Introduction to

Atomic Physics

مقرر

Atomic Physics

الفرقة .....الثانية

شعبة... طبيعة وكيمياء

قسم الفيزياء – كلية العلوم بقنا

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Contents

PART ONE

ATOM STRUCTURE

Atoms

The word 'atom' comes from the Greek word 'a-tomio' which means 'uncuttable' or 'non-divisible'. Scientists believed that atoms were indivisible for the longest time. However, in the early 20th century, some scientists showed that atoms can be further divided into smaller parts such as electrons, protons, and neutrons. These are called sub-atomic particles. Want to know how these sub-atomic particles were discovered? Let's dive in!

Development of Atomic Theory

Democritus

The Greek philosopher Democritus (460 B.C. – 370 B.C.) was among the first to suggest the existence of atoms. Democritus explained the nature of matter. He also proposed that all substances are made up of matter. He stated atoms are constantly moving, invisible, minuscule particles that are different in shape, size, and temperature and cannot be destroyed.

- He believed that atoms were indivisible and indestructible
- His ideas did agree with later scientific theory, but did not explain chemical behavior, and was not based on the scientific method – but just philosophy

Greek Scientists

The idea of atomic theory has emerged in Greek society, and scientists gather that its founder, the Greek atomic theory was based on the following hypotheses:

- The atoms constantly move in the vacuum and affect each other by pushing and pressing.
- The material changes due to separation or contact of the atoms.
- Different materials can be explained by the shape and size of their constituent atoms.

Aristotle

In the fourth century BC Aristotle introduced a different concept of maize, refusing the idea of atom. Therefore, scientists believed that cheap materials (such as iron or copper) could be converted into precious materials (such as gold) by changing the proportions of the four components this belief prevailed for 2,000 years until 1600 AD.

Boyle

Boyle wrote that all matter is composed of solid particles arranged into molecules to give material its different properties.

He explained that all things are *made of one Catholick Matter common to them all, and…differ but in the shape, size, motion or rest, and texture of the small parts they consist of.* So, he refuses the concept of Aristotle and Set the first definition of element.

Element: *A simple pure substance that cannot be analyzed to the simplest of them by known chemical methods*.

Isaac Newton

Forty years later Isaac Newton expressed a typical 18th-century view of the atom that was similar to that of Democritus and Boyle. In the last query in his book *Opticks* (1704), Newton stated:

All these things being considered, it seems probable to me that God in the Beginning formed Matter in solid, massy, hard, impenetrable, moveable Particles, of such Sizes and Figures, and with such other Properties, and in such Proportion to Space, as most conduced to the End for which he formed them; and that these primitive Particles being Solids, are incomparably harder than any porous Bodies compounded of them; even so very hard, as never to wear or break in pieces; no ordinary Power being able to divide what God himself made one in the first Creation.

Atomic Models

In the era of Newton's particle theory, the perception of the shape of the atom was a small, solid, indivisible particle. Thus, 19th-century physicists accepted the idea that chemical elements were composed of atoms, but they did not know much about the atoms themselves.

All of these concepts have aided in the atomic structure in the theory of motion in gases. Electric discharge through gases has led to the discovery of the electron as one of the basic components of the atoms of all elements by the world Thompson, giving the first light to the atomic structure.

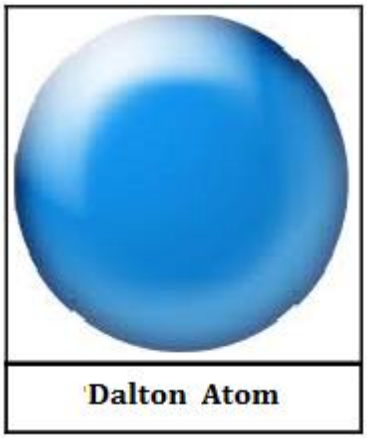
Dalton's Atomic Model

John Dalton (in 1805) proposes his Atomic Theory to explain the results of the quantitative studies of several scientists.

- Elements consist of tiny, indivisible particles called atoms.
- All the atoms of a given element are identical (e.g., same mass, same chemical behavior).
- The atoms of different elements differ in fundamental ways (e.g., different masses, different chemical behavior).
- Compounds form when atoms of different elements join together in simple whole number ratios. Thus, a given

compound always contains the same relative number and types of atoms.

- During a chemical reaction, atoms are neither created nor destroyed. Instead, reactions involve the reorganization of the atoms – a change in the way they are grouped together.
- The atoms themselves are unaltered.



Dalton Atom

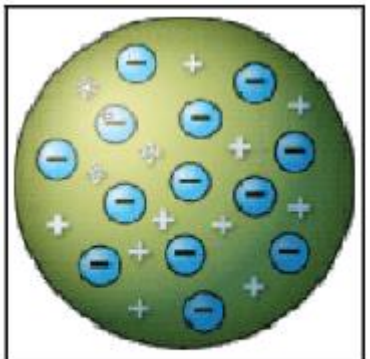
This was the first truly scientific theory of the atom, since Dalton reached his conclusions by experimentation and examination of the results in an empirical manner.

In the 19th century, physicists accepted the idea that chemical elements were composed of atoms, but they did not know much about the atoms themselves. The study of electric discharge led to the discovery of the electron as one of the basic components of the atoms of all elements, which shed light on the atomic structure. The electrons carry a negative electrical charge and the atom of the element is electrically neutral. This requires the atom to contain positive electrical charges sufficient to equal the negative charges of the electron. It was necessary to understand the shape of the atom. Where and how is this charge distributed? To answer this question, an American atomic scientist named J. C. Thomson is a simplified model of the 1898 atom,

Thomson's Atomic Model

The description of Thomson's atomic model is one of the many scientific models of the atom. It was proposed by J.J Thomson in the year 1904 just after the discovery of electrons. However, at that time the atomic nucleus was yet to be discovered. So, he proposed a model on the basis of known properties available at that time. The known properties are:

- According to the postulates of Thomson's atomic model, an atom resembles a sphere of positive charge with electrons (negatively charged particles) present inside the sphere.
- The positive and negative charge is equal in magnitude and therefore an atom has no charge as a whole and is electrically neutral (Atoms are neutrally charged).



Thomson Atom

- Thomson's atomic model resembles a spherical plum pudding as well as a watermelon. It resembles a plum pudding because the electrons in the model look like the dry fruits embedded in a sphere of positive charge just like a spherical plum pudding. The model has also been compared to a watermelon because the red edible part of a watermelon was

compared to the sphere having a positive charge and the black seeds filling the watermelon looked similar to the electrons inside the sphere.

- Thomson's atomic model failed to explain how the positive charge holds on the electrons inside the atom. It also failed to explain an atom's stability.
- The theory did not mention anything about the nucleus of an atom.
- It was unable to explain the scattering experiment of Rutherford.

SOLVED QUESTIONS FOR YOU

Question: Who gave the first model of an atom?

1. Rutherford **2.** J.J Thomson **3.** Eugen Goldstein
4. Neils Bohr

Solution: The answer is **2** (J.J. Thomson)

Ernest Rutherford

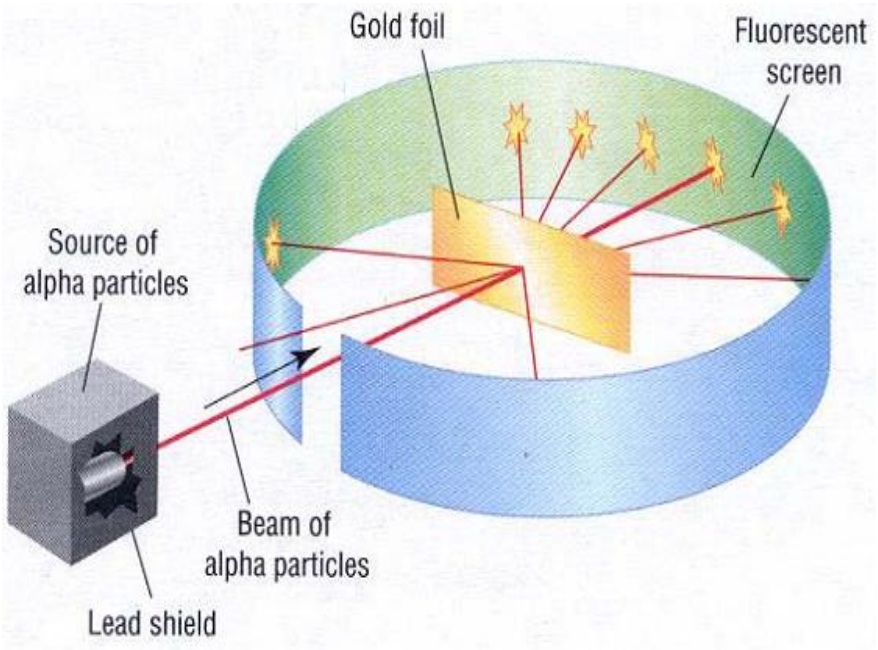
He Born in New Zealand (1871-1937) and tested Thomson's theory of atomic structure with the "gold foil" experiment in 1910 in University of Manchester, England.

