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Atomic structure

Inorganic chemistry deals with the physical and chemical properties of the elements of the periodic table.

An element

This is a substance that cannot be split into simpler substances. It is composed of discrete particles called atoms.

An atom is the smallest electrically neutral particle of an element that can take part in a chemical reaction.

The structure of the atom

The atom is composed of three basic subatomic particles, namely the electron, the proton and the neutron; the characteristics of which are given in a table1.1 below:

Table 1.1The three main subatomic particles.

Particle	Symbols	Charge	Mass
Electron	e	-1	1/1837
Proton	p	+1	1
Neutron	n	No charge	1

All atoms of the same element contain the same number of protons (usually equal to the number of electrons). The number of protons in an atom is characteristic of an element and is called the **atomic number**. **The atomic mass** is the sum of the number of protons and neutrons in an atom (element). Atoms with the same atomic number but different atomic masses (due to the difference in the number of neutrons in an atom) are called **isotopes**. It is now known that many more sub-atomic particles exist, e.g. the positron, the neutrino, the hyperon, etc., but in Chemistry only those listed in table 1.1. generally, need to be considered.

The atomic theory

An atom is composed of a nucleus (proton and neutrons) situated at the center of the atom and surrounded by a system of electrons. The electrons rotate around the nucleus in definite orbitals called energy levels or **quantum shells**; which are specified by giving them numbers, i.e., 1, 2, 3, 4, --------- or letters (capital) K, L, M, ------------ (fig. 1).

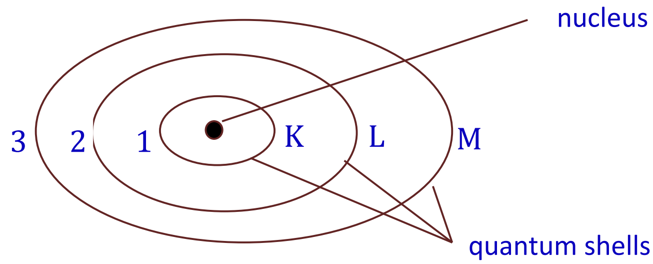


Fig. 1 Quantum shells

The numbers, 1, 2, 3 ----------- are called **principal quantum numbers**. Each energy level is associated with a definite amount of energy and the energy of these levels increases in the order K<L<M<N and an electron in any of these energy levels is associated with a definite quantity of energy.

If extra energy is supplied to an electron in ground state (this is usually done by heating, by collision with a fast-moving particle or by electrical discharge), the electron absorbs this extra energy and jumps to one of the higher energy levels. In this situation, the electron is said to be **excited** (fig. 2).

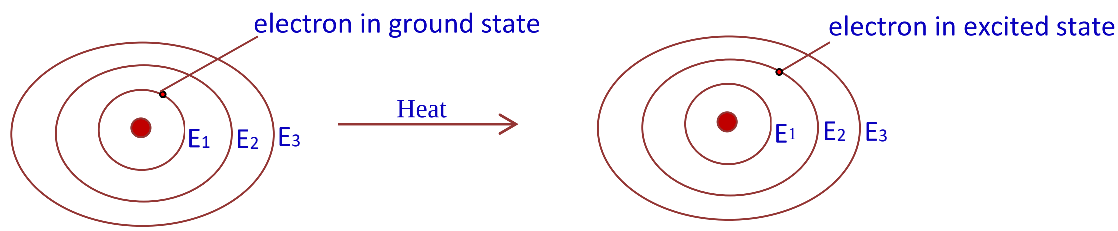


Fig 2 Excited state of an electron

If the electron falls back to the lower energy level, it emits energy in form of radiation. The energy absorbed or emitted is equal to the energy difference between the two levels.

Atomic spectra

A spectrum is a characteristic pattern of radiations absorbed or emitted by a substance.

Significance of atomic spectra

(ii) The fact that atomic spectra are line spectra suggests that only specific amounts of energy absorptions and emissions are possible indicating existence of energy levels within an atom and that an electron occupies certain energy levels around the nucleus. The various lines in each series are suggestive that the main energy levels are subdivided into lower energy levels.

(a) Absorption spectra

All atoms and molecules absorb light of certain wavelengths. When light is passed through a gaseous substance at low pressure, dark lines appear in the outgoing spectrum, where there were wavelengths of light that have been absorbed by the substance. The pattern of frequencies of light absorbed by the substance is called its **absorption spectrum**. The absorption spectrum is characteristic and can be used to identify a substance.

(b) Atomic emission or line spectrum

If atoms and molecules are heated to sufficiently high temperatures, they emit light of certain wave lengths. The pattern of frequencies of light emitted by a substance is called its **atomic emission or line spectrum**. Each line or color corresponds to a definite wavelength of radiation. If the frequency of the radiation is in the visible region of the radio magnetic spectrum; then it is seen as a colored light on a black background.

