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Atomatic structure (9 marks)

2 + 2 + 5
S.Q S.Q L.Q

~~Atoms~~ ~~are~~ is the smallest unit of element which can take part in chemical reaction. Atom is composed of 3 elementary particals called proton, electron & neutron. The word atom is derived from Greek word that is invisible which cannot be breakdown into small particles. In 1808 AD, John Dalton give the concept about the atom which is given below:

1) Every matter consist large number of tiny particles called atom.
2) Atoms can neither be created nor be destroyed by any chemical process.
3) The atoms of different element are differ in ~~(from)~~ size & mass.
4) The atom of same element is identical in nature

Sub-atomic particles

Atom consists of electron, proton, neutron, positron, mesons etc. These particals are sub-atomic particles. Now, it has been reported that 40 sub-atomic particles are found in atom. Among them, electron, proton & neutron are the fundamental particles.

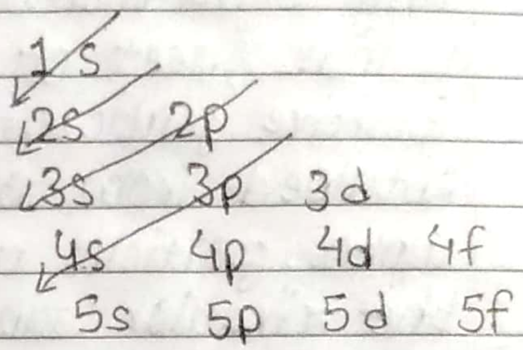
Aufbau principle

Aufbau is a ~~journal~~ german word which means building ~~house~~ or construction. It states that the process of filling of electron in an orbital from lower energy level to higher energy level.

The sequence of increasing order of energy of orbital is given below.

$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s$

- Ca → 20
  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
  s-block (metal)



- Ar → 18
  atomic number (Z) = 18
  $1s^2 2s^2 2p^6 3s^2 3p^6$
  p-block (non-metal)

- N → 7
  atomic number (Z) = 7
  $1s^2 2s^2 2p^3$

- ~~F~~ K → 19
  ato~~mat~~ic number (Z) = 19
  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

- Chromonium (Cr)
  Ato~~mat~~ic number (Z) = 24
  [Ar] $4s^1 3d^5$ [due to ~~external~~ extra stability]

copper

• $Cu \rightarrow 29$

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$

$[Ar]\ 4s^2\ 3d^9$ Aufbau rule (incorrect)

$[[Ar]\ 4s^1 3d^{10}]$ correct due to extra stability.

V.V.I

Rutherford's atomic model (alpha particle scattering experiment)

This model is put forward by Ernest Rutherford in 1911 with the help of an alpha particle scattering experiment. In this experiment, he bombarded a thin sheet of gold foil with alpha particles (positive particles equal to helium ion) which were obtained from a radioactive substance. The scattered alpha particles were observed on the surface of the circular zinc sulphide screen. Lead can absorb alpha particles, so lead plate with a slit was used to obtain a beam of alpha particles.

