Unit 5 – Questions - ANSWERS

Part 1 : Multiple-Choice questions

1) a	6) b	11) b
2) c	7) d	12) d
3) d	8) b	13) c
4) a	9) c	14) b
5) d	10) c	

Part 2 : Short-Answer Questions

15) The blue color of the sky results from the scattering of sunlight by air molecules. The blue light has a frequency of about 7.5×10^{14} Hz.

a. Calculate the wavelength, in nm, associated with this radiation.

$\nu = 7.5 \times 10^{14}$ Hz (=1/s)
$\lambda = ?$
$c = \lambda \cdot \nu$
3.00×10^8 m/s $= \lambda \cdot 7.5 \times 10^{14}$ s^{-}
$\lambda = 4.0 \times 10^{-7} \text{ m} \times \dfrac{1 \text{ nm}}{1 \times 10^{-9} \text{ m}} = \mathbf{4.0 \times 10^2 \text{ nm}}$

b. Calculate the energy, in joules, of a single photon associated with this frequency.

$E = h \cdot \nu$
$= (6.626 \times 10^{-34} \text{ J} \cdot \text{s})(7.5 \times 10^{14} \text{ s}^{-})$
$= 4.9695 \times 10^{-19}$ J / photon = $\mathbf{5.0 \times 10^{-19}}$ **J / photon**

c. Calculate the energy in kJ per mole of photons with this frequency.

$6.022 \times 10^{23} \text{ photons} \times \dfrac{5.0 \times 10^{-19} \text{ J}}{\text{photon}} \times \dfrac{1 \text{ kJ}}{1000 \text{ J}} = 301.1 \text{ kJ} = \mathbf{3.0 \times 10^2 \text{ kJ}}$

16) An energy of 3.3×10^{-22} kJ/atom is required to cause a cesium atom on a metal surface to lose an electron.

a. Calculate the longest possible wavelength of light that can ionize a cesium atom.

$E = h \cdot \nu$
$E = h \cdot (c / \lambda)$
$3.3 \times 10^{-22} \times 10^3 \text{ J} = \dfrac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{\lambda}$
$\lambda = 6.0 \times 10^{-7}$ m x 1 nm / 10^{-9} m = $\mathbf{6.0 \times 10^2}$ **nm**

b. In what region of the electromagnetic spectrum is this radiation?

Light with a wavelength of at least 600 nm, which is in the **yellow** region of the visible light spectrum.

17) The most prominent line in the spectrum of mercury is at 253.652 nm. Other lines are located at 365.015 nm, 404.656 nm, 435.833 nm, and 1013.975 nm.

a. Which of these lines represents the most energetic light?

The most energetic line is the one with the smallest wavelength. That would be **253.652 nm**.

b. What is the frequency of the most prominent line?

$\nu = c / \lambda$
$= 3.00 \times 10^8 \text{ m} \cdot \text{s}^{-} / 253.652 \times 10^{-9} \text{ m} = \mathbf{1.18282 \times 10^{15} \text{ s}^{-}}$

c. What is the energy of one photon with this wavelength?

$E = h \cdot \nu$

$= (6.626 \times 10^{-34}\ J{\cdot}s)(1.18282 \times 10^{15}\ s^{-})$

$\mathbf{= 7.836 \times 10^{-19}\ J/photon}$

18) Calculate the frequency (Hz) and wavelength (nm) of the photon that is emitted when an electron falls from the $n = 4$ to the $n = 2$ level in a hydrogen atom.

$$\frac{1}{\lambda} = R_\infty \cdot \left(\frac{1}{n_i^2} - \frac{1}{n_{out}^2}\right)$$

$$\frac{1}{\lambda} = 1.097 \times 10^{-2} nm^{-} \cdot \left(\frac{1}{2^2} - \frac{1}{4^2}\right)$$

$$\frac{1}{\lambda} = 0.0020569\ nm^{-}$$

$\boldsymbol{\lambda = 486\ nm}$

$$\nu = c / \lambda$$

$$= \frac{3.00 \times 10^8 m/s}{486 \times 10^{-9} m}$$

$\mathbf{= 6.17 \times 10^{14}\ s^{-}}$

19) Protons can be accelerated to speeds near that of light in particle accelerators. What is the wavelength (in nm) of such a proton moving at 2.90×10^8 m/s. (Mass of a proton = 1.673×10^{-27} kg)

