

CHEM 101

Chapter 7 tutorial

7.55 An electron in a certain atom is in the $n = 2$ quantum level. List the possible values of L and m_L that it can have.

$n=2, l=0,1$,

if $l=0$ then $m_l=0$, if $l=1$ then $m_l = -1,0,+1$

7.56 An electron in an atom is in the $n = 3$ quantum level. List the possible values of L and m_L that it can have.

$n=3, l=0,1,2$,

if $l=0 \rightarrow m_l=0$, if $l=1 \rightarrow m_l = -1,0,+1$, if $l=2 \rightarrow m_l = -2,-1,0,+1,+2$)

7.57 Give the values of the quantum numbers associated with the following orbitals:

- (a) $2p$: $n = 2$, $l = 1$, $m_l = 1, 0, \text{ or } -1$
- (b) $3s$: $n = 3$, $l = 0$, $m_l = 0$ (only allowed value)
- (c) $5d$: $n = 5$, $l = 2$, $m_l = 2, 1, 0, -1, \text{ or } -2$

An orbital in a subshell can have any of the allowed values of the magnetic quantum number for that subshell. All the orbitals in a subshell have exactly the same energy.

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

- (b) $4p$: $n = 4$; $l = 1$; $m_l = -1, 0, 1$; $m_s = +1/2, -1/2$
- (c) $3d$: $n = 3$; $l = 2$; $m_l = -2, -1, 0, 1, 2$; $m_s = +1/2, -1/2$

7.59 Discuss the similarities and differences between a $1s$ and a $2s$ orbital.

A $2s$ orbital is larger than a $1s$ orbital. Both have the same spherical shape. The $1s$ orbital is lower in energy than the $2s$.

7.60 What is the difference between a $2p_x$ and a $2p_y$ orbital?

The two orbitals are identical in size, shape, and energy. They differ only in their orientation with respect to each other.

7.61 List all the possible subshells and orbitals associated with the principal quantum number n , if $n = 5$.

- $l = 0, 1, 2, 3, 4$. these correspond to the $5s, 5p, 5d, 5f$, and $5g$ subshells
- $l=0, m_l = 0$ (1 orbital)
- $l=1, m_l = -1, 0, +1$ (3 orbitals)
- $l=2, m_l = -2, -1, 0, +1, +2$ (5 orbitals)
- $l=3, m_l = -3, -2, -1, 0, +1, +2, +3$ (7 orbitals)
- $l=4, m_l = -4, -3, -2, -1, 0, +1, +2, +3, +4$ (9 orbitals)

7.62 List all the possible subshells and orbitals associated with the principal quantum number n , if $n = 6$.

- $l = 0, 1, 2, 3, 4, 5$ these correspond to the $6s$, $6p$, $6d$, $6f$, $6g$ and $6h$ subshells
- $l=0, m_l = 0$ (1 orbital)
- $l=1, m_l = -1, 0, +1$ (3 orbitals)
- $l=2, m_l = -2, -1, 0, +1, +2$ (5 orbitals)
- $l=3, m_l = -3, -2, -1, 0, +1, +2, +3$ (7 orbitals)
- $l=4, m_l = -4, -3, -2, -1, 0, +1, +2, +3, +4$ (9 orbitals)
- $l=5, m_l = -5, -4, -3, -2, -1, 0, +1, +2, +3, +4, +5$ (11 orbitals)

7.64 What is the total number of electrons that can be held in all orbitals having the same principal quantum number n ?

<u>n value</u>	<u>orbital sum</u>	<u>total number of electrons</u>
1	1	2
2	$1 + 3 = 4$	8
3	$1 + 3 + 5 = 9$	18
4	$1 + 3 + 5 + 7 = 16$	32
5	$1 + 3 + 5 + 7 + 9 = 25$	50
6	$1 + 3 + 5 + 7 + 9 + 11 = 36$	72

In each case the total number of orbitals is just the square of the n value (n^2). The total number of electrons is $2n^2$.

7.66 Indicate the total number of

(a) *p* electrons in N ($Z = 7$);

(b) *s* electrons in Si ($Z = 14$)

(c) *3d* electrons in S ($Z = 16$)

The electron configurations for the elements are

(a) N: $1s^2 2s^2 2p^3$ There are three *p*-type electrons.

(b) Si: $1s^2 2s^2 2p^6 3s^2 3p^2$ There are six *s*-type electrons.

(c) S: $1s^2 2s^2 2p^6 3s^2 3p^4$ There are no *d*-type electrons.

7.68 Why do the $3s$, $3p$, and $3d$ orbitals have the same energy in a hydrogen atom but different energies in a many-electron atom?

In the many-electron atom, the $3p$ orbital electrons are more effectively shielded by the inner electrons of the atom (that is, the $1s$, $2s$, and $2p$ electrons) than the $3s$ electrons. The $3s$ orbital is said to be more “penetrating” than the $3p$ and $3d$ orbitals. In the hydrogen atom there is only one electron, so the $3s$, $3p$, and $3d$ orbitals have the same energy.

7.69 For each of the following pairs of hydrogen orbitals, indicate which is higher in energy:

(a) $1s$, $2s$; (b) $2p$, $3p$; (c) $3d_{xy}$, $3d_{yz}$; (d) $3s$, $3d$; (e) $4f$, $5s$.