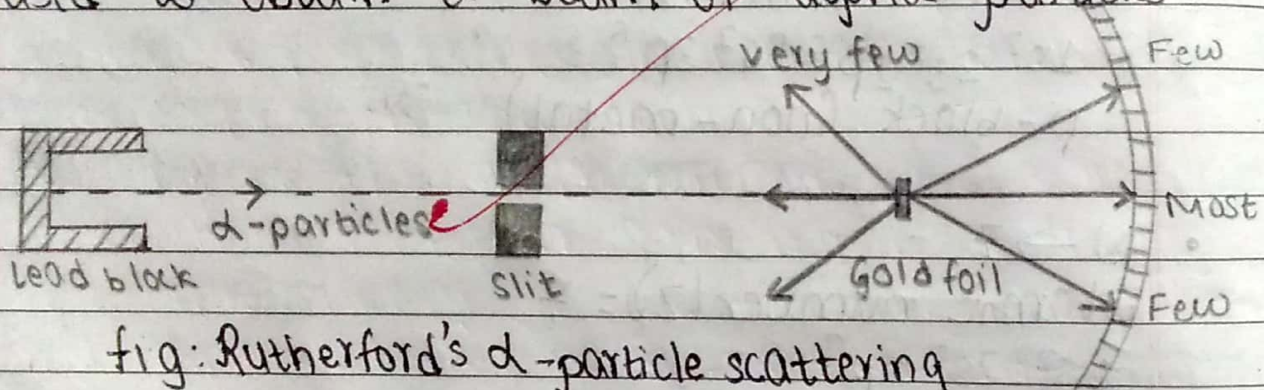


fig: Rutherford's α-particle scattering experiment

Observations

1. Most of the alpha particles passed through the gold foil without deflection.
2. Some of the alpha particles deflected through small angles.
3. Very few alpha particles were deflected through

angles more than 90° or bounced back.

Inference

1. Most of the space inside the atom is empty.
2. There is a presence of a heavy positively charged body at the centre of the atom.
~~3. Very few alpha particles were deflected through angles more than 90° or bounced back~~
3. There is a close encounter of alpha particles with a positively charged body.

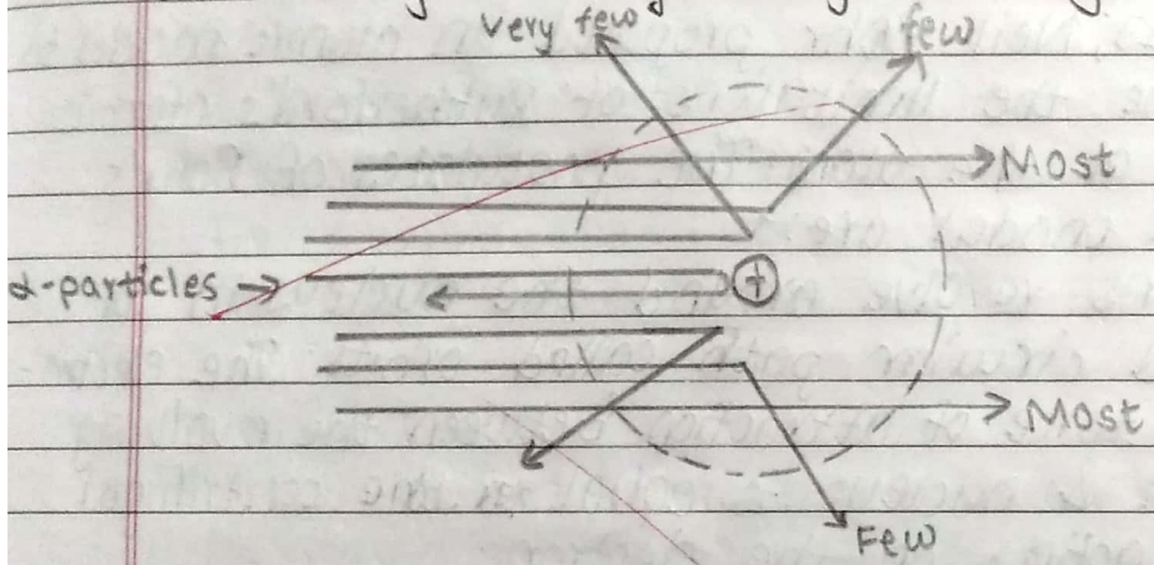


fig: scattering of alpha particles by a single atom

Postulates

1. An atom consists of a positively charged nucleus at which the entire mass is concentrated & its size is extremely small as compared to the size of the atom.
2. The space between the nucleus & revolving electron is empty.
3. The centrifugal force of the revolving electron is balanced by the electrostatic force of attraction between the electron and nucleus.