Rutherford's Gold Foil Experiment

In Rutherford's experiment, he bombarded high energy streams of α-particles on a thin gold foil of 100 nm thickness. The streams of α-particles were directed from a radioactive source. He conducted the experiment to study the deflection produced in the trajectory of α-particles after interaction with the thin sheet of gold. To study the deflection, he placed a screen made up of zinc sulfide around the gold foil. The observations made by Rutherford contradicted the plum pudding model given by J.J. Thomson.

Rutherford and coworkers aimed a beam of alpha particles at a sheet of gold foil surrounded by a florescent screen.

Most particles passed through with no deflection, while some were highly deflected Rutherford concluded that most particles passed through because the atom is mostly empty space.

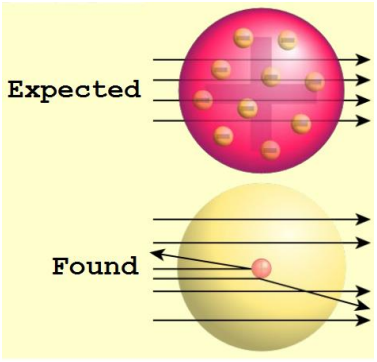


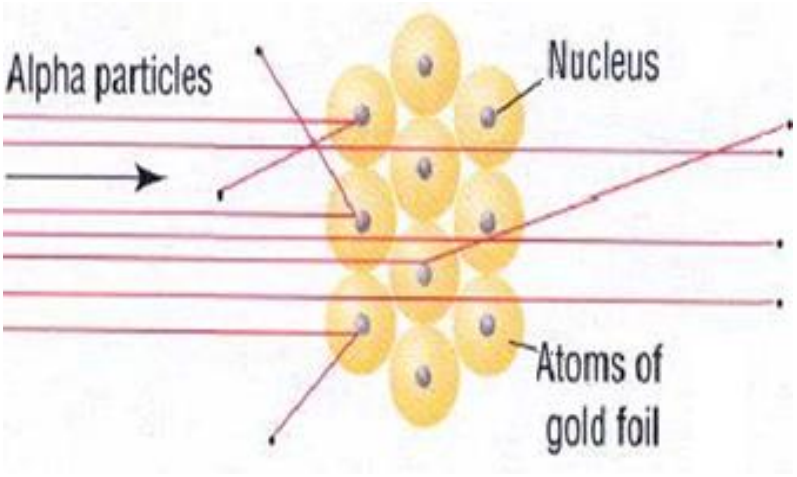
Observations of Rutherford Model Experiment

On the basis of the observations made during the experiment, Rutherford concluded that:

- Major space in an atom is empty – A large fraction of α-particles passed through the gold sheet without getting deflected. Therefore, the major part of an atom must be empty.
- The positive charge in an atom is not distributed uniformly and it is concentrated in a very small volume – Few α-particles when bombarded were deflected by the gold sheet. They were deflected minutely and at very small angles. Therefore he made the above conclusion.

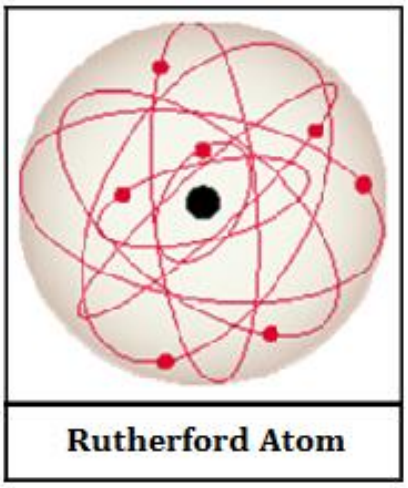
- Very few α-particles had deflected at large angles or deflected back. Moreover, very few particles had deflected at 180o. Therefore, he concluded that the positively charged particles covered a small volume of an atom in comparison to the total volume of an atom.





Postulates of Rutherford atomic model based on observations

- An atom is composed of positively charged particles. Majority of the mass of an atom was concentrated in a very small region. This region of the atom was called as the nucleus of an atom. It was found out later that the very small and dense nucleus of an atom is composed of neutrons and protons.
- Atoms nucleus is surrounded by negatively charged particles called electrons. The electrons revolve around the nucleus in a fixed circular path at very high speed. These fixed circular paths were termed as "orbits."
- An atom has no net charge or they are electrically neutral because electrons are negatively charged and the densely concentrated nucleus is positively charged. A strong electrostatic force of attractions holds together the nucleus and electrons.



Rutherford Atom

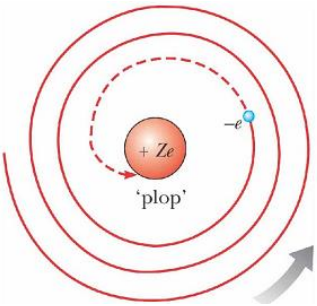
  - Electrons - negative charge- are attracted to the nucleus about which they orbit (just as planets orbit the sun due to attractive 1/r2 force)

- The size of the nucleus of an atom is very small in comparison to the total size of an atom.
  - nuclei ~ 10^{-14} m (calculated from fraction of α-particles that scatter more than 900 in a foil of given thickness)
  - atom ~ 10^{-10} m (from the mass density and number of atoms in a mole – Avogadro's number)

Difficulties of the Rutherford Model

Since ***electron*** travels in a circular orbit, it is constantly ***accelerated*** (even though its speed is constant.) Thus, the electron emits EM radiation, which carriers away energy. The energy of the atoms is reduced.

Thus the electrons has a lower potential energy and moves closer to the Nucleus Thus, classically, the Rutherford Atom is Unstable



Another problem is that the spectrum of the emitted EM radiation would be continuous. Classical approach gives the following expression

- As *r* decrease, the emission wavelength changes continuously, so this model predicts that the emission spectrum of atoms is broad

$$E_{total} = -\frac{(Ze)e}{8\pi\varepsilon_0 r}$$

- But sharp spectral lines are observed, not a continuum

Bohr's Model of Atom

Quantum mechanics, Quantum physics, the theory of relativity, etc are the modern subjects that interests, astound, and confuse almost everybody. These topics form the basis of modern physics. However, the very first-time quantum theory was ncorporated in Bohr's Model of an atom or Bohr atomic model. Later this model became the predecessor of complete quantum mechanical models.

The physicist Niels Bohr said, "Anyone who is not shocked by quantum theory has not understood it." He also said, "We must be clear that when it comes to atoms, language can only be used as in poetry." So what exactly is this Bohr atomic model? Let us find out!

Bohr atomic model and the models after that explain the properties of atomic electrons on the basis of certain allowed possible values. The model explained how an atom absorb or emit radiation when electrons on subatomic level jump between the allowed and stationary states. German-born physicists James Franck and Gustav Hertz obtained the experimental evidence of the presence of these states.

A Danish physicist named Neil Bohr in 1913 proposed the Bohr atomic model. He modified the problems and limitations associated with Rutherford's model of an atom. Earlier in Rutherford Model, Rutherford explained in an atom a nucleus is positively charged and is surrounded by electrons (negatively charged particles).

The electrons move around in a predictable path called orbits. Bohr modified Rutherford's model where he explained that electrons move around in fixed orbital shells. Furthermore, he explained that each orbital shell has fixed energy levels. Therefore, Rutherford basically explained a nucleus of an atom whereas Bohr took the model one step ahead. He explained about electrons and the different energy levels associated with it.

