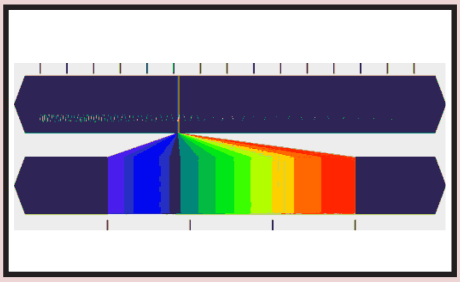
Chapter

8

Structure of Atom



You have studied in the previous class about sub-atomic particles like negatively charged electrons, positively charged protons and electrically neutral neutrons.

- How do these sub-atomic particles coexist in an electrically neutral atom?

You have acquired some fundamental ideas about atomic models suggested by J J Thomson, Ernest Rutherford and Niels Bohr which have been presented in Class 9.

Activity 1

On the basis of your previous knowledge can you prepare a model of an atom?

- Can you arrange the sub-atomic particles in any other way in an atom?

(Take the help of your friends, your teacher and the Internet). Observe your atomic model and the models of your friends carefully and try to answer the questions.

- Do all atoms have the same sub-atomic particles?
- Why is an atom of one element different from the atoms of other elements?
- How are the electrons distributed in the space of an atom?

To answer the above questions, we need to understand the nature of light, coloured flames and their characteristics.

Spectrum

You must have observed the formation of a rainbow.

- How many colours are there in a rainbow?

There are seven colours namely violet, indigo, blue, green, yellow, orange and red (VIBGYOR) in a rainbow.

You can find the colours spreading continuously and the intensity of each colour varies from one point to the other.

Wave nature of light

- When you throw a stone into a still pond, you observe ripples, which are transmitting the disturbance in the form of waves on the surface of water.
- You know that sound waves are produced when something vibrates; like a drum.
- In the same way electromagnetic waves are produced when an electric charge vibrates (moves back and forth).
- How do the vibrating electric and magnetic fields around the charge become a wave that travel through space?

A vibrating electric charge creates a change in the electric field. The changing electric field creates a changing magnetic field.

This process continues, with both the created fields being perpendicular to each other and at right angles to the direction of propagation of the wave.

Visible light is an electromagnetic wave and the speed of light (c) is 3×10^8 m s^{-1}.

- What are the characteristics of electromagnetic waves?

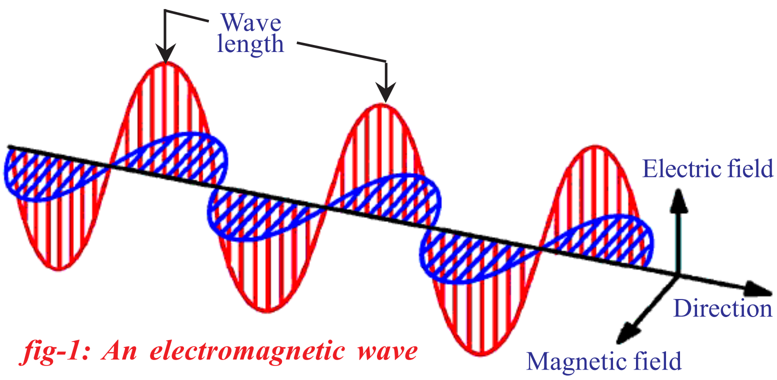


fig-1: An electromagnetic wave

Electromagnetic energy travelling through a vacuum behaves in some way like ocean waves travelling through water. Like ocean waves, electromagnetic energy is characterized by wavelength (λ), and frequency (υ).

The wavelength (λ) of the wave is the distance from one wave peak to the next. The frequency (υ) of a wave is simply the number of wave peaks that pass by a given point per unit time, expressed in units of reciprocal seconds (1/s or s^{-1}). The relation between these quantities is given by

$\lambda \alpha\ 1/\upsilon$ or $c = \upsilon\lambda$

- Can we apply this equation to a sound wave?

Yes. It is a universal relationship and applies to all waves. As the frequency increases, the wavelength becomes smaller.