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Limitations

1. It ~~is~~ could not explain the stability of an atom

Bohr's atomic model

In 1913, Neil Bohr proposed an atomic model to overcome the limitations of Rutherford's atomic model of the atom. The postulates of Bohr's atomic model are:-

1. Electrons revolve around the nucleus in a defined circular path called orbits. The electro-static force of attraction between the revolving electron & nucleus is equal to the centrifugal force acting on the electron.
2. As long as the atom remains in a particular orbit, it will neither gain nor lose energy and hence the energy of the electron in a particular orbit remains constant. This means these orbits are non-radiating & thus called stationary state or energy level & designated as K, L, M, N etc.
3. Only those orbits are permitted in an atom whose angular momentum of the electron is equal to the whole number multiple of $h/2\pi$ where h is Plank's constant.

-Angular momentum $(mvr) = \frac{nh}{2\pi}$

where, m = mass of electron
v = velocity of electron
r = radius of orbit
$n = 1, 2, 3, \ldots$
h = Planck's constant $= 6.67 \times 10^{-27}$ erg

4. Energy is emitted or absorbed by the electron in the form of a photon only when it jumps from one energy level to another. The quantum or photon of energy absorbed or emitted is the difference between the higher & lower energy level.

$$\Delta E = E_2 - E_1 = h\nu$$

ΔE = energy emitted or absorbed
E_2 = higher energy level
E_1 = lower energy level

Energy is absorbed when an electron jumps from lower to higher energy level & energy is emitted when an electron jumps from higher to lower energy level.

release

Origin of hydrogen spectrum

When hydrogen gas is taken in a discharge tube & high voltage is applied, then there is the dissociation of hydrogen molecule into hydrogen atoms. Although hydrogen consists of only one electron, it produces different types of spectra. When energy is supplied, the electron of the hydrogen atom absorbs energy & jumps to a high energy level. The electrons in high energy level (excited state) are unstable & jump to lower energy level with the emission of electromagnetic radiation which give rise to line spectra called spectral series. The equation to calculate the wavelength of different spectral series is:

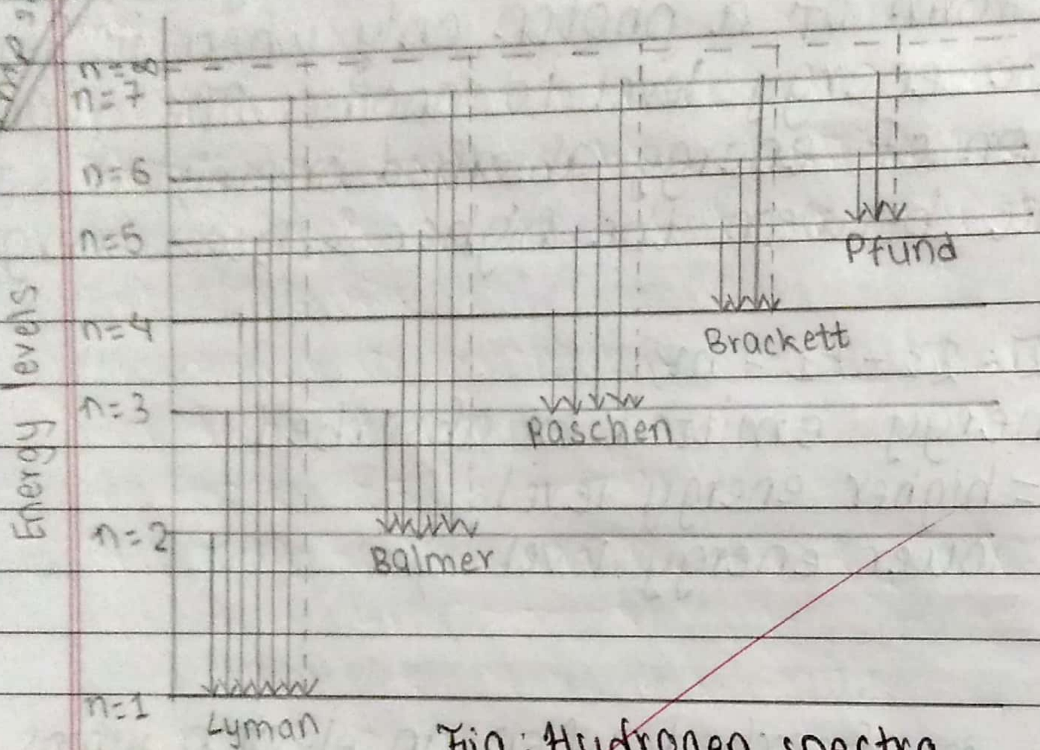

Fig: Hydrogen spectra

1. Lyman series (92-120 nm):
It is observed when an electron jumps from a higher energy level to the first energy level. It lies in the ultraviolet region.

2. Balmer series (400-650 nm):
It is observed when an electron jumps from a higher energy level to a second energy level. It lies in the visible region.

3. Paschen series (950-1875 nm):
It is observed when an electron jumps from a higher energy level to a third energy level. It lies

in the infrared region.

4. Brackett series (1945 - 4050 nm):
It is observed when an electron jumps from a higher energy level to a fourth energy level. It lies in the infrared region.

5. Pfund series (above 4050 nm):
It is observed when an electron jumps from a higher energy level to a fifth energy level. It lies in the infrared region.

Significances of Bohr's atomic model

1. It explains the stability of the atom.
2. It explains the origin of line spectra of the hydrogen atom.

Limitations of Bohr's atomic model

1. No explanation for the multi-electron system:
This model only explained the origin of spectra of the hydrogen atom and hydrogen-like ions like He^+, Li^{++}, etc. but it doesn't explain the origin of the spectra of the multielectron system.

Defects of Bohr's Atomic model

1) No explanation of spectrum of multielectronic system.
2) No explanation of dual nature of electron
Bohr's only treat electron as particles but De-broglie (1924) suggested that electrons behave as dual nature i.e wave & particles.

iii) No explanation about Zeeman effects & stark effect.

Bohr's unable to explain why spectral line get further split into thin line in presence of magnetic effect (zeeman effects) & electric field (starks effects).

Dual nature of electron (Debroglic equation)

Microscopic particles like electron, has dual nature i.e mass particle & v of light wave but it cannot behave both at sametime.

According to quantum theory, energy of photon is equal to hf

wave

$$E = hf \quad \text{——} \; (i)$$

and from Einstein mass - energy relationship:

$$E = mc^2 \quad \text{——} \; (ii)$$

Combining equation (i) & eq^n (ii)

constant $hf = mc^2$ momentum

$$h\frac{c}{\lambda} = mc^2$$

$$\frac{h}{\lambda} = mc$$

$$\therefore \frac{h}{\lambda} = p$$

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Quantum number (4 marks)

The number which gives position & energy of revolving electron in an atom is called quantum number.

Quantum number is further divided into
i) Principle quantum number (n)
ii) Azimuthal quantum number (l)
iii) Magnetic quantum number (m)
iv) Spin quantum number (s)

i) Principle quantum number (n)

This number gave the position of electron in different orbit. For example for K-shell n=1, for L-shell n=2, for M-shell n=3, for N-shell n=4 etc. For example, $_{11}Na$

K	L	M	N
2	8	1	

n=3

ii) Azimuthal quantum number (l)

Azimuthal quantum number describes the number of subshell that is orbital. The value of l lies from 0 to (n-1).