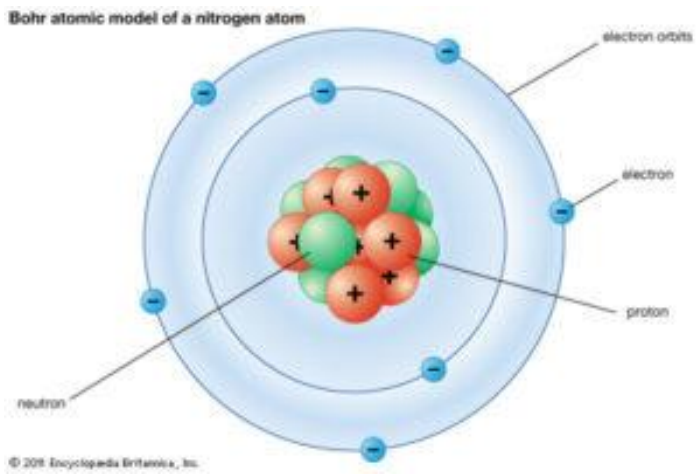
According to Bohr Atomic model, a small positively charged nucleus is surrounded by revolving negatively charged electrons in fixed orbits. He concluded that electron will have more energy if it is located away from the nucleus whereas the electrons will have less energy if it located near the nucleus.

SOLVED QUESTIONS FOR YOU

Question: Name the part of an atom discovered by Rutherford α-particles scattering experiment

1. Electrons **2.** Protons **3.** Neutrons **4.** Nucleus

Answer: The answer is 4. Rutherford α-particles scattering experiment led to the discovery of nucleus.



Bohr's Model of an Atom (Source Credit: Britannica)

Postulates of Bohr Atomic Model

- Electrons revolve around the nucleus in a fixed circular path termed "orbits" or "shells" or "energy level."
- The orbits are termed as "stationary orbit."
- Every circular orbit will have a certain amount of fixed energy and these circular orbits were termed orbital shells. The electrons will not radiate energy as long as they continue to revolve around the nucleus in the fixed orbital shells.
- The different energy levels are denoted by integers such as n=1 or n=2 or n=3 and so on. These are called as quantum numbers. The range of quantum number may vary and begin from the lowest energy level (nucleus side n=1) to highest energy level. Learn the concept of an Atomic number here.
- The different energy levels or orbits are represented in two ways such as 1, 2, 3, 4… or K, L, M, N….. shells. The lowest

energy level of the electron is called the ground state. Learn the concept of Valency here in detail.

- The change in energy occurs when the electrons jump from one energy level to other. In an atom, the electrons move from lower to higher energy level by acquiring the required energy. However, when an electron loses energy it moves from higher to lower energy level.

Therefore,

- 1st orbit (energy level) is represented as K shell and it can hold up to 2 electrons.
- 2nd orbit (energy level) is represented as L shell and it can hold up to 8 electrons.
- 3rd orbit (energy level) is represented as M shell and it can contain up to 18 electrons.
- 4th orbit (energy level) is represented as N Shell and it can contain maximum 32 electrons.

The orbits continue to increase in a similar manner.

Distribution of Electrons in Orbits or Shells:

Electronic distribution of various orbits or energy levels can be calculated by the formula 2n2. Here, 'n' denotes the number of orbits.

- The number of electrons in K shell (1st orbit) can be calculated by 2n2= 2 x 12 = 2. Thus, maximum number of electrons in 1st orbit = 2

- Similarly, The number of electrons in L shell (2nd orbit)= 2 x 22 = 8. Thus, maximum number of electrons in 2nd orbit = 8

We can determine the maximum number of electrons in a similar way.

Read about Thomson's Model of an Atom, the very first model of an Atom by J.J. Thomsons.

Limitations of Bohr's Model of an Atom:

Bohr atomic model had few limitations. They are:

- Failure to explain Zeeman Effect (how atomic spectra are affected by magnetic fields).
- It contradicts Heisenberg Uncertainty Principle.
- Unable to explain how to determine the spectra of larger atoms.

What are Isotopes? Learn the concept of Isotopes and Isobars.

Quantum Theory of Hydrogen Atom

The success of the Bohr model depends on the extent to which the results from its hypotheses are consistent with the results of the practical experiments. Here we will derive the theoretical relations based on the Bohr hypothesis and compare them with the practical results.

Find the radius and velocity of the electron around the nucleus

Let us assume that the mass of the electron is negligible for the mass of the nucleus and therefore we assume that the nucleus is fixed in the vacuum. Of the law of keeping the electron in orbit, it is under the influence of two forces equal in size and opposite in direction (Coulombic force and centripetal force).

$$\frac{1}{4\pi\varepsilon_o}\frac{Ze^2}{r^2}=\frac{mv^2}{r}$$

From which we find

$$Ze^2=4\pi\varepsilon_o mv^2 r \qquad \text{.....(3)}$$

where v is the speed of the electron in its orbit around the nucleus, and *r* is Orbit radius.

Using Bohr's second hypothesis of the electron's atomic momentum around the nucleus, we find that

$$L=\frac{nh}{2\pi}=n\left(\frac{h}{2\pi}\right)=n\hbar$$

As well

$$L = mvr$$

Thus it is

$$mvr = n\hbar$$

$$v = \frac{n\hbar}{mr}$$

$$v^2 = \frac{n^2\hbar^2}{m^2r^2} \qquad \text{..... ..... (4)}$$

replacing the value of v^2 form equation (4) in equation (3) we obtain the

$$Ze^2 = 4\pi\varepsilon_o m v^2 r$$

$$= 4\pi\varepsilon_o \frac{n^2\hbar^2}{mr}$$

Thus, the radius of the orbit

$$r = 4\pi\varepsilon_o \frac{n^2\hbar^2}{mZe^2} \qquad \text{..... ..... (5)}$$

The speed of the electron in orbit is

$$v = \frac{n\hbar}{mr} = \frac{1}{4\pi\varepsilon_o}\frac{Ze^2}{n\hbar} \qquad \text{..... .....(6)}$$

The quantification of the orbital angular momentum of the electron around the nucleus has quantified the possible orbits of the electron around the nucleus by determining the radius of the possible orbits of the electron that exists in it and finding that the radius of the orbit is directly proportional to the quantum number box. In equation (5) for n = 1, the radius value of the first orbit of the electron around the nucleus can be calculated for the

hydrogen atom where Z = 1, the radius is equal to $r = 5.3x10^{-11}\ m = 0.5\ A$

This value is acceptable to determine the radius of the atom and is within the value that Rutherford adopted in his model of the atom when he estimated the radius of the nucleus.

In equation (6), the velocity of the electron around the nucleus can be calculated in the first orbit n = 1 for the hydrogen atom. In order to compensate for the constants in the equation, the velocity of the electron is equal to

$$v = 2.2x10^{6}\ \ m/\sec$$

This is the largest electron speed around the nucleus because the velocity is inversely proportional to the quantum number of the orbit. When we talk about atoms that have an atomic number greater than the hydrogen atom Z> 1, the velocity becomes close to the speed of light. Here, the Bohr model is not verified for those atoms because it did not deal with speeds close to the speed of light.

Find the total energy of the electron in orbit around the nucleus

To calculate the total energy of the electron in any of the permitted orbits around the nucleus, we will collect the energy of the situation (resulting from the attraction between the positive charge charge and the negative electron charge assuming that the energy of the situation is zero when the electron is in the infinity) with the energy of the electron movement.

The potential energy:

$$V = -\frac{1}{4\pi\varepsilon_o}\frac{Ze^2}{r}$$

The negative signal of the energy of the situation indicates that the reciprocal force between the nucleus and the electron is a gravitational force and that there is a negative function to bring the electron from the infinity to its orbit around the nucleus.

The kinetic energy(K):

$$K = \frac{1}{2}mv^2$$

After compensation for the value of mv^2 *equation (3)*

$$K = \frac{1}{2}mv^2 = \frac{1}{2}\frac{1}{4\pi\varepsilon_o}\frac{Ze^2}{r}$$

Total energy is given from the equation

The total enery (E) :

$$\therefore \quad E = K + V =$$

$$= \left(\frac{1}{2}\frac{1}{4\pi\varepsilon_o}\frac{Ze^2}{r}\right) + \left(-\frac{1}{4\pi\varepsilon_o}\frac{Ze^2}{r}\right)$$

$$= \left(\frac{1}{4\pi\varepsilon_o}\frac{Ze^2}{r}\right)\cdot\left(\frac{1}{2}-1\right)$$

$$= -\frac{1}{4\pi\varepsilon_o}\frac{Ze^2}{2r}$$

To compensate for the value of r equation (5) in the equation of total energy we get:

$$E = -\left(\frac{Ze^2}{4\pi\varepsilon_o}\right)\cdot\frac{1}{2r}$$

$$= -\left(\frac{Ze^2}{4\pi\varepsilon_o}\right)\cdot\frac{mZe^2}{2(4\pi\varepsilon_o)n^2\hbar^2}$$

$$= -\frac{mZ^2e^4}{2(4\pi\varepsilon_o)^2 n^2\hbar^2}$$

$$= -\frac{1}{(4\pi\varepsilon_o)^2}\frac{mZ^2e^4}{2\hbar^2}\frac{1}{n^2} \qquad \text{..... ..... } (7)$$

Equation (7) shows energy quantification. The following diagram shows the quantified energy levels of the hydrogen atom. The values on the right of the diagram show the quantum number n and the values on the left side show the energy value corresponding to each energy level by equation (7). Note that the lowest level of energy is the level with the smallest quantum number n = 1 and the greater the n the total energy is less negative and the total energy is equal to zero from n to infinity.