Electromagnetic waves can have a wide variety of frequencies. The entire range of electromagnetic wave frequencies is known as the *electromagnetic spectrum*.

The familiar example of the visible spectrum in nature is the formation of a rainbow.

Each colour in a rainbow is characterized by a specified wavelength from red (higher wavelength) to violet (shorter wavelength). These colours (wavelengths), that the human naked eye is sensitive to, are called *visible light*. The range of wavelengths covering red colour to violet colour is called the *visible spectrum*.

- Are there any other wavelengths of light other than visible spectrum?

Electromagnetic spectrum

Electromagnetic waves can have a wide variety of wavelengths. The entire range of wavelengths is known as the electromagnetic spectrum.

The electromagnetic spectrum consists of a continuous range of wavelengths of gamma rays at the shorter wavelength to radio waves at the longer wavelength. But our eyes are sensitive only to visible light.

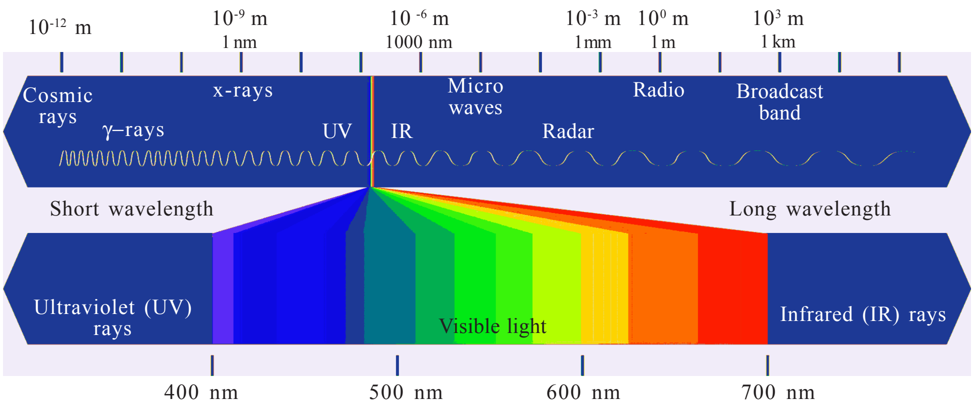


fig-2: Electromagnetic Spectrum

- What happens when you heat an iron rod on a flame?
- Do you find any change in colour on heating an iron rod?

When you heat an iron rod, some of the heat energy is emitted as light. First it turns red (lower energy corresponding to higher wavelength) and as the temperature rises it glows orange, yellow, blue(higher energy and of lower wavelength) or even white (all visible wavelengths) if the temperature is high enough.

- Do you observe any other colour at the same time when one colour is emitted?

When the temperature is high enough, other colours will also be emitted, but due to higher intensity of one particular emitted colour (e.g., red), others cannot be observed.

Max Planck broke with the 'continuous energy' tradition of electromagnetic energy by assuming that the energy is always emitted in multiples of $h\upsilon$;

For example: $h\upsilon$, $2\,h\upsilon$, $3\,h\upsilon$... $nh\upsilon$

That is, the energy for a certain frequency E can be represented by the equation $\mathbf{E} = \boldsymbol{h\upsilon}$**,** where 'h' is *Planck's constant* which has the value 6.626×10^{-34}Js and 'υ' is the frequency of the radiation absorbed or emitted.

The energy (E) for the red colour (higher wavelength or lower frequency) is lower compared to the energy of blue colour (lower wavelength or higher frequency). The energy emitted from a material body increases with increase in heat energy.

The significance of Planck's proposal is that, *electromagnetic energy can be gained or lost in discrete values and not in a continuous manner*.

Hence, emission or absorption of light spectrum is a collection of a group of wavelengths.

- Do you enjoy Deepavali fireworks?

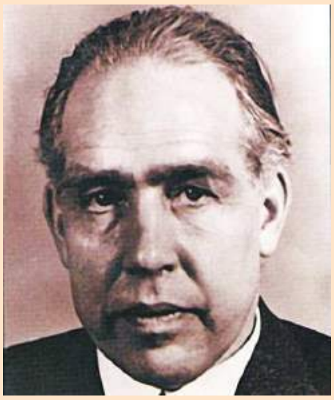
  Variety of colours is seen from fireworks.
- How do these colours come from fireworks?