Application of atomic spectra

1. Identification of elements in salts by flame test
2. Street advertisement light
3. fireworks


(a)

(b)

Fig.3(a) A yellow flame of a sodium salt; (b) 'Streetfeet shoes'; a stylish use of atomic emission spectrum of rare gases in advertising

The atomic spectrum of hydrogen

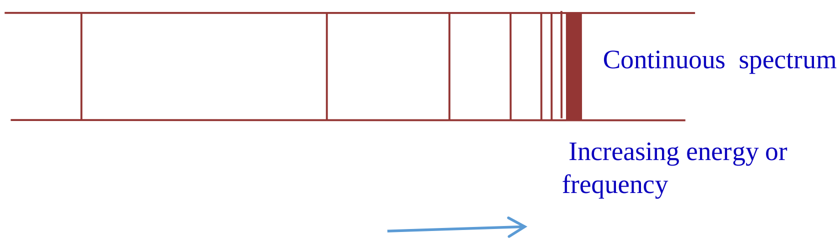
(i) Absorption spectrum

It is observed as a pattern of dark lines on a black background when white light is shone through gaseous hydrogen. The absorption spectrum is caused by hydrogen atoms absorbing energy corresponding to certain wavelengths from the light.

(ii) The hydrogen emission spectrum

It is observed as a pink glow when an electric discharge is passed through hydrogen at low pressure. Analyzed through a spectrometer, the emission is seen to be a number of separate sets of lines or series of lines.

In each series, the intervals between the frequencies of the lines become smaller and smaller towards the high frequency end of the spectrum until the lines run together or converge to form a continuum of light or continuum.



Cause of hydrogen spectra

When a hydrogen atom is struck by light, an electron absorbs energy of certain wavelength/frequency. The absorbed frequency appears as dark line in the absorption spectrum.

After absorbing energy, an electron jumps to a higher energy level. Energy absorbed is equal to the energy difference between the two energy levels.

When the excited electron drops to orbits of lower energy, it emits this energy in form of radiation giving an emission spectrum corresponding to the frequency of radiation emitted.

The discrete amount of energy (quantum) absorbed or emitted by an atom is proportional to the frequency of radiation. This energy is expressed as follows

$$E = hf \text{ (Plank's equation)}$$

Where E is the energy in joules, f, is the frequency of radiation (velocity of light/wavelength) and h is called Plank's constant. Its value is $6.625 \times 10^{-34} J\,s$

When the lines corresponding to a particular series are examined, they are seen that they all fit in the equation.

$$\frac{1}{\lambda} = \mathrm{RH}\left[\frac{1}{n^2} - \frac{1}{m^2}\right]$$

Where λ is the wavelength of a particular line, R^{H} is the Rydberg's constant and, *n* and *m* are integers, (*m* > *n*).

The complete atomic spectrum of hydrogen is resolved into five definite series characterized by one value of n for each series and with varying values of m. These five series named after their respective discoverers are shown in table 2 with their corresponding values of *n* and *m*. Electron transitions responsible for the different lines in each series are shown below (Table 2).

Table 2 Tabular representation of atomic hydrogen spectrum.

Lyman series	n = 1	m = 2, 3, 4, etc.	in the ultra violet region.
Balmer series	n = 2	m = 3, 4, 5, etc.	in the visible region
Paschen series	n = 3	m = 4, 5, 6, etc.	in the infrared region.
Brackett series	n = 4	m = 5, 6, 7, etc.	in the far infrared region.
Pfund series	n = 5	m = 6, 7, 8, etc.	in the far infrared region.

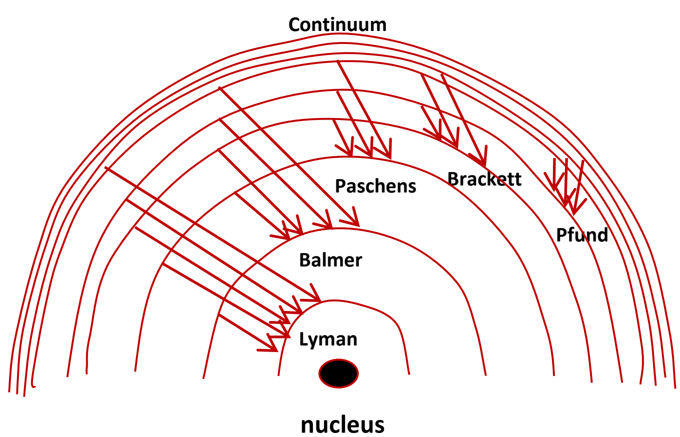


Fig. 4 The energy transitions among orbitals in each series.