$$1\ Joule = \frac{1\ kg{\cdot}m^2}{s^2}$$

$$h = 6.626 \times 10^{-34}\ J{\cdot}s$$

$$= 6.626 \times 10^{-34} \frac{(kg{\cdot}m^2)}{s^{\cancel{2}}}\cancel{s}$$

$$\lambda = h / (m \cdot v)$$

$$= \frac{6.626 \times 10^{-34}\ kg{\cdot}m^2}{\cancel{s}} \times \frac{\cancel{s}}{(1.673 \times 10^{-27}\ kg \cdot 2.90 \times 10^8\ m)}$$

$$= 1.37 \times 10^{-15}\ m \times \frac{1\ nm}{1 \times 10^{-9}\ m}$$

$\mathbf{= 1.37 \times 10^{-6}\ nm}$

20) a) Calculate the wavelength, in nanometers, associated with a 1.0×10^{-2}g golf ball moving at a speed of 30. m/s.

$$\lambda = h / (m \cdot v)$$

$$= \frac{6.626 \times 10^{-34}\ kg{\cdot}m^2{\cdot}s^{-}}{(1.0 \times 10^{-2} g \times 1\ kg/1000g)(30\ m/s)}$$

$$= 2.2 \times 10^{-30}\ m \times 1\ nm / 10^{-9}\ m$$

$\mathbf{= 2.2 \times 10^{-21}\ nm}$

b) At what speed must the ball travel to have a wavelength of 5.6×10^{-3} nm?

$$\lambda = \frac{h}{(m \cdot v)}$$

$$5.6 \times 10^{-3} nm \times \frac{1 \times 10^{-9} m}{1\ nm} = \frac{6.626 \times 10^{-34}\ kg{\cdot}m^2{\cdot}s^{-}}{\left(1.0 \times 10^{-2} g \times \frac{1\ kg}{1000\cancel{g}}\right) \times v}$$

$$5.6 \times 10^{-12}\ m \times v = 6.626 \times 10^{-29}\ m^2{\cdot}s^{-}$$

$\mathbf{v = 1.2 \times 10^{-17}\ m/s}$

21) An electron in an atom is in the *n* = 3 quantum level. List the possible values of ℓ and mℓ that it can have.

n = 3
ℓ = 0 → (n-1)
= 0, 1, 2
mℓ = - ℓ → + ℓ,
= -2, -1, 0, +1, +1
when ℓ = 0, mℓ = 0
ℓ = 1, mℓ -1, 0, +1
ℓ = 2, mℓ -2, -1, 0, +1, +2

22) Complete the following on quantum numbers:

a. When *n* = 4, what are the possible values of ℓ?
n = 4
ℓ = 0 → (n-1) *4 subshells*
= 0, 1, 2, 3

b. When ℓ is 2, what are the possible values of *m*ℓ?
ℓ = 2
mℓ = - ℓ → + ℓ,
= -2, -1, 0, +1, +2

c. For a 4*s* orbital, what are the possible values of *n*, ℓ, and *m*ℓ?
n = 4
ℓ = 0 → (n-1) = s, therefore = 0
mℓ = 0

d. For a 4*f* orbital, what are the possible values of *n*, ℓ, and *m*ℓ?
n = 4
ℓ = 0 → (n-1) = f, therefore = 3
mℓ = - ℓ → + ℓ = -3, -2, -1, 0, +1, +2, +3

23) What are the possible values of *m*ℓ for an electron with ℓ = 4?
ℓ = 4
mℓ = - ℓ → + ℓ = -4, -3, -2, -1, 0, +1, +2, +3, +4 *9 orbitals*

24) What is the difference between the following subshells and orbitals?

a. the 2*s* subshell and a 2*s* orbital
The s subshell contains only one orbital. The 2s orbital is the 2s subshell.

b. the 2*p* subshell and a 2*p* orbital.
The 2p subshell contains 3 orbitals:

-1	0	1
2px	2py	2pz

25) Answer these questions related to orbitals, shells, and electrons of atoms:

a. How many orbitals are there in the fourth shell of an atom?

ℓ (0 → n-1)	**mℓ** (-ℓ → + ℓ)	**# orbitals**
0 (s)	0	1
1 (p)	-1, 0, +1	3
2 (d)	-2, -1, 0, +1, +2	5
3 (f)	-3, -2, -1, 0, +1, +2, +3	7
	Total	**16**
	or $n^2 = 4^2 =$	**16**

b. The number of electrons per shell equals $2n^2$, where n is the shell number. How many electrons can be held in the fourth shell of an atom?
$2(n)^2 = 2(4)^2 = 32$
OR
Since each orbital can hold up to 2 electrons, and there are 16 orbitals in the 4^{th} shell, there are a total of 32 electrons.