Activity 2

Take a pinch of cupric chloride in a watch glass and make a paste with concentrated hydrochloric acid. Take this paste on a platinum loop and introduce it into a non-luminous flame.

- What colour do you observe?

  Carry out similar activity with strontium chloride.



Niels Henrik David Bohr was a Danish physicist who made foundational contributions to understanding atomic structure and quantum theory, for which he received the Nobel Prize in Physics in 1922. Bohr was also a philosopher and a promoter.

Cupric chloride produces a green colour flame while strontium chloride produces a crimson red flame.

- Do you observe yellow light in street lamps?

Sodium vapours produce yellow light in street lamps.

- Why do different elements emit different flame colours when heated by the same non-luminous flame?

Scientists found that each element emits its own characteristic colour. These colours correspond to certain discrete wavelengths of light and are called *line spectra*.

The lines in atomic spectra can be used to identify unknown atoms, just like fingerprints are used to identify people.

Bohr's model of hydrogen atom and its limitations

Let us examine the spectrum of hydrogen atom.

fig-3: Hydrogen Spectrum

- What does a line spectrum tell us about the structure of an atom?

Niels Bohr proposed that *electrons in an atom occupy 'stationary' orbits (states) of fixed energy at different distances from the nucleus.*

When an electron 'jumps' from a lower energy state (ground state) to higher energy states (excited state) it absorbs energy or emits energy when such a jump occurs from a higher energy state to a lower energy state.

The energies of an electron in an atom can have only certain values E_1, E_2, E_3 ……; that is, the energy is quantized. The states corresponding to these energies are called stationary states and the possible values of the energy are called energy levels.

- The lowest energy state of the electron is known as ground state.
- What happens when an electron gains energy?

  The electron moves to a higher energy level, the excited state.

- Does the electron retain the energy forever?

The electron loses the energy and comes back to its ground state. The energy emitted by the electron is seen in the form of electromagnetic energy and when the wavelength is in the visible region it is visible as an *emission line*.

Bohr's model explains all the line spectra observed in the case of hydrogen atom. It is a successful model as far as line spectra of hydrogen atom is concerned.

But the line spectrum of hydrogen atom when observed through a high resolution spectroscope appears as groups of finer lines.

- Did Bohr's model account for the splitting of line spectra of a hydrogen atom into finer lines?

  Bohr's model failed to account for splitting of line spectra.

Bohr-Sommerfeld model of an atom

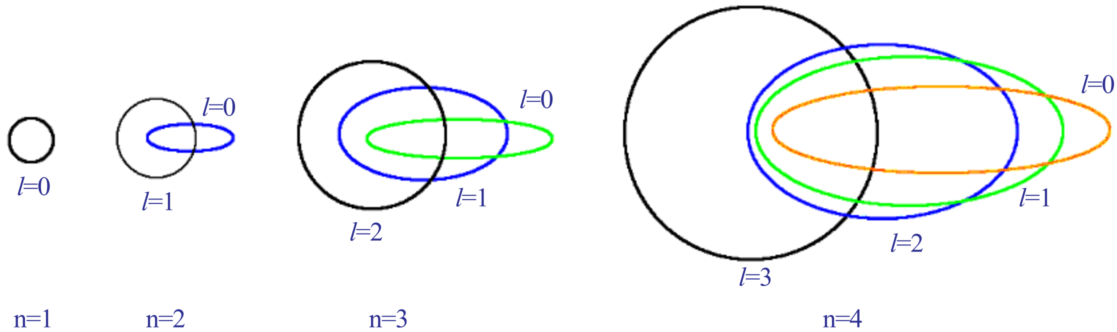


fig-4: The allowed electronic orbits for the main Quantum numbers by Bohr - Sommerfeld model

In an attempt to account for the structure (splitting) of line spectra known as fine spectra, Sommerfeld modified Bohr's atomic model by adding elliptical orbits. While retaining the first of Bohr's circular orbit as such, he added one elliptical orbit to Bohr's second orbit, two elliptical orbits to Bohr's third orbit, etc., such that the nucleus of the atom is one of the principal foci of these elliptical orbits. He was guided by the fact that, in general, periodic motion under the influence of a central force will lead to elliptical orbits with the force situated at one of the foci.