The lowest level of energy is the most stable for the electron, the level n = 1 in the case of hydrogen atom.

The electron in the normal state is at the lowest level of energy and in the hydrogen atom is at the level n = 1. Thus, to extract the electron from the nucleus of the hydrogen atom, we must overcome the energy of its correlation with the nucleus. The energy of its level and the electron release makes the atom positively charged. . To calculate the energy of the electron in the first level, we compensate for n = 1 in equation (7) ie :

$$E = -2.17x10^{-18}\ j = -13.6\ \ eV$$

This is the energy value of the first level, the energy of the electron bonding in the nucleus, which is called Binding energy. The energy of the other levels can be calculated based on the energy value in the first level of the following relationship:

$$E_n = -\frac{13.6\ \ eV}{n^2}$$

SOLVED QUESTIONS FOR YOU

Question: Calculate the charge of one mole of electrons.

Solution: 1 mole = 6.022×10^{23} atoms, charge of one electron = 1.6022×10^{-19} C Therefore, charge of one mole of electrons = $1.6022 \times 10^{-19} \times 6.022 \times 10^{23} = 9.65 \times 10^{4}$C.

Find the frequency of electromagnetic radiation resulting from electron transmission between energy levels

Pohr's fourth hypothesis is that the electromagnetic spectrum is emitted from the atom when the electron moves from the ni orbit to the nf orbit as follows:

$$h\upsilon = E_i - E_f$$

To compensate for both Ei and Ef using equation (7) we obtain the frequency as follows:

$$\upsilon = \frac{1}{h}\left(E_i - E_f\right)$$
$$= \left(\frac{1}{2\pi\hbar}\right)\cdot\frac{1}{\left(4\pi\varepsilon_o\right)^2}\frac{mZ^2e^4}{2\hbar^2}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$$
$$= \frac{1}{\left(4\pi\varepsilon_o\right)^2}\frac{mZ^2e^4}{4\pi\hbar^3}\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$$

Using an inverted wavelength (wave number) k

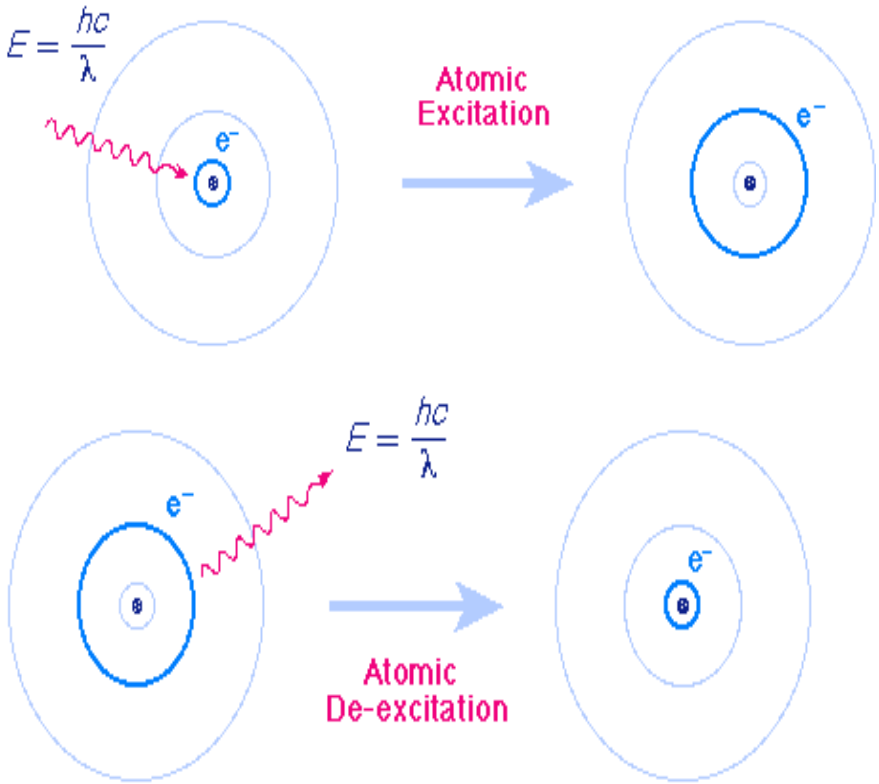
$$\because \nu = \frac{c}{\lambda} \quad , k = \frac{1}{\lambda} \quad \Rightarrow \quad \therefore k = \frac{\nu}{c}$$

$$k = \frac{1}{\lambda} = \left(\frac{1}{4\pi\varepsilon_o}\right)^2 \frac{me^4}{4\pi\hbar^3 c} Z^2\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right] \qquad \ldots\ldots\ldots(8)$$

Equations (7) and (8) are the most important conclusions of the Bohr model. Using these equations, the electromagnetic spectrum emitted from the hydrogen atom can be explained.

(1) The first energy level n = 1 is the lowest energy level called the ground state .

(2) When a hydrogen atom is stirred by, for example, electric discharge, the electron at level n = 1 will acquire energy from the collisions and move to a higher energy level (n> 1). Here the hydrogen atom becomes excited state.



(3) The atom is excreted from the excitation state by emitting a photon that carries the energy difference between the two levels. The transfer can be carried out through a series of transitions until reaching n = 1 level and in every fun move to a lower power level. For example, if the electron is raised to the level n = 7 then

it moves to n = 4 then moves to n = 2 and then to n = 1, in which case we obtain three spectral lines with a wavelength that can be calculated from equation (8) For n_i =7 and n_f =4 for the first spectral line and the second line for n_i = 4 and n_f = 2 and the third line for n_i =2 and n_f = 1.

SOLVED QUESTIONS FOR YOU

Question: Which of the options represent the properties of electrons?

1. The relative charge of -1
2. negligible mass
3. The charge is opposite and equal to a proton
4. all of these

SOLVED QUESTIONS FOR YOU

Question: What is the net charge of an atom?

Answer: There is no net charge of an atom. Electrons are the negatively charged particle whereas protons are the positively charged particles. The equal positive charge of the proton and the negative charge of the electron cancel each other. Therefore, the atom has no net charge. In an atom that is neutral, the number of electrons revolving around the nucleus and the number of protons inside the nucleus are equal in number.

Important Properties of Sub-atomic Particles

Name	Symbol	Absolute charge/C	Relative charge	Mass/kg	Mass/u	Approx. mass/u
Electron	e	-1.6022×10^{-19}	-1	9.10939×10^{-31}	0.00054	0
Proton	p	$+1.6022 \times 10^{-19}$	+1	1.67262×10^{-27}	1.00727	1
Neutron	n	0	0	1.67493×10^{-27}	1.00867	1

CSOLVED QUESTIONS FOR YOU

Question: If the atomic number of sodium is 11. Find out how many electrons and protons are present in a calcium atom.

Solution: We know that "Atomic Number = No. of Protons = No. of Electrons" Thus, number of electrons=11 and number of protons =11

Atomic Number

We know that an atom consists of electrons, protons, and neutrons. Atomic number is one of the fundamental properties of an atom. Each atom can be characterized by a unique atomic number. It is represented by the letter "Z."

The total number of protons present in the nucleus of an atom represents the atomic number of a particular atom. Every atom of a particular element is composed of the same number of protons and therefore have the same atomic number. However, atoms of different elements have unique atomic numbers that vary from one element to the other.

An atom does not have any net charge and is thus electrically neutral. This means that the number of electrons will be equal to the number of protons present in an atom thereby making an atom electrically neutral.