For n=1 l=0, to (1-1)=0 s-orbital
For n=2 l=0, to (2-1)=1 s-orbital, p-orbital
n=3 l=0, to (3-1)=2 s-orbital, p-orbital, d-orbital
n=4 l=0, to (4-1)=3 s-orbital, p-orbital, d-orbital, f-orbital

iii) Magnetic quantum number (m)

It describes about the orientation of orbital in three dimension space. It's value lies from -l to +l

$n=1$, $l=0$, $m=0$

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

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

4) Spin - quantum number (s)

Spin-quantum number describe the rotation of revolving electron in clockwise or anti-clockwise direction. If the value is $+\frac{1}{2}$ then it is clockwise rotation & if the value is $-\frac{1}{2}$ then it is anti-clockwise rotation.

V.IMP

1) What is quantum number? An atom has 20 electron find

i) It's atomic no. ii) no. of p-electron

iii) azumuthal quantum iv) position in periodic table

i) It's atomic no is 20.

ii) $Ca \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

no. of p-electron = 12

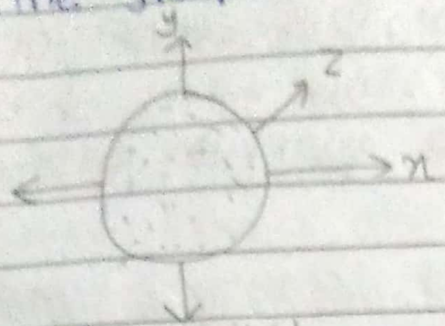
iii) Ca

K	L	M	N
2	8	8	2

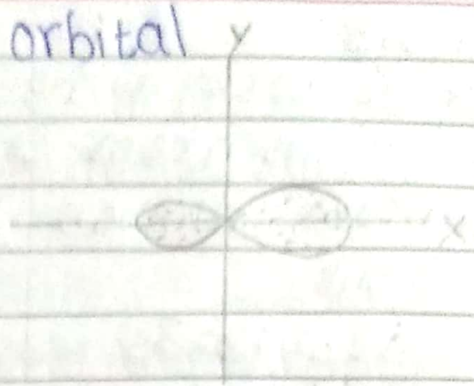
$n=4$ $l=0,1,2,3$

iv) position in periodic table = 4th period.

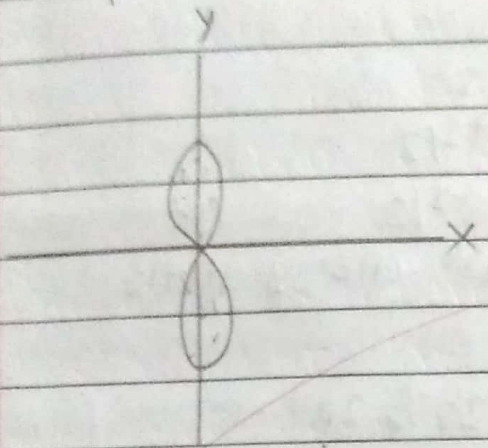
The shape of different orbital



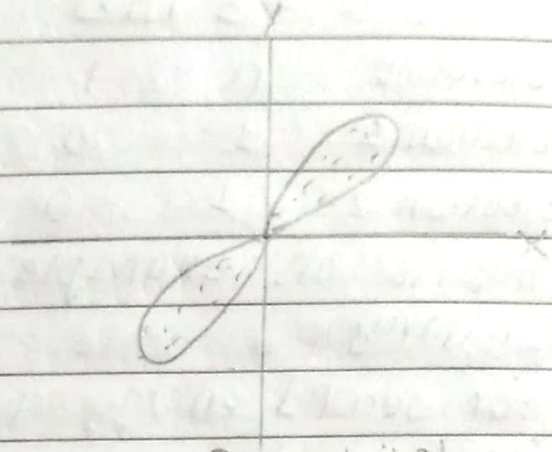
s-orbital
(spherical shape)



px orbital



Py orbital
(dumb best shape)



Pz orbital

Imp

Pauli's exclusion Principle

The distribution of quantum numbers among the electrons was enplain by Pauli's exclusion Principle. It states that "No two electrons in an atom can have same sets of 4 identical quantum numbers." It means that two electrons present in any orbitals can have same value of principle quantum numbers, subsidiary quantum numbers & same value of magnetic quantum numbers but spin quantum number is not same. One electron has $+1/2$ value & another electron has $-1/2$ value which indicate anti-clockwise & clock-wise direction respectively.

