# 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)** $2p$: $n = 2$, $l = 1$, $m_l = 1, 0,$ or $-1$

**(b)** $3s$: $n = 3$, $l = 0$, $m_l = 0$ (only allowed value)

**(c)** $5d$: $n = 5$, $l = 2$, $m_l = 2, 1, 0, -1,$ 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 2*s* orbital is larger than a 1*s* orbital. Both have the same spherical shape. The 1*s* orbital is lower in energy than the 2*s*.

**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 5*s*, 5*p*, 5*d*, 5*f*, and 5*g* 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 6*s*, 6*p*, 6*d*, 6*f*, 6g and 6*h* 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*?

| *n* value | orbital sum | total number of electrons |
|---|---|---|
| 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 $\mathbf{2n^2}$.

**7.66** Indicate the total number of

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

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

(c) 3*d* 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 3*s*, 3*p*, and 3*d* orbitals have the same energy in a hydrogen atom but different energies in a many-electron atom?

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

**7.69** For each of the following pairs of hydrogen orbitals, indicate which is higher in energy:
(a) 1*s*, 2*s*; (b) 2*p*,3*p*; (c) 3*dxy*, 3*dyz*; (d) 3*s*, 3*d*; (e) 4*f*, 5*s*.

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

**(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: $[He]2s^2 2p^1$ (1 unpaired electron)

P: $[Ne]3s^2 3p^3$ (3 unpaired electrons)

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

Kr: (0 unpaired electrons)

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

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

Ne: (0 unpaired electrons, Why?)

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

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

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

I: $[Kr]5s^2 4d^{10} 5p^5$ (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? S1, S, or S2. Explain how you arrive at your answer

| $3s^2$ | $3p^3$ |
|---|---|
| ↑↓ | ↑ ↑ ↑ |

$S^+$ (5 valence electrons)
3 unpaired electrons

| $3s^2$ | $3p^4$ |
|---|---|
| ↑↓ | ↑↓ ↑ ↑ |

S (6 valence electrons)
2 unpaired electrons

| $3s^2$ | $3p^5$ |
|---|---|
| ↑↓ | ↑↓ ↑↓ ↑ |

$S^-$ (7 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 4*p*, 4*d*, and 4*f* 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 3*d* 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 2*s* orbital).

**(e)** $n = 4$, $l = 3$ and $m_l = -2$, specifies one 4*f* orbital. This orbital can hold **2 electrons**.