"Atomic Number = No. of Protons = No. of Electrons"

For example, each atom of oxygen has 8 protons in their nucleus so the atomic number is 8. Similarly, the atomic number of carbon is 6 because all atoms of carbon have 6 protons in their nucleus.

Atomic numbers are whole numbers because it is the total number of protons and protons are generally units of matter. It ranges from 1 to 118. It starts with hydrogen and ends with the heaviest known element Oganesson (Og).

Theoretically, the atomic numbers can be increased if more elements are discovered. However, with the addition of more number of protons and neutrons, the elements become prone to radioactive decay.

Importance of Atomic Numbers

- Helps in the identification of a particular element of an atom.
- Forms the basis of the arrangement of the elements. The elements are arranged in increasing order of the atomic numbers of the elements.
- Helps in the determination of the properties of any element. However, valence electron determines the chemical bondingbehaviourr of an element.

Examples of Atomic Numbers

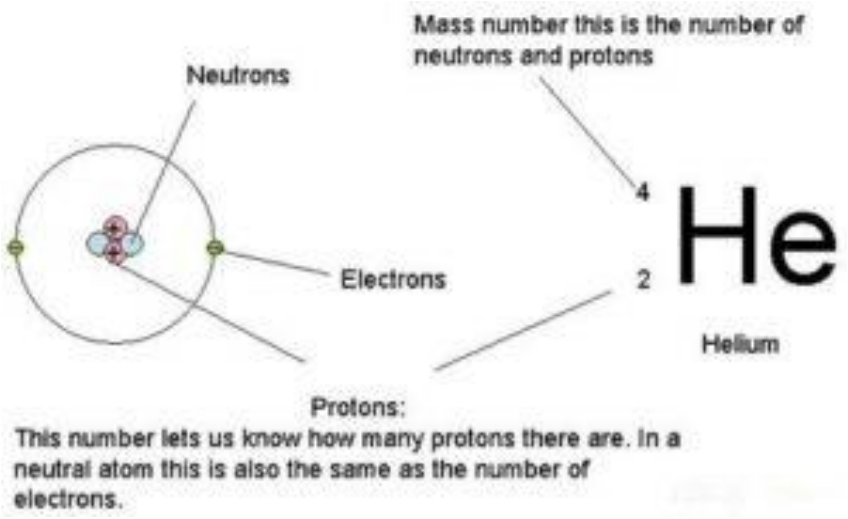
It does not matter how many electrons or neutrons present in an atom. Atomic number is always determined by the number of protons. For example

- If an atom has one proton then the atomic number is 1 and the element is hydrogen.
- Every carbon atom will have 6 protons and the atomic number of carbon is 6.
- Every cesium atom has 55 protons and hence the atomic number is 55.

Finding the Atomic Number

It can be calculated on the basis of information given. Firstly, an atomic number can be calculated by the number of protons in an atom of an element.

Secondly, if an element name or symbol is given then it can be identified by using a periodic table chart. It is a simple positive and whole number and is denoted by the letter z. For example, if the symbol Al (aluminum) is given. If you search the periodic table you can see the atomic number of Al is 13. It is easy to find out in a periodic table because the elements are present in increasing order.



Thirdly it can be identified by the isotope symbol. Isotope symbol can be written in many ways but the symbol is always given while writing an isotope. For example, if an isotope of carbon is given as ^{14}C or $^{14}{}_{6}C$.

- ^{14}C- If the isotope is written this way we know that the symbol represents carbon atom so the atomic number of carbon atom is 6.
- $^{14}_{6}C$- If the isotope is written this way the atomic number will be the whole number and smaller of the two numbers which is 6. It is usually the subscript of the element symbol.

SOLVED QUESTIONS FOR YOU

Question: What is the atomic number of chlorine?

a. 18 **b.** 19 **c.** 17 **d.** 16

Solution: The correct answer is "C". The atomic number of chlorine is 17 because the number of protons in a chlorine atom is 17.

Mass Number

Did you know that protons and neutrons together are termed as nucleons? Nucleons reside inside the nucleus and are denoted by the mass number is denoted as "A." All atoms of a particular chemical element will have the same atomic number, number of protons present in the nucleus of an atom, but different mass numbers. But what exactly are these mass numbers? Let's know more about it.

It is defined as the sum of protons and neutrons. The mass number is almost equal to the atomic mass of a particular atom. Therefore it can be written as:

Mass no. of an atom = No. of protons + No. of neutrons

Thus, it represents the total number of neutrons present in the nucleus of an atom.

Use of Mass Number

Mass number helps to give an idea of the isotopic mass. Isotopic mass is measured in atomic mass units or "u." An isotope of an element will have the same atomic number but a different mass number. Isotope mainly differs in the number of neutrons. Different isotopes of the same element will have a different mass number. However, isotopes of different elements can have the same mass number such as carbon-14 (6 protons + 8 neutrons) and nitrogen-14 (7 protons + 7 neutrons).

Properties of Mass Number

The various properties of mass number are enumerated here. Let us have a look,

- Sum of protons and neutrons provide this number of a certain element.
- It is represented by the letter A.
- Protons and Neutrons are together termed as nucleons.

  Example: Atoms of a carbon consist of 6 protons and 6 neutrons. Therefore, the mass number of Carbon is 12.

- The number of neutrons may vary in an element. However, the total number of protons is same in all atoms of an element. Therefore, the atoms of the same element with a same atomic number but a different mass number are termed as isotopes.
- Generally, atomic mass and mass numbers are two different terms and may vary slightly. In most cases, they are not the same. However, the weight of an electron is almost negligible so we can consider the atomic mass of an atom to be almost equal to its mass number.

X= Chemical Symbol of a particular element

N= Neutron number

Z= Atomic Number= Number of protons

A= Mass number= Z+N

SOLVED QUESTIONS FOR YOU

Question: Calculate the number of neutrons for the symbol $^{35}Cl_{17}$

Solution: We know that, Z= Atomic Number= Number of protons
Atomic number=Z= 17
Given, Mass no= A= 35
We know that Mass no. of an atom = No. of protons + No. of neutrons
35= 17 + no of neutrons Or, No of neutrons= 35-17= 18
Therefore, Number of neutrons=18

SOLVED QUESTIONS FOR YOU

Question: Match the columns:

1. Mass number	**a.** Nuclear Model of Atom
2. J.J. Thomson	**b.** Number of protons
3. Rutherford	**c.** Number of nucleons
4. Atomic number	**d.** Plum pudding model

Solution: 1 → c, 2 → d, 3 → a, 4 → b

As we know now, a positive charge on the nucleus is due to the protons. Also, the charge on the proton is equal but opposite to that of the electron. **Atomic Number (Z)** is the number of protons present in the nucleus. For example, the number of protons in sodium is 11 whereas it is 1 in hydrogen, Therefore, the atomic numbers of sodium and hydrogen are 11 and 1, respectively. Also, to maintain electrical neutrality, the number of electrons in an atom is equal to the number of protons (atomic number, Z). Therefore, the number of electrons in sodium and hydrogen is 11 and 1, respectively.

Atomic number = the number of protons in the nucleus of an atom

= the number of electrons in a neutral atom

The positive charge on the nucleus is due to protons, but the mass of the atom is due to protons and neutrons. They are collectively known as **nucleons**. **Mass number (A)** of the atom is the total number of nucleons.

Mass number (A) = the number of protons (Z) + the number of neutrons (n)

Therefore, the composition of an atom is represented using the element symbol (X) with the mass number (A) as super-script on the left and atomic number (Z) as sub-script on the left – ${}^{A}_{Z}X$.

How are Electrons Distributed in Different Orbits (Shells)?

By now, we know that, in an atom, electrons revolve around the nucleus while protons and neutrons are inside the nucleus. The question is, how do electrons revolve? Do they move randomly? Or do they follow a specific route? How are they arranged? The answer to these questions is Electronic Configuration. Let's learn about the arrangement of electrons around the nucleus.

Distribution of Electrons in Different Orbits

Neils Bohr gave the planetary model of an atom. He was the first person to suggest the periodicity in the properties of the elements. "Bohr atomic model" forms the basis of the electronic structure of an atom. He was the person to describe the arrangement of electrons (electronic configuration) in different orbits/shells.

He proposed that electrons are distributed in circular electronic shells (orbits). These electrons revolve in the orbits around the nucleus from a fixed distance. In this topic, we will learn more about the electronic configuration of different elements.