For eg:
Let us take a helium atom, two electrons present in helium atom can have same values of three quantum numbers but spin quantum number is different.

i.e. He = 3
$= 1s^2$
$= \boxed{\uparrow\downarrow}$

Electrons	n	l	m	s
Electron 1	1	0	0	$+1/2$
Electron 2	1	0	0	$-1/2$

This is an example of 1^{st} energy levels.
Similarly,
For second energy level (2s & 2p)

For 2s-orbital

n	l	m	s
2	0	0	$+1/2$
2	0	0	$-1/2$

For 2p-orbital

n	l	m	s	
2	1	-1	$+1/2$	For $2p_x$ orbital
2	1	-1	$-1/2$	
2	1	0	$+1/2$	For $2p_y$ orbital
2	1	0	$-1/2$	
2	1	+1	$+1/2$	For $2p_z$ orbital
2	1	+1	$-1/2$	

And so on.

Significance of Pauli's exclusion principle

1. It is used to determine or verify the numbers of electrons present in any orbitals or sub-shell or shell

2. Any orbitals contains only two electrons with anti-parallel spin but not parallel spin.

Question
1. State Pauli's experiment principle (2) with eg.
2. State & explain Pauli's exclusion principle. Also write its significance. (5)

Imp
Hund's rule OR Hund's rule of Maximum spin multiplicity
The filling of electrons in degenerate orbitals was explained by Hund's rule. It states that "while filling electrons in degenerate orbitals, at first electrons are filled singly then pairing takes place. Here, degenerate orbital means orbitals having same size, shape & energy but only differ direction in a space. p_x, p_y & p_z are the examples of degenerate orbitals in case of P sub shell.

Example of Hund's rule
Let us take a nitrogen atom. The filling of electrons in Nitrogen atom is

$$N = 7$$
$$= 1s^2, 2s^2\ 2p^3$$
$$= \boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \underset{(\checkmark)}{\boxed{\uparrow|\uparrow|\uparrow}}\ \left(\text{or } \underset{(\times)}{\boxed{\uparrow\downarrow|\uparrow|\ \ }}\right)$$

Similarly, the filling of electron in oxygen atom is:

$$O = 8$$
$$= 1s^2\ 2s^2\ 2p^4$$
$$= \boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \underset{(\checkmark)}{\boxed{\uparrow\downarrow|\uparrow|\uparrow}}\ \left(\text{or } \underset{(\times)}{\boxed{\uparrow\downarrow|\uparrow\downarrow|\ \ }}\right)$$

And so on.

Since, the maximum number of unpair electrons are obtained while filling electrons according to Hund's rule. So, this rule is also called Hund's rule of maximum spin multiplicity.

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Significance of Hund's rule

1. It is helpful to determine the paramagnetic & dimagnetic behaviour of an atom.

Questions:

1. What do you mean by Hund's rule of maximum spin multiplicity? (2 marks)
2. state & explain Hund's rule of maximum spin multiplicity? (2)
3. state Hund's rule. (2)
4. what is degenerate orbital? Write an example of it.

Bohr & Bury rule

To explain the arrangement of electrons in an atom having more than electron, Bohr & Bury proposed a new scheme for the distribution of electrons in an atom. The main points of the scheme are:-

The maximum number of electrons that can be accomodated by any shell or orbit is $2n^2$ rule where n is number of orbit. For eg:-

K-shell ($n=1$) can accomodate, $2 \times 1^2 = 2$ electrons.

L-shell

M-shell

N-shell

It is not necessary for an orbit to be completely filled before another orbit begins to fill. A new orbit begins when the outermost orbit attain eight electrons.