(a) $2s > 1s$

(b) $3p > 2p$

(c) equal

(d) equal

(e) $5s > 4f$

7.70 Which orbital in each of the following pairs is lower in energy in a many-electron atom?

(a) $2s, 2p$; (b) $3p, 3d$; (c) $3s, 4s$; (d) $4d, 5f$

(a) $2s < 2p$

(b) $3p < 3d$

(c) $3s < 4s$

(d) $4d < 5f$

7.75 Indicate which of the following sets of quantum numbers in an atom are unacceptable and explain why:

(a) $(1, 0, 1/2, 1/2)$, (b) $(3, 0, 0, +1/2)$, (c) $(2, 2, 1, +1/2)$,
(d) $(4, 3, -2, +1/2)$, (e) $(3, 2, 1, 1)$

(a) is wrong because the magnetic quantum number m_l can have only whole number values.

(c) is wrong because the maximum value of the angular momentum quantum number l is $n - 1$.

(e) is wrong because the electron spin quantum number m_s can have only half-integral values.

7.78 Indicate the number of unpaired electrons present in each of the following atoms: B, Ne, P, Sc, Mn, Se, Kr, Fe, Cd, I, Pb.

You should write the electron configurations for each of these elements to answer this question. In some cases, an orbital diagram may be helpful.

B: $[\text{He}]2s^2 2p^1$ (1 unpaired electron)

P: $[\text{Ne}]3s^2 3p^3$ (3 unpaired electrons)

Mn: $[\text{Ar}]4s^2 3d^5$ (5 unpaired electrons)

Kr: (0 unpaired electrons)

Cd: $[\text{Kr}]5s^2 4d^{10}$ (0 unpaired electrons)

Pb: $[\text{Xe}]6s^2 4f^{14} 5d^{10} 6p^2$ (2 unpaired electrons)

Ne: (0 unpaired electrons, Why?)

Sc: $[\text{Ar}]4s^2 3d^1$ (1 unpaired electron)

Se: $[\text{Ar}]4s^2 3d^{10} 4p^4$ (2 unpaired electrons)

Fe: $[\text{Ar}]4s^2 3d^6$ (4 unpaired electrons)

I: $[\text{Kr}]5s^2 4d^{10} 5p^5$ (1 unpaired electron)

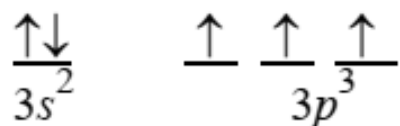
7.91 The electron configuration of a neutral atom is $1s^2 2s^2 2p^6 3s^2$. Write a complete set of quantum numbers for each of the electrons. Name the element.

There are a total of twelve electrons:

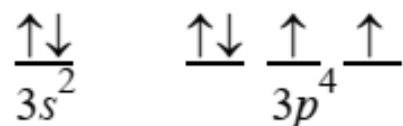
Orbital	n	l	m_l	m_s
1s	1	0	0	$+\frac{1}{2}$
1s	1	0	0	$-\frac{1}{2}$
2s	2	0	0	$+\frac{1}{2}$
2s	2	0	0	$-\frac{1}{2}$
2p	2	1	1	$+\frac{1}{2}$
2p	2	1	1	$-\frac{1}{2}$
2p	2	1	0	$+\frac{1}{2}$
2p	2	1	0	$-\frac{1}{2}$
2p	2	1	-1	$+\frac{1}{2}$
2p	2	1	-1	$-\frac{1}{2}$
3s	3	0	0	$+\frac{1}{2}$
3s	3	0	0	$-\frac{1}{2}$

The element is magnesium.

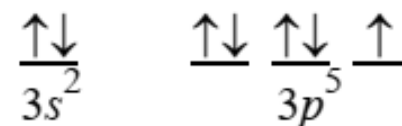
7.92 Which of the following species has the most unpaired electrons? S⁺, S, or S²⁻. Explain how you arrive at your answer



S⁺ (5 valence electrons)
3 unpaired electrons



S (6 valence electrons)
2 unpaired electrons



S²⁻ (8 valence electrons)
1 unpaired electron

S⁺ has the most unpaired electrons

7.96 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) $n = 2, m_s = +1/2$; (b) $n = 4, m_l = +1$; (c) $n = 3, l = 2$;

(d) $n = 2, l = 0, m_s = -1/2$; (e) $n = 4, l = 3, m_l = -2$.

- (a) With $n = 2$, there are n^2 orbitals $= 2^2 = 4$. $m_s = +1/2$, specifies 1 electron per orbital, for a total of **4 electrons**.
- (b) $n = 4$ and $m_l = +1$, specifies one orbital in each subshell with $l = 1, 2$, or 3 (i.e., a $4p$, $4d$, and $4f$ orbital). Each of the three orbitals holds 2 electrons for a total of **6 electrons**.
- (c) If $n = 3$ and $l = 2$, m_l has the values $2, 1, 0, -1$, or -2 . Each of the five orbitals can hold 2 electrons for a total of **10 electrons** ($2 e^-$ in each of the five $3d$ orbitals).
- (d) If $n = 2$ and $l = 0$, then m_l can only be zero. $m_s = -1/2$ specifies 1 electron in this orbital for a total of **1 electron** (one e^- in the $2s$ orbital).
- (e) $n = 4, l = 3$ and $m_l = -2$, specifies one $4f$ orbital. This orbital can hold **2 electrons**.