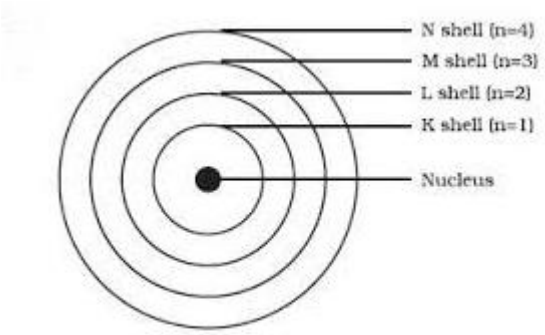
Bohr Bury Schemes

The distribution of electrons in an atom is called as Electronic Configuration. Formula $2n^2$ helps in the determination of the maximum number of electrons present in an orbit, here n= orbit number. The formula helps in determination of arrangement of electrons and is known as "Bohr Bury Schemes."

Electrons are negatively charged subatomic particles arranged like a cloud of negative charges outside the nucleus of an atom. The arrangement depends upon of their potential energies in different orbits. The different energy levels are known as 1, 2, 3, 4….. and the corresponding shells are known as K, L, M, N and so on.

For instance,

- 1^{st} energy level- K shell/orbit
- 2^{nd} energy level- L shell/orbit
- 3^{rd} energy level- M shell/orbit and so on.



A few energy levels in an atom

SOLVED QUESTIONS FOR YOU

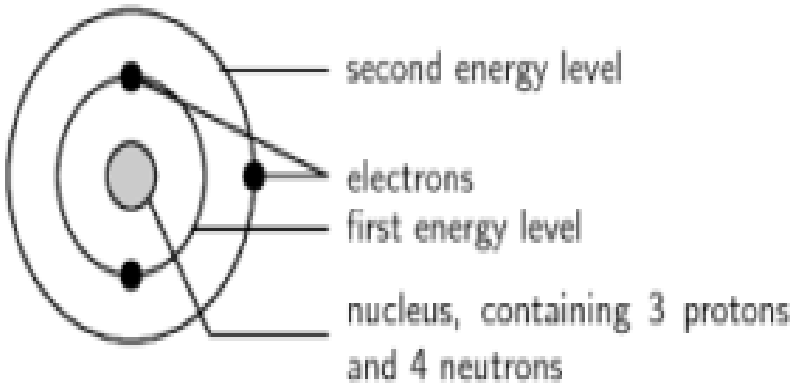
Question: Calculate the number of protons, neutrons, and electrons in $^{56}_{26}Fe$.

Solution: In $^{56}_{26}Fe$, atomic number (Z) = 26, mass number (A) = 56.
Number of protons = number of electrons = Z = 26.
Number of neutrons = A – Z = 56 – 26 = 30.

The Arrangement of Electrons in Different Orbits

The shells begin from the centre and gradually move outwards. So K shell will always have minimum energy. Similarly, L shell is a little away from nucleus so it will have higher energy than K shell. The outermost shell will have maximum energy. Now it is important to understand the distribution and arrangement of electrons in the atoms of any elements in the different energy levels.

An atom of any element is most stable when it has minimum energy. An atom will first fill the lowest energy level so as to attain the state of minimum energy. Gradually, the electrons will fill the higher energy levels. Therefore, electrons will first fill K shell, then L shell, M shell, N shell, and so on.



Electronic Configuration of Elements

According to the postulate of Neils Bohr, “electrons revolve around the centre of an atom (nucleus) in a predictable pathway named orbits”. The representation of the orbits is done by letters and numbers such as K, L, M, N, O…. and 1,2,3,4…. respectively. The arrangement and distribution of electrons in different orbits was given by Bohr and Bury.

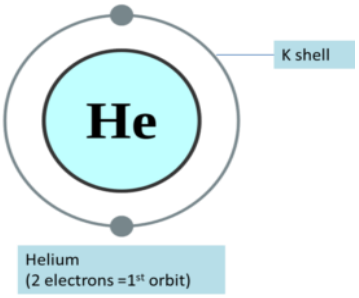
The arrangement of electrons in different shells and sub-shells is known as the electronic configuration of a particular element. The electronic configuration diagram represents an element in its ground state or stable state. There are a set of rules to remember while distribution off electrons in different orbits.

- **Rule 1**: The maximum number of electrons present in a particular shell is calculated by the formula $2n^2$, where “n” represents the shell number. For instance, K shell is the first shell and it can hold up to $2(1)^2 = 2$ electrons. Similarly, L shell is the second shell and it can hold up to $2(2)^2 = 8$ electrons. This formula helps to calculate the maximum number of electrons that an orbit can accommodate.
- **Rule 2**: The maximum capacity to hold electrons in the outermost shell is 8.
- **Rule 3**: The electrons will fill the inner shells before the outer shells. First electrons will fill the K-shell and then L shell and so on. Thus, electronic configuration of elements follows an ascending order.

Examples of the Electronic Configuration

1) Helium

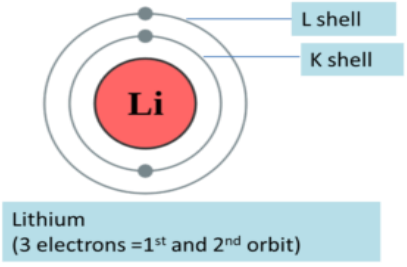
The atomic number of the element = 2. The total number of electrons present in Helium = 2. The maximum number of electrons in K shell (1st orbit) = 2. Therefore, shells needed = 1.



Electronic Configuration of Helium

2) Lithium

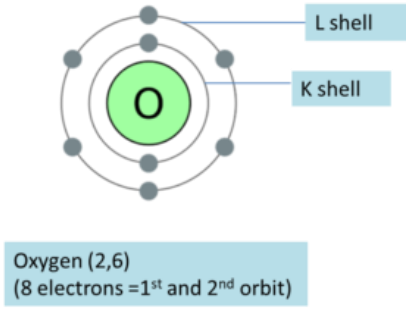
The atomic number of the element = 3. Lithium has 3 electrons. We can apply rule number 3 to fill the electrons in different in different orbits. The maximum number of electrons accommodated in K shell (1st orbit) will be 2. The second orbit will accommodate rest of the electrons. Electronic configuration of Lithium= 2, 1. Therefore, the total number of shells required = 2.



Electronic Configuration of Lithium

3) Oxygen

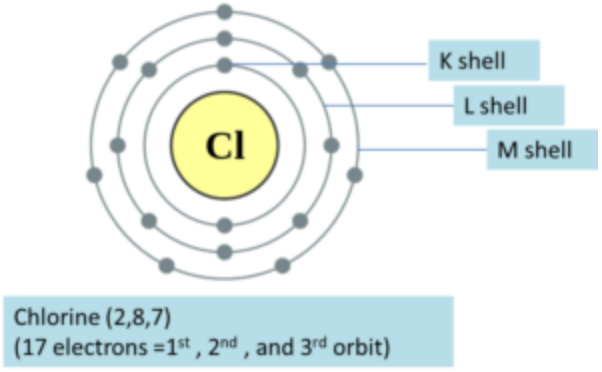
The atomic number of the element = 8. Oxygen has 8 electrons. The maximum number of electrons accommodated in the K shell (1st orbit) will be 2. The second orbit will accommodate rest of the electrons left (6 electrons). Electronic configuration of Oxygen = 2, 6. Therefore, the total number of shells required = 2 (1st and 2ndshell/orbit).



Electronic Configuration of Oxygen

4) Chlorine

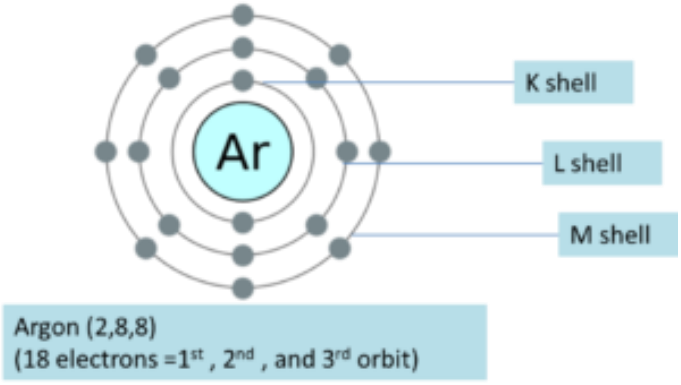
The atomic number of the element= 17. Chlorine has 17 electrons. The maximum number of electrons accommodated in the K shell (1st orbit) will be 2. The second orbit will fill up to 8 electrons. Finally, The third orbit will accommodate rest of the electrons left. Electronic configuration of Chlorine = 2, 8, 7. Therefore, the total number of shells required = 3 (1st, 2nd, and 3rd shell).