c. How many electrons can be held in the fourth shell of an atom before the fifth shell starts to fill?
- Order for filling orbitals: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^{10}$ $4p^6$ $5s^2$
- The 5s subshell begins to fill before the 4d subshell. So, only the 4s and 4p subshells are filled before the 5^{th} shell starts.
- 4s subshell can hold up to 2 electrons
- 4p subshell can hold up to 6 electrons
- Total = 8 electrons

26) What are the permitted values of *ms* for an electron with $n = 6$, $l = 4$, and $ml = -2$?
No matter what values the other quantum numbers, ms is always = +1/2, -1/2

27) For each of the following sets of quantum numbers, briefly explain why they are not possible.
a. $n = 2$, $\ell = 2$, $m\ell = 0$
$\ell = 2$ is not possible because if n = 2, then $\ell = 0 \rightarrow (n - 1)$

b. $n = 3$, $\ell = 0$, $m\ell = -2$
$m\ell = -2$ is not possible because if $\ell = 0$, then $m\ell$ can only = 0 ($m\ell = -\ell \rightarrow +\ell$)

c. $n = 6$, $\ell = 2$, $m\ell = +3$
$m\ell = +3$ is not possible because if $\ell = 2$, then $m\ell$ can only = -2, -1, 0, +1, +2 ($m\ell = -\ell \rightarrow +\ell$)

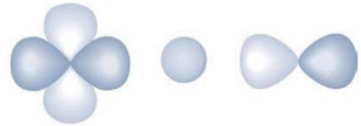
28) What is the maximum number of orbitals that can be identified by each of the following sets of quantum numbers? When "none" is the correct answer, explain your reasoning.
a. $n = 3$, $\ell = 0$, $m\ell = +1$
None because this set of quantum numbers is not possible.
$m\ell = -\ell \rightarrow +\ell$, so when $\ell = 0$, $m\ell$ can only equal 0.
b. $n = 5$, $\ell = 1$
Three orbital because $m\ell$ = -1, 0, +1
c. $n = 7$, $\ell = 5$
Eleven orbitals because $m\ell = -\ell \rightarrow +\ell$, and so = -5, -4, -3, … +3, +4, +5
d. $n = 4$, $\ell = 2$, $m\ell = -2$
$m\ell$ = -2, so 1 orbital

29) Fill in the blanks to complete the following statements.
a. The quantum number *n* describes the size and energy of an atomic orbital.
b. The shape of an atomic orbital is given by the quantum number _ ℓ _.
c. A photon of green light has more energy than a photon of orange light.
d. The maximum number of orbitals that may be associated with the set of quantum numbers $n = 4$ and $\ell = 3$ is _7_. *when $\ell = 3$, then there are 7 f orbitals:*

e. The maximum number of orbitals that may be associated with the quantum number set $n = 3$, $\ell = 2$, and $m\ell = -2$ is _1_. *when $m\ell=-2$, then this is specifying one orbital:*

-2	-1	0	+1	+2

f. Label each of the following orbital pictures with the appropriate letter:



d ($\ell = 2$)	s ($\ell = 0$)	p ($\ell = 1$)

g. When $n = 5$, the possible values of $\ell = 0 \rightarrow (n-1)$, so 0, 1, 2, 3, 4

h. The number of orbitals in the $n = 4$ shell is $n^2 = 4^4 = 16.$

30) What are the values of *n* and *l* in each of the following subshells?

a. 2*p*: n = 2, $\ell = 1$
b. 3*s:* n = 3, $\ell = 0$
c. 5*d:* n = 5, $\ell = 2$
d. 4*f*: n = 4, $\ell = 3$

31) What is the maximum number of electrons that can be identified with each of the following sets of quantum numbers? In some cases, the answer is "none." Explain why this is true.

a. $n = 4, \ell = 3, m\ell = 1$

Since $\ell = 3$, then f orbitals

$m\ell = 1$, then 1 f orbital and **2 electrons.**

-3	-2	-1	0	+1	+2	+3
				↑↓		

b. $n = 6, \ell = 1, m\ell = -1, ms = -1/2$

Since $\ell = 1$, then p orbitals

$m\ell = -1$

ms = -1/2, which specifies the spin number for **one electron**.

-1	0	+1
↓		

c. $n = 3, \ell = 3, m\ell = -3$

Since n = 3 and $\ell = 3$, then this doesn't exist. **No electrons**.

$\ell = 0 \rightarrow (n-1)$, so 0, 1, and 2 are possible values for ℓ.