If an electron can be made to jump beyond the highest energy level, it will not return to the atom and in that case, the atom is said to be **ionized;** i.e., the electron will no longer be bound to the atom. Therefore, the continuum represents complete removal of an electron from the atom. The energy required to excite an electron from its ground state above the highest energy level is the **ionization energy** for that electron.

Examples:

1. Calculate the wave number for the 1st and 3rd lines in the Lyman's series

(RH= 109678 cm^{-1})

Solution:

But for the Lyman's series *n*=1 and first line *m*=2

$$Wavenumber, \frac{1}{\lambda} = RH\left[\frac{1}{n^2} - \frac{1}{m^2}\right] = 109678\left[\frac{1}{1^2} - \frac{1}{2^2}\right] = 83359cm^{-1} \text{or } \lambda = 1.216 \times 10^{-5}cm$$

For the third line in Lyman's series m = 4

$$Wavenumber, \frac{1}{\lambda} = RH\left[\frac{1}{n^2} - \frac{1}{m^2}\right] = 109678\left[\frac{1}{1^2} - \frac{1}{4^2}\right] = 102823cm^{-1} \text{ or } \lambda = 9.725 \times 10^{-6}cm$$

2. Calculate the wave number for the line 3rd line in Balmer series. (RH = 109678 cm^{-1})

NB in Balmer series

lines	n	m
1st	2	3
2nd	2	4
3rd	2	5

Then

$Wavenumber, \frac{1}{\lambda} = R_H\left[\frac{1}{n^2} - \frac{1}{m^2}\right] = 109678\left[\frac{1}{2^2} - \frac{1}{5^2}\right] = 4387cm^{-1}$ or $\lambda = 2.3 \times 10^{-4}cm$

Trial 1

(a) Describe the spectra of a hydrogen atom.
Use a diagrams to illustrate your answer. (7marks)

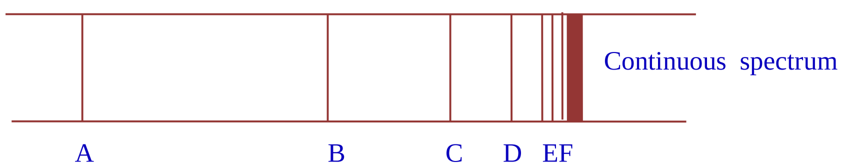
(b) Explain how the spectra of a hydrogen atom:
(i) are formed. (4marks)
(ii) provide evidence for existence of energy levels in atoms. (7marks)

(c) The frequency of hydrogen at the point of ionization is 32.8×10^{14} *Hz.*
Calculate the ionization energy of hydrogen.
(Plank's constant = $6.6 \times 10^{-34} J\ s$) (2marks)

Trial 2

The hydrogen emission spectrum consists of several series of lines. Each line represents the electron transitions between different energy levels characterized by different values of ***n***.

Part of the highest energy Lyman's series in the hydrogen atomic spectrum is shown in the figure below. A is the first line of this series.



(a)(i) On the diagram, draw an arrow to show the direction of increasing energy and label it "energy". (ii) Draw another arrow of increasing frequency and label it frequency.

(b) Why does the spectrum consist of lines?

(c) Name the symbol, ***n.***

(d) What do the transitions in the same series all have in common?

(e) When, if any, of the lines A to F shown above correspond to each of the following transitions? If none of the lines corresponds, answer 'none'. In each case, explain your answer.
(i) transition n=3 to n=1.
(ii) the transition n=3 to n=2.
(iii) the transition n=1 to n=3.

Subdivision of the main energy levels

The main energy levels of an atom designated by the principal quantum number, n = 1, 2, 3, 4… or by letters, K, L, M, N…, are further subdivided into sub- energy levels designed by s, p, d, f as follows;

Principal quantum shell	Subsidiary quantum shell
1	1s
2	2s 2p
3	3s 3p 3d
4	4s 4p 4d 4f
5	5s 5p 5d 5f
6	6s 6p 6d 6f
7	7s 7p 7d 7f

The relative energies of the sub shells (orbital) increase in the following order:
1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d.

This order of energy levels can easily be worked out from fig. 1.24.

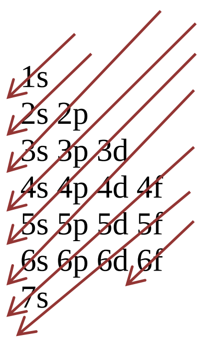


Fig. 1.24. Method for determining the order of energy levels

Electronic configurations

The electronic configuration (table below) of an element is constructed, by assuming that electrons occupy the lowest possible energy level, being available. The maximum number of electrons that can be accommodated in any given sub energy level is given in the table below

Sub energy level	Maximum electron it can carry
s	2
p	6
d	10
f	14

Table 5 The electronic configurations of the first ten elements

Element	Atomic number	Electronic configuration
H	1	$1s^1$
He	2	$1s^2$
Li	3	$1s^2 2s^1$
Be	4	$1s^2 2s^2$
B	5	$1s^2 2s^2 2p^1$
C	6	$1s^2 2s^2 2p^2$
N	7	$1s^2 2s^2 2p^3$
O	8	$1s^2 2s^2 2p^4$
F	9	$1s^2 2s^2 2p^5$
Ne	10	$1s^2 2s^2 2p^6$

The number of electron in the outermost shell determines **the group** of an element in the periodic table. For instance, oxygen has six electrons in the outermost shell (2 electrons in 2s-orbital and 4 electrons in2p-orbital) and is in group 6.