Electronic Configuration of Chlorine

5) Argon

The atomic number of the element= 18. Argon has 18 electrons. The maximum number of electrons accommodated in K shell is 2. The second orbit will fill up to 8 electrons. The third orbit can fill up to 18 electrons and it will accommodate left electrons of the element. Electronic configuration of Argon = 2, 8, 8. Therefore, the total number of shells = 3.



Electronic Configuration of Argon

Uses of the Electronic Configuration

- Electronic Configuration helps to understand the structure of periodic table with respect to each element.
- It also helps in understanding and explanation of the chemical bonds between the atoms.
- It explains the different properties and peculiar properties of certain elements. For example, electronic configuration explains the reason for the unique properties of lasers and semiconductors.

Importance of the Electronic Configuration

The electronic configuration is very important and basic part of understanding chemistry. It forms the basis of the periodic table. Additionally, the stability of any orbital will depend upon the electronic configuration of that element. It also helps us to understand the arrangement of elements in different periods and groups.

Electronic Distribution of First 18 Elements

Element	Symbol	Atomic No.	No. of Electrons	Electronic configuration				Valence
				Shells				
				K	L	M	N	
Hydrogen	H	1	1	1				1
Helium	He	2	2	2				0
Lithium	Li	3	3	2	1			1
Beryllium	Be	4	4	2	2			2
Boron	B	5	5	2	3			3
Carbon	C	6	6	2	4			4
Nitrogen	N	7	7	2	5			3
Oxygen	O	8	8	2	6			2
Fluorine	F	9	9	2	7			1
Neon	Ne	10	10	2	8			0
Sodium	Na	11	11	2	8	1		1
Magnesium	Mg	12	12	2	8	2		2
Aluminum	Al	13	13	2	8	3		3
Silicon	Si	14	14	2	8	4		4
Phosphorus	P	15	15	2	8	5		3
Sulphur	S	16	16	2	8	6		2
Chlorine	Cl	17	17	2	8	7		1
Argon	Ar	18	18	2	8	8		0

The reason behind the arrangement of electrons in such a manner is due to the presence of subshells. An atom always tends to remain in its stable state. Furthermore, it is necessary to arrange the electrons in the sub-shells in such a way that the element gains stability to attain the lowest energy level. There are separate principles to fill the electrons in its subshells. Hence, the fourth orbit will accommodate the one electron left. Thus, the Electronic configuration of potassium = 2, 8, 8, 1

SOLVED QUESTIONS FOR YOU

Question: The atomic numb er of aluminum is 13 and atomic mass is 27u. Calculate the number of protons and neutrons.

Solution: Given, Atomic number= 13 .
We know that, Z= Atomic Number= Number of protons. Therefore, number of proton=13
We know that Mass no. of an atom = No. of protons + No. of neutrons
Mass no= atomic mass= 27 u . 27 u = 13 + n Or, n = 27-13 = 14
Thus, Number of proton=13 and number of neutron= 14

SOLVED QUESTIONS FOR YOU

Question: Find the Electronic Configuration of Potassium (K).

1. 2,8,8,4 **2.** 2,8,8,1 **3.** 2, 8, 7 **4.** 2,8,8,3

Solution: The correct answer is option 2 (2, 8, 8, 1).
Atomic number of the potassium = 19. Potassium has 18 electrons. Applying rule number 3, the maximum number of electrons accommodated in K shell is 2. After filling the first orbit, the second orbit will consist of 8 electrons. Although it can fill up to 18 electrons the 3rd orbit will fill up to 8 electrons.

Orbits and energy levels

Unlike planets orbiting the Sun, electrons cannot be at any arbitrary distance from the nucleus; they can exist only in certain speci}c locations called allowed orbits. This property, first explained by Danish physicist Niels Bohr in 1913, is another result of quantum mechanics—specifically, the requirement that the angular momentum of an electron in orbit, like everything else in the quantum world, come in discrete bundles called quanta.

In the Bohr atom electrons can be found only in allowed orbits, and these allowed orbits are at different energies. The orbits are analogous to a set of stairs in which the gravitational potential energy is different for each step and in which a ball can be found on any step but never in between.

The laws of quantum mechanics describe the process by which electrons can move from one allowed orbit, or energy level, to another. As with many processes in the quantum world, this process is impossible to visualize. An electron disappears from the orbit in which it is located and reappears in its new location without ever appearing any place in between. This process is called a quantum leap or quantum jump, and it has no analog in the macroscopic world.

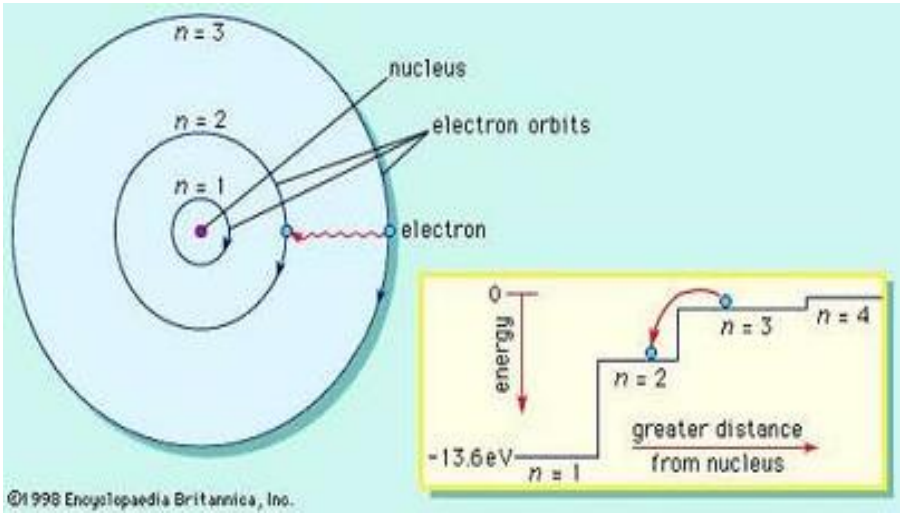
Because different orbits have different energies, whenever a quantum leap occurs, the energy possessed by the electron will be different after the jump. For example, if an electron jumps from a higher to a lower energy level, the lost energy will have to

go somewhere and in fact will be emitted by the atom in a bundle of electromagnetic radiation. This bundle is known as a photon, and this emission of photons with a change of energy levels is the process by which atoms emit light. *See also* laser.

In the same way, if energy is added to an atom, an electron can use that energy to make a quantum leap from a lower to a higher orbit. This energy can be supplied in many ways. One common way is for the atom to absorb a photon of just the right frequency. For example, when white light is shone on an atom, it selectively absorbs those frequencies corresponding to the energy differences between allowed orbits. Each element has a unique set of energy levels, and so the frequencies at which it absorbs and emits light act as a kind of fingerprint, identifying the particular element.

This property of atoms has given rise to spectroscopy, a science devoted to identifying atoms and molecules by the kind of radiation they emit or absorb. This picture of the atom, with electrons moving up and down between allowed orbits,accompanied by the absorption or emission of energy, contains the essential features of the Bohr atomic model, for which Bohr received the Nobel Prize for Physics in 1922. His basic model does not work well in explaining the details of the structure of atoms more complicated than hydrogen, however. This requires the introduction of quantum mechanics. In quantum mechanics each orbiting electron is represented by a mathematical expression known as a wave function—something

like a vibrating guitar string laid out along the path of the electron's orbit. These waveforms are called orbitals.



The Bohr atom The electron travels in circular orbits around the nucleus. The orbits have quantized sizes and energies. Energy is emitted from the atom when the electron jumps from one orbit to another closer to the nucleus. Shown here is the first Balmer transition, in which an electron jumps from orbit $n = 3$ to orbit $n = 2$, producing a photon of red light with an energy of 1.89 eV and awavelength of 656 nanometres.