Bohr-Sommerfeld model, though successful in accounting for the fine line structure of hydrogen atomic spectra, does not provide a satisfactory picture of the structure of atom in general.

This model failed to account for the atomic spectra of atoms of more than one electron.

- Why is the electron in an atom restricted to revolve around the nucleus at certain fixed distances?

Max Karl Ernst Ludwig Planck was a German theoretical physicist who originated *quantum theory*, which won him the Nobel Prize in Physics in 1918. Planck made many contributions to theoretical physics, but his fame rests primarily on his role as originator of the *quantum theory*. This theory revolutionized human understanding of atomic and subatomic processes.

Quantum mechanical model of an atom

- Do the electrons follow defined paths around the nucleus?

If the electron revolves around the nucleus in defined paths or orbits, the exact position of the electron at various times will be known. For that we have to answer two questions:

- What is the velocity of the electron?
- Is it possible to find the exact position of the electron?

Electrons are invisible to naked eye. Then, how do you find the position and velocity of an electron?

To find articles during dark nights we take the help of torchlight. Similarly, we can take the help of suitable light to find the position and velocity of electron. As the electrons are very small, light of very short wavelength is required for this task.

This short wavelength light interacts with the electron and disturbs the motion of the electron. Hence, simultaneously the position and velocity of electron cannot be measured accurately.

From the above discussion, it is clear that electrons do not follow definite paths in an atom.

Do atoms have a definite boundary, as suggested by Bohr's model?

If the electrons are not distributed in orbits around the nucleus this means that an atom does not have a definite boundary.

As a result, it is not possible to pinpoint an electron in an atom.

Under these circumstances in order to understand the properties of electrons in an atom, a quantum mechanical model of atom was developed by Erwin Schrodinger.

According to this model of an atom, instead of orbits of Bohr's model, the electrons are thought to exist in a particular region of space around the nucleus at a given instant of time

- What do we call the region of space where the electron might be, at a given time?

The region of space around the nucleus where the probability of finding the electron is maximum is called an orbital.

In a given space around the nucleus, only certain orbitals can exist. Each orbital of a stable energy state for the electron is described by a particular set of quantum numbers.

Quantum numbers

Each electron in an atom is described by a set of three numbers n, l, and m_l .These numbers are called ***quantum numbers***. These numbers indicate the probability of finding the electron in the space around the nucleus.

- What information do the quantum numbers provide?

The quantum numbers describe the space around the nucleus where the electrons are found and also their energies. These are called atomic orbitals.

- What does each quantum number signify?

1. Principal Quantum Number (n)

The *principal* quantum number is related to the size and energy of the *main shell*.

'n' has positive integer values of 1, 2, 3,…

As 'n' increases, the shells become larger and the electrons in those shells are farther from the nucleus.

An increase in 'n' also means higher energy. n = 1, 2, 3, ... are often represented by the letters *K*, *L*, *M*... For each 'n' value there is one main shell.

Shell	K	L	M	N
n	**1**	**2**	**3**	**4**

2. The angular - momention quantum number (l)

'l' has integer values from 0 to n-1 for each value of 'n'. Each 'l' value represents one sub-shell.

Each value of 'l' is related to the shape of a particular sub-shell in the space around the nucleus.

The value of 'l' for a particular sub-shell is generally designated by the letters *s, p, d* . . . as follows:

l	0	1	2	3
Name of the sub-shell	s	p	d	f

When n = 1, there is only one sub-shell with l = 0. This is designated as '1*s*' orbital.

When n = 2, there are two sub-shells, with l = 0, the '2*s*' sub-shell and with l = 1, the '2*p*' sub-shell.

- What is the maximum value of 'l' for n=4?
- How many values can 'l' have for n = 4?