The number of filled quantum shell indicates **the period** in the periodic table into which an element falls. Oxygen atom has 6 electron in principal quantum shell 1 and 2; so it is in period 2

Trial 3

The electronic configuration of element X is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$.

(a) State with reason the period and group to which X belongs. (2marks)

(b) What is the valence of element X? Explain your answer. (1 ½ marks)

(c) Write electron configurations of atoms with the following atomic numbers; 24, 13 and 56.

Suggested answers to the trials

Trial 1

Hydrogen produces both absorption and emission spectrum

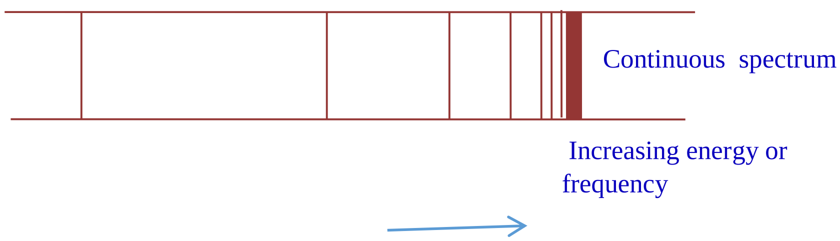
(i) Absorption spectra

It is observed as dark lines on a black background when white light is passed through gaseous hydrogen. This is caused by hydrogen atoms absorbing energy corresponding to certain wavelengths from the light.

(ii) Emission

It is observed as a pink glow when an electric discharge is passed through hydrogen at low pressure. Analyzed through a spectrometer, the emission is seen to be a number of separate sets of lines or series of lines.

In each series, the intervals between the frequencies of the lines become smaller and smaller towards the high frequency end of the spectrum until the lines run together or converge to form a continuum of light or continuum.



(b)(i) When a hydrogen atom in ground state is struck by light, its electron absorbs energy of certain wavelength/frequency. The absorbed frequency appears as a dark line in absorption spectrum

After absorbing energy, an electron jumps to a higher energy level. Energy absorbed is equal to the energy difference between two the energy levels. When the excited electron drops to the orbits of a lower energy level, it emits this energy in form of radiation giving, an emission spectrum corresponding to the frequency of radiation emitted.

(iii) The fact that the hydrogen spectrum is a line spectrum, suggests that only specific amounts of energy absorptions and emissions are possible suggesting existence of energy levels within an atom and that an electron occupies certain energy levels around the nucleus. The various lines in each series are suggestive that the main energy levels are subdivided into lower energy levels.

(iv) The spacing between adjacent lines in each series of the spectrum decreases with in the direction of decreasing wavelength showing that the series lines are produced from radiations of differing wavelengths that are emitted as a result of electron transitions from different energy levels.

(v) The existence of a continuum between the series of lines is suggestive of demarcation of energy levels around the nucleus and provides evidence for the existence of different energy levels.

(c) Energy = hf (h is Plank's constant, f = frequency)

$= 6.6 \times 10^{-34} \times 32.8 \times 10^{14}$

$= 2.165 \times 10^{-18}$J or 2.165×10^{-21} kJ

But 1 mole contains 6.023particles, according to Avogadro

$E = 2.165 \times 10^{-21} \times 6.0^{23} \times 10^{23} = 1304$ kJ

Trial 2

(a) (i) Arrow from left to right.
    (ii) Arrow from left to right.
(b) Because electron transitions produce radiations of definite energy.
(c) Principal quantum number.
(d) All involve transitions from higher energy levels down to the same energy level.
(e) (i) B
    (ii) None, transition to level 2 are Balmer series.
    (iii) None, an emission spectrum is caused by a transition from an orbital of a higher valve to another orbital of lower *n*.

Trial 3

(a) X is in period 4 because it has four energy levels/quantum shells and in group 1 because it contains only one electron in the outermost shell.
(b) X has a valence of +1 because it loses one electron to form a stable configuration.

(c) $24 = 1s^22s^22p^63s^23p^64s^13d^5$.
$13 = 1s^22s^22p^63s^23p^1$.
$56 = 1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^2$.