32) (a) Depict the electron configuration for magnesium using an orbital box diagram and noble gas notation. (b) Give a complete set of four quantum numbers for each of the electrons beyond those of the preceding noble gas.

(a) magnesium: 12 electrons

$1s^2$	$2s^2$	$2p^6$			$3s^2$
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓
1s	2s	2p			3s

= [Ne] $3s^2$

(b) For the two electrons in the 3s orbital :

n = 3, $\ell = 0$, $m\ell = 0$, ms = +1/2

n = 3, $\ell = 0$, $m\ell = 0$, ms = -1/2

33) What is the maximum number of electrons in an atom that can have the following quantum numbers? Specify the orbitals in which the electrons would be found.

a. $n = 2, m_s = +1/2$

- n =2nd shell
- $m_s = +1/2$ specifies the spin of an electron.
- No other information is given regarding the subshell (ℓ) or the orbital orientation ($m\ell$).
- Since n = 2, $\ell = 0 \rightarrow (n-1) = 0, 1$:

↑

- $\ell = 0$ ___ 2s One electron in a 2s orbital

  or

- $\ell = 1$

↑	↑	↑

2p

One electron in each 2p orbitals.

b. $n = 4$, $m\ell = +1$

- n =4th shell
- $m\ell = +1$ specifies the orientation of the orbital
- No other information is given regarding the subshell (ℓ) or the spin number (m_s).
- Since $m\ell = +1$:
  - If $\ell = 1$, this means the p subshell.

			↑↓
$m\ell$	-1	0	+1

4p

    - Two electrons in a 4p orbital.

  - If $\ell = 2$, this means the d subshell.

				↑↓	
$m\ell$	-2	-1	0	+1	+2

4d

    - Two electrons in a 4d orbital.

  - If $\ell = 3$, this means the f subshell.

					↑↓		
$m\ell$	-3	-2	-1	0	+1	+2	+3

4f

    - Two electrons in a 4f orbital.

c. $n = 3$, $\ell = 2$

- n =3rd shell
- $\ell = 2$ specifies the subshell of the orbital = d
- No other information is given regarding the orientation ($m\ell$) or the spin number (m_s).
- Since $\ell = 2$:
  - $m\ell = -2, -1, 0, +1, +2$

↑↓	↑↓	↑↓	↑↓	↑↓
-2	-1	0	+1	+2

3d

  - Two electrons in each of the five 3d orbitals.

d. $n = 2$, $\ell = 0$, $m_s = -1/2$

- n = 2nd shell
- $\ell = 0$ specifies the subshell of the orbital = s
- $m_s = -1/2$ specifies the spin number
- It doesn't specify the orientation ($m\ell$), but there is only one option.
- Since $\ell = 0$, $m\ell = 0$:
  - One electron in a 2s subshell.

↓

34) The ground-state electron configurations listed below are **incorrect**. Explain what mistakes have been made in each and write the correct electron configurations.

a. Al: $1s^2 2s^2 2p^4 3s^2 3p^3$
Correct electron configuration for Al (atomic number 13): $1s^2 2s^2 2p^6 3s^2 3p^1$

b. B: $1s^2 2s^2 2p^5$
Correct electron configuration for B (atomic number 5): $1s^2 2s^2 2p^1$

c. F: $1s^2 2s^2 2p^6$
Correct electron configuration for F (atomic number 9): $1s^2 2s^2 2p^5$

35) Give the values of the four quantum numbers of an electron in the following orbitals:

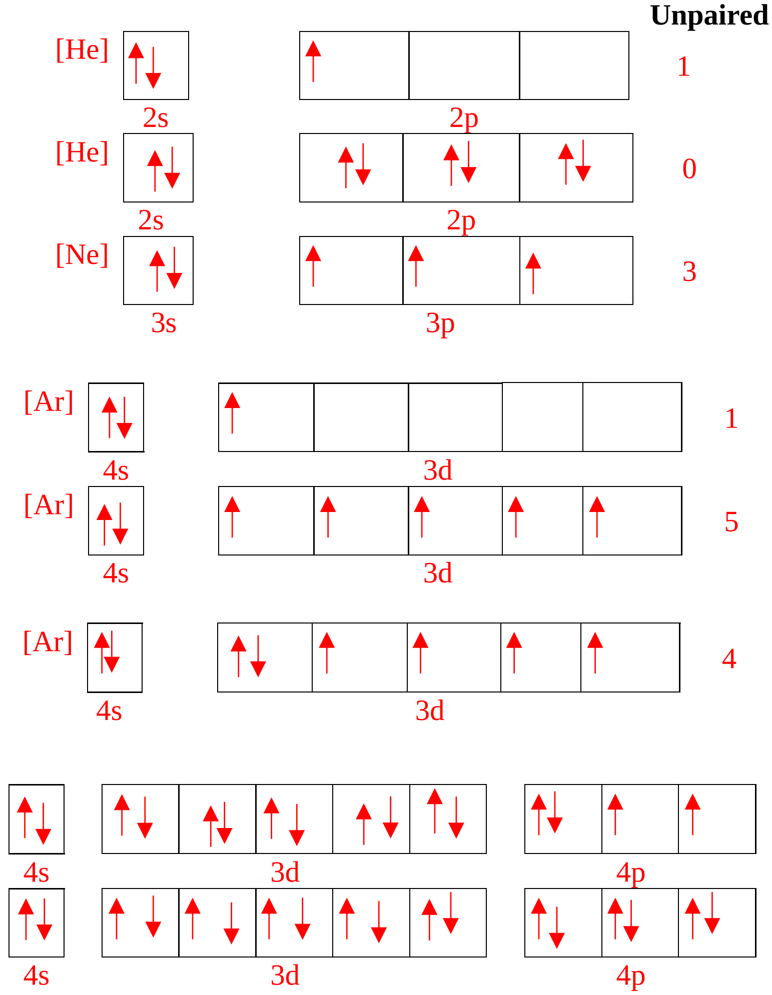
(a) 3*s* : n = 3, ℓ = 0, mℓ = 0

(b) 4*p* : n = 4, ℓ = 1, mℓ = -1, 0, +1

(c) 3*d*; n = 3, ℓ = 2, mℓ = -2, -1, 0, +1, +2

36) Indicate the number **of unpaired electrons** in each of the following atoms: B, Ne, P, Sc, Mn, Se, Kr, Fe.

	Unpaired e^-
B: [He] $2s^2$2p1	1
Ne: [He] $1s^2 2p^6$	0
P: [Ne] $3s^2 3p^3$	3
Sc: [Ar] $4s^2 3d^1$	1
Mn: [Ar] $4s^2 3d^5$	5
Fe: [Ar] $4s^2 3d^6$	4
Se: [Ar]$4s^2 3d^{10} 4p^4$	2
Kr: [Ar]$4s^2 3d^{10} 4p^6$	0



37) Write the detailed electronic configurations for K, S, and Y.

K	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p^6$			$4s^1$
$19e^-$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$

S	$1s^2$	$2s^2$	$2p^6$			$3s^2$	$3p^4$		
$16e^-$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$

Y	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$	$4p^6$	$5s^2$	$4d^1$
$39e^-$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow$	$\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$

38) Write detailed electronic configurations for N, P, As, and Sb. What makes their chemical properties similar?

N [He] $2s^2\ 2p^3$

P [Ne] $3s^2\ 3p^3$

As [Ar] $4s^2\ 3d^{10}\ 4p^3$

Sb [Kr] $5s^2\ 4d^{10}\ 5p^3$

Their outermost electronic configurations are similar: $ns^2\ np^3$

39) What neutral atom is represented by each of the following configurations?

a. $1s^22s^22p^63s^23p^4$

The superscripts equal the total number of electrons in this species. Since there are 16, this configuration is for sulfur (atomic number =16).

b. $1s^22s^22p^3$

Seven electrons, so this species is nitrogen.

c. $1s^22s^22p^63s^23p^64s^2$

Calcium

40) Write the short-hand electronic configuration for Iodine (atomic number = 53).

$\underbrace{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6}_{[Kr]\ (36\ electrons)}\ 5s^2\ 4d^{10}\ 5p^5$

Shortened form for I: [Kr] $5s^2\ 4d^{10}\ 5p^5$

$5s^2$	$4d^{10}$	$5p^5$
$\uparrow\downarrow$	$\uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow\ \uparrow\downarrow$	$\uparrow\downarrow\ \uparrow\downarrow\ \uparrow$

41) Using orbital box diagrams, depict an electron configuration for each of the following ions:

- Mg^{2+} Z = 12, lost 2 electrons → $10e^-$: $1s^2\ 2s^2\ 2p^6$
- O^{2-} Z = 8, gained 2 electrons → $10e^-$: $1s^2\ 2s^2\ 2p^6$
- K^+ Z = 19, lost 1 electron → $18e^-$: $1s^2\ 2s^2\ 2p^6\ 3s^23p^6$
- Cl^- Z = 17, gained 1 electron → $18e^-$: $1s^2\ 2s^2\ 2p^6\ 3s^23p^6$