3. The magnetic quantum number (m_l)

The magnetic quantum number (m_l) has integer values between -l and l, including zero. Thus for a certain value of l, there are (2l +1) integer values of m_l as follows:

$-l, (-l+1) \ldots, -1, 0, 1, \ldots (+l - 1), +l$

These values describe the orientation of the orbital in space relative to the other orbitals in the atom.

When l = 0, (2l+ 1) = 1 and there is only one value of m_l, thus we have only one orbital i.e., 1s.

When l= 1, (2l+ 1) = 3, that means m_l has three values, namely, -1, 0, and 1 or three p orbitals, with different orientations along x, y, z axes. These are labelled as p_x, p_y, and p_z.

- Do these three p-orbitals have the same energy?

The number of 'm_l' values indicates the number of orbitals in a sub-shell with a particular l value. Orbitals in the sub-shell belonging to the same shell possess same energy.

Fill the table-1 with the number of orbitals per sub-shell using (2l+1) rule.

Table-1

l	Sub-shell	Number of orbitals
0	s	
1	p	
2	d	
3	f	

s-orbital is spherical in shape, **p**-orbital is dumbell-shaped and **d**-orbital are double dumbell-shaped as shown below.

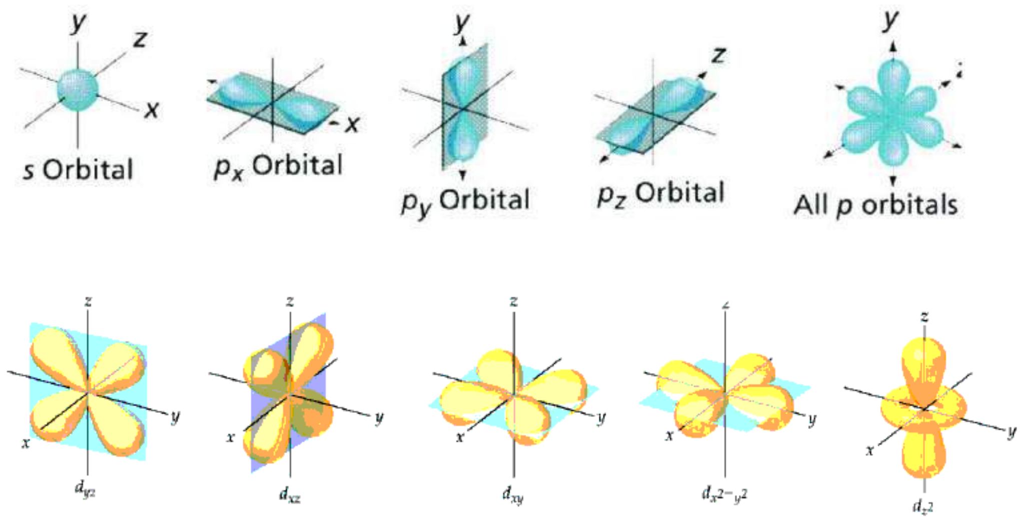


fig-5: Shapes of orbitals in s, p and d subshells

The following table-2 represents the shells, sub-shells and the number of orbitals in the sub-shells.

Table-2

n	l	m_l	sub-shell notation	No of orbitals in the subshell
1	0	0	1s	1
2	0	0	2s	1
	1	-1,0,+1	2p	3
3	0	0	3s	1
	1	-1,0,+1	3p	3
	2	-2,-1,0,+1,+2	3d	5
4	0	0	4s	1
	1	-1,0,+1	4p	3
	2	-2,-1,0,+1,+2	4d	5
	3	-3,-2,-1,0,+1,+2,+3	4f	7

Each sub-shell holds a maximum of twice as many electrons as the number of orbitals in the sub-shell.

The maximum number of electrons that can occupy various sub-shells is given in the following table.

Table-3

Sub shells	Number of orbitals ($2l+1$)	Maximum number of electrons
s (l=0)	1	2
p (l=1)	3	6
d (l=2)	5	10
f (l=3)	7	14

4. Spin Quantum Number (m_s)

The three quantum numbers *n, l,* and m_l describe the size (energy), shape, and orientation, respectively, of an atomic orbital in space.

