$, (d) $n = 3$, $l = 2$
2. What are the possible subshells when n = 4? How many orbitals are contained by each of these subshells?
3. What are the values of principal quantum number (n) and azimuthal quantum number (l) for 3s and 4p orbitals?
4. If a shell contains a maximum of 32 electrons, what is the principal quantum number, n?

Explain, giving reasons, which of the following sets of quantum numbers are not possible.

(a) $n = 0$, $l = 0$, $ml = 0$, $ms = + ½$

(b) n = 1, l = 0, ml = 0, ms = – ½

(c) n = 1, $l = 1$, $ml = 0$, $ms = + ½$

(d) $n = 2$, $l = 1$, $ml = 0$, $ms = – ½$

(e) n = 3, l = 3, ml = –3, ms = + ½

(f) n = 3, $l = 1$, $ml = 0$, $ms = + ½$

5. How many electrons in an atom may have the following quantum numbers?

(a) $n = 4, m=2, s = -½$ (b) $n = 3, l = 0$ (c) n=5, l = 2, m =1, s= +1/2, (d) n=3, l=3, s= +1/2

Exercise 4 :

1. Using complete subshell notation (no abbreviations), predict the electron configuration of each of the following atoms:

$_{6}C$, $_{15}P$, $_{23}V$, $_{51}Sb$, $_{62}Sm$.

2. Using short hand notation, predict the electron configuration of each of the following atoms:

$_{64}Cd$, $_{29}Cu$, $_{44}Ru$, $_{7}N$, $_{16}Si$, $_{52}Te$

3. Use an orbital diagram to describe the electron configuration of the valence shell of each of the following atoms: $_{7}N$, $_{16}Si$, $_{26}Fe$, $_{52}Te$, $_{42}Mo$
4. Which atoms have the electron configurations:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^2$, $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7 4s^2$

5. Which of the following atoms contains only three valence electrons: $_{3}Li$, $_{5}B$, $_{7}N$, $_{9}F$, $_{10}Ne$.
6. Which of the following has two unpaired electrons :$_{12}Mg$, $_{14}Si$, $_{16}S$; Both Mg and S ; Both Si and S.
7. Identify the element whose highest energy electron would have the following four quantum nombers :

(3.1.-1.+1/2) ; (4, 2, +1, +1/2); (6, 1, 0, -1/2) ; (2, 1, +1, -1/2)