As you have observed in the case of street lights (sodium vapour lamp), yellow light is emitted. This yellow light is comprised of a very closely spaced doublet when analyzed using high resolution spectroscope.

Alkali and alkaline earth metals show such type of lines.

To account for such a behavior of electron an additional quantum number is introduced. This is spin quantum number. This represents the property of the electron. It is denoted by 'm_s'.

This quantum number refers to the two possible orientations of the spin of an electron, one clockwise and the other anticlockwise spin. These are represented by +1/2 and -1/2. If both are positive values, then the spins are parallel otherwise the spins are anti-parallel.

The importance of the spin quantum number is seen when electrons occupy specific orbitals in multi-electron atoms.

- How do electrons in an atom occupy shells, sub-shells and orbitals?

The distribution of electrons in shells, sub-shells and orbital in an atom is known as *electronic configuration*.

Electronic Configuration

Let us first consider the hydrogen atom for understanding the arrangement of electrons, as it contains only one electron.

The shorthand notation consists of the principal energy level (n value), the letter representing sub-level (l value), and the number of electrons (x) in the sub-shell is written as a superscript as shown below:

$$nl^{x}$$

For the hydrogen (H) atom having atomic number (Z) = 1, the number of electrons is one, then the electronic configuration is $1s^1$.

Denotes the number of electrons in orbital

Denotes the principle quantum number — $1s^1$

Denotes the angular momentum quantum number

The electron configuration can also be represented by showing the spin of the electron.

For the electron in H, as you have seen, the set of quantum numbers is: $n = 1$, $l = 0$, $m_l = 0$, $m_s = ½$ or $– ½$.

$\boxed{\uparrow}$

For many-electron atoms, we must know the electron configuration of the atom. The distribution of electrons in various atomic orbitals provides an understanding of the electronic behavior of the atom and, in turn, its reactivity. Let us consider the helium (He) atom.

- Helium (Z=2) atom has two electrons.
- How are these two electrons arranged?

To describe the electronic configuration for more than one electron in the atom, we need to know three principles:

Those are the *Pauli Exclusion Principle, Aufbau principle* and *Hund's Rule.*

Let us discuss them briefly.

The Pauli Exclusion Principle

Helium atom has two electrons. The first electron occupies '1s' orbital. The second electron joins the first in the 1s-orbital, so the electron configuration of the ground state of 'He' is $1s^2$.

- Then the question is: What are the spins of these two electrons?

According to Pauli Exclusion Principle no two electrons of the same atom can have all four quantum numbers the same.

If n, l, and m_l are same for two electrons then m_s must be different. In the helium atom the spins must be paired.

Electrons with paired spins are denoted by '↑↓'. One electron has $m_s = +1/2$, the other has $m_s = -1/2$. They have anti-parallel spins.

- How many electrons can occupy an orbital?

The major consequence of the exclusion principle involves orbital occupancy. Since only two values of m_s are allowed, an orbital can hold only two electrons and they must have opposite spins.

Hence, the electronic configuration of helium atom is:

$$\boxed{\uparrow\downarrow}$$

$1s^2$

Aufbau Principle

As we pass from one element to another one of next higher atomic number, one electron is added every time to the atom.

The maximum number of electrons in any shell is '$2n^2$', where 'n' is the principal quantum number.

The maximum number of electrons in a sub-shell (s, p, d or f) is equal to $2(2l+1)$ where $l = 0, 1, 2, 3\ldots$ Thus these sub-shells can have a maximum of 2, 6, 10, and 14 electrons respectively.

In the ground state the electronic configuration can be built up by placing electrons in the lowest available orbitals until the total number of electrons added is equal to the atomic number. This is called the Aufbau principle (The German word "Aufbau" means "building up.").Thus orbitals are filled in the order of increasing energy.

Two general rules help us to predict electronic configurations.

1. Electrons are assigned to orbitals in order of increasing value of $(n+l)$.

2. For sub-shells with the same value of (n+l), electrons are assigned first to the sub-shell with lower 'n'.

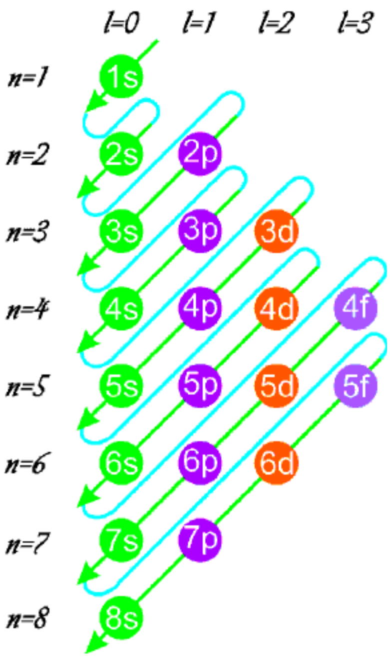


fig-6: The filling order of atomic orbitals (Moeller Chart)

The following diagram shows the increasing value of (n+l). Ascending order of energies of various atomic orbitals is given below.

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p < 8s …

The electronic configuration of some elements in the increasing atomic number (Z) value is given below.

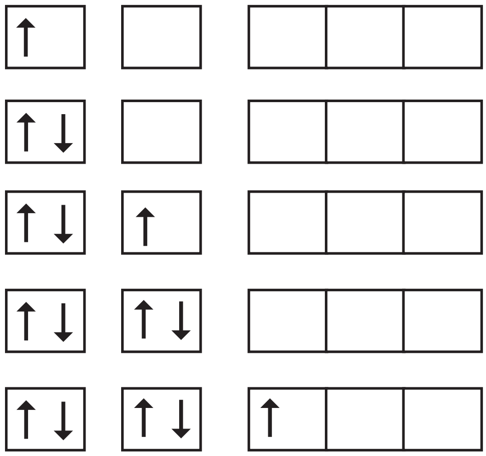
H(Z=1) $1s^1$

He(Z=2) $1s^2$

Li(Z=3) $1s^2 2s^1$

Be(Z=4) $1s^2 2s^2$

B(Z=5) $1s^2 2s^2 2p^1$

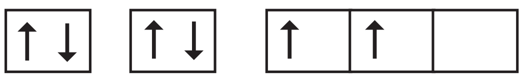


- For carbon (C) atom (Z=6), where does the 6th electron go?
- Whether the electron pairs up in the same *p*-orbital or will it go to the next *p*-orbital?

Hund's Rule

According to this rule electron pairing in orbitals starts only when all available empty orbitals of the same energy (*degenerate orbitals*) are singly occupied.

The configuration of Carbon (C) atom (Z=6) is $1s^2 2s^2 2p^2$. The first four electrons go into the 1*s* and 2*s* orbitals. The next two electrons go into separate 2*p* orbitals, with both electrons having the same spin



Note that the unpaired electrons in the 2*p* orbitals are shown with parallel spins.

Activity 3

Complete the electronic configuration of the following elements.

Table.4

Element	Atomic number (Z)	Electronic configuration of elements
C	6	
N	7	
O	8	
F	9	
Ne	10	
Na	11	
Mg	12	
Al	13	
Si	14	
P	15	
S	16	
C*l*	17	
Ar	18	
K	19	
Ca	20	

Key words

Wave, spectrum, intensity, discrete energy, line spectrum, orbital, quantum numbers, shell, sub-shell, shapes of orbitals electron spin, electronic configuration, the Pauli's exclusion principle, Aufbau principle, Hund's rule.

What we have learnt

- Light can be characterized by its *wavelength* (λ) and frequency (υ), and these quantities are related to the speed of light (c) as: $c = \upsilon\lambda$.
- Spectrum is a group of wavelengths.
- Electromagnetic energy (Light) can have only certain discrete energy values which is given by the equation $E=h\upsilon$
- Electrons in an atom can gain energy by absorbing a particular frequency of light and can lose energy by emitting a particular frequency.

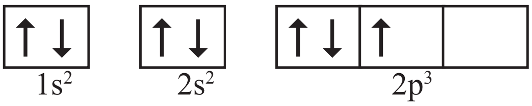
- Bohr's model of atom: Electrons are present in stationary states. The electron moves to higher energy level if it absorbs energy in the form of electromagnetic energy or moves to a lower energy state by emitting energy in the form of electromagnetic energy of appropriate frequency.
- Atomic line spectra arise because of absorption /emission of certain frequencies of light energy.
- It is not possible to measure accurately the position and velocity of an electron simultaneously.
- The space around the nucleus where the probability of finding the electron is maximum is called orbital.
- The three quantum numbers n, l, m_l describe the energy, shape and orientation respectively, of an atomic orbital.
- Spin is an intrinsic property of an electron.
- The arrangement of electrons in shells, sub-shells and orbitals in an atom is called the electron configuration.
- According to Pauli Exclusion Principle no two electrons of the same atom can have all the four quantum numbers same.
- Aufbau principle: The lowest-energy orbitals are filled first.
- Hund's rule: The orbitals of equal energy (degenerate) are occupied with one electron each before pairing of electrons starts.

Improve your learning

1. What information does the electronic configuration of an atom provide? (AS1)
2. a. How many maximum number of electrons that can be accommodated in a principal energy shell?
   b. How many maximum number of electrons that can be accommodated in a sub shell?
   c. How many maximum number of electrons can be accommodated in an orbital?
   d. How many sub shells present in a principal energy shell?
   e. How many spin orientations are possible for an electron in an orbital?
3. In an atom the number electrons in M-shell is equal to the number of electrons in the K and L shell. Answer the following questions. (AS1)
   a. Which is the outer most shell?
   b. How many electrons are there in its outermost shell?
   c. What is the atomic number of element?
   d. Write the electronic configuration of the element.
4. Rainbow is an example for continuous spectrum – explain.(AS1)
5. How many elliptical orbits are added by Sommerfeld in third Bohr's orbit? What was the purpose of adding these elliptical orbits?(AS1)
6. What is absorption spectrum?
7. What is an orbital? How is it different from Bohr's orbit?(AS1)
8. Explain the significance of three Quantum numbers in predicting the positions of an electron in an atom.(AS1)
9. What is nl^x method? How it is useful? (AS1)

10. Following orbital diagram shows the electron configuration of nitrogen atom. Which rule does not support this? (AS1)

N (Z = 7)



11. Which rule is violated in the electronic configuration $1s^0\ 2s^2\ 2p^4$?
12. Writ e the four quantum numbers for the differentiating electron of sodium (Na) atom? (AS1)
13. What is emisssion spectrum?
14. i. An electron in an atom has the following set of four quantum numbers to which orbital it belong to: (AS2)

n	l	m_l	m_s
2	0	0	+ ½

ii. Write the four quantum numbers for $1s^1$ electron.(AS1)

15. Which electronic shell is at a higher energy level K or L? (AS2)
16. Collect the information regarding wave lengths and corresponding frequencies of three primary colours red, blue and green. (AS4)
17. The wave length of a radio wave is 1.0m. Find its frequency.(AS7)

Fill in the blanks

1. If n = 1 then angular momentum quantum number (l) = ...
2. If a sub-shell is denoted as 2p then its magnetic quantum number values are ...., ..., ...
3. Maximum number of electrons that an M-shell contain is/are ....
4. For 'n', the minimum value is .... and the maximum value is ....
5. For 'l', the minimum value is ....and the maximum value is ....
6. For 'm_l' the minimum value is .... and the maximum value is ....
7. The value of 'm_s' for an electron spinning in clock-wise direction is .... and for anti-clockwise direction is ....

Multiple choice questions

1. An emission spectrum consists of bright spectral lines on a dark back ground. Which one of the following does not correspond to the bright spectral lines? [ ]
   a) Frequency of emitted radiation b) Wave length of emitted radiation
   c) Energy of emitted radiations d) Velocity of light
2. The maximum number of electrons that can be accommodated in the L – shell of an atom is :
   a) 2 b) 4 c) 8 d) 16 [ ]
3. If l = 1 for an atom then the number of orbitals in its sub-shell is [ ]
   a) 1 b) 2 c) 3 d) 0
4. The quantum number which explains about size and energy of the orbit or shell is: [ ]
   a) n b) l c) m_l d) m_s