Electron shells

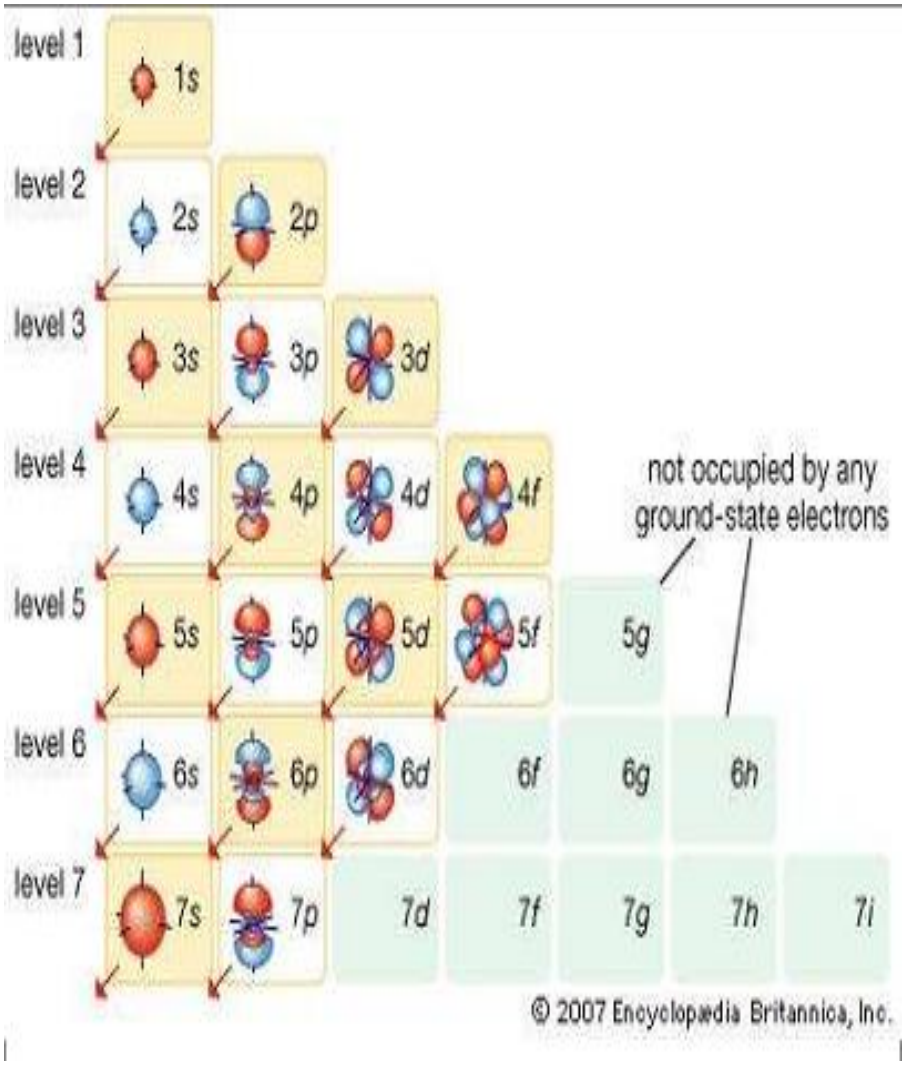
In the quantum mechanical version of the Bohr atomic model, each of the allowed electron orbits is assigned a quantum number *n* that runs from 1 (for the orbit closest to the nucleus) to in}nity (for orbits very far from the nucleus). All of the orbitals that have the same value of *n* make up a shell. Inside each shell there may be subshells corresponding to different rates of rotation and orientation of orbitals and the spin directions of the electrons. In general, the farther away from the nucleus a shell is, the more subshells it will have. *See* the table. This arrangement of possible orbitals explains a great deal about the chemical properties of different atoms.

The easiest way to see this is to imagine building up complex atoms by starting with hydrogen and adding one proton and one electron (along with the appropriate number of neutrons) at a time. In hydrogen the lowest-energy orbit—called the ground state—corresponds to the electron located in the shell closest to the nucleus. There are two possible states for an electron in this shell, corresponding to a clockwise spin and a counterclockwise spin (or, in the jargon of physicists, spin up and spin down).

The next most-complex atom is helium, which has two protons in its nucleus and two orbiting electrons. These electrons fill the two available states in the lowest shell, producing what is called a filled shell. The next atom is lithium, with three electrons. Because the closest shell is filled, the third electron goes into the

next higher shell. This shell has spaces for eight electrons, so that it takes an atom with 10 electrons (neon) to fil the first two levels. The next atom after neon, sodium, has 11 electrons, so that one electron goes into the next highest shell.

In the progression thus far, three atoms—hydrogen, lithium, and sodium—have one electron in the outermost shell. As stated above, it is these outermost electrons that determine the chemical properties of an atom. Therefore, these three elements should have similar properties, as indeed they do. For this reason, they appear in the same column of the periodic table of the elements (*see* periodic law), and the same principle determines the position of every element in that table. The outermost shell of electrons— called the valence shell— determines the chemical behaviour of an atom, and the number of electrons in this shell depends on how many are left over after all the interior shells are filled.



Atomic orbitals Electrons fill in shell and subshell levels in a semiregular process, as indicated by the arrows above. After filling the first shell level (with just an *s* subshell), electrons move into the second level *s* subshell and then into the *p* subshell, before starting on another shell level. Because of its lower energy state, the 4*s* orbital fills before the 3*d*, and similarly for later *s* orbitals (for example, 6*s* fills before 4*f*).

The Scrödinger Model

- Spectral analysis of elements suggested that electron shells don't have the same amounts of energy.
- Scrödinger called for sub-shells.
- Each shell is divided into subshells (also called sub-energy levels) -regions within a shell that have a similar amount of energy
- These have characteristic shapesand different amounts of energy s < p < d < f < g

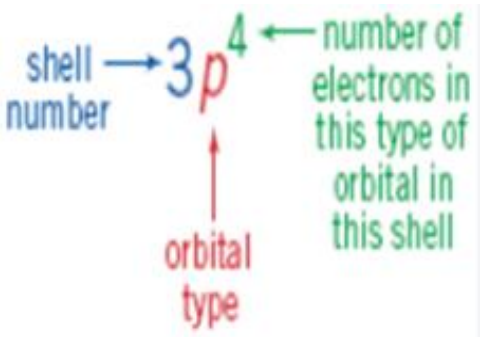
  n = number of subshells

  Sharp (s) =2

  Principal (p) = 6

  Diffuse (d) = 10

  Fine (f) = 14



The Subshells

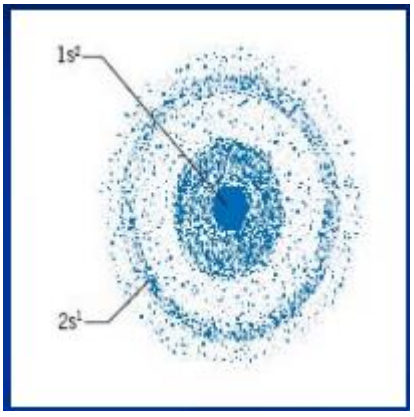
Electron shell	Number of shell (n)	Total number of subshells present in the shell	Names of subshells	Maximum number of electrons in the subshell	Maximum number of electrons present in the shell
K	1	1	1s	2	2
L	2	2	2s 2p	2 6	8
M	3	3	3s 3p 3d	2 6 10	18
N	4	4	4s 4p 4d 4f	2 6 10 14	32

Atomic Orbitals

Electrons are not confined to circular orbits.

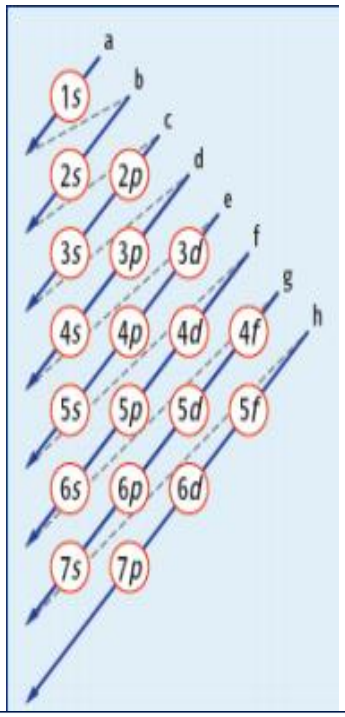
Instead they move in regions of space called atomic orbitals. Each subshell is made up of orbitals.

An orbital is the region of space in which the electron travels within the subshell.



Orbitals are assigned a letter according to the subshell it is in (s, p, d, f, g)

There can be: 1 s orbital, 3 p orbitals, 5 d orbitals, 7 f orbitals, 9 g orbitals.


a
b
c
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e
f
g
h
1s
2s
2p
3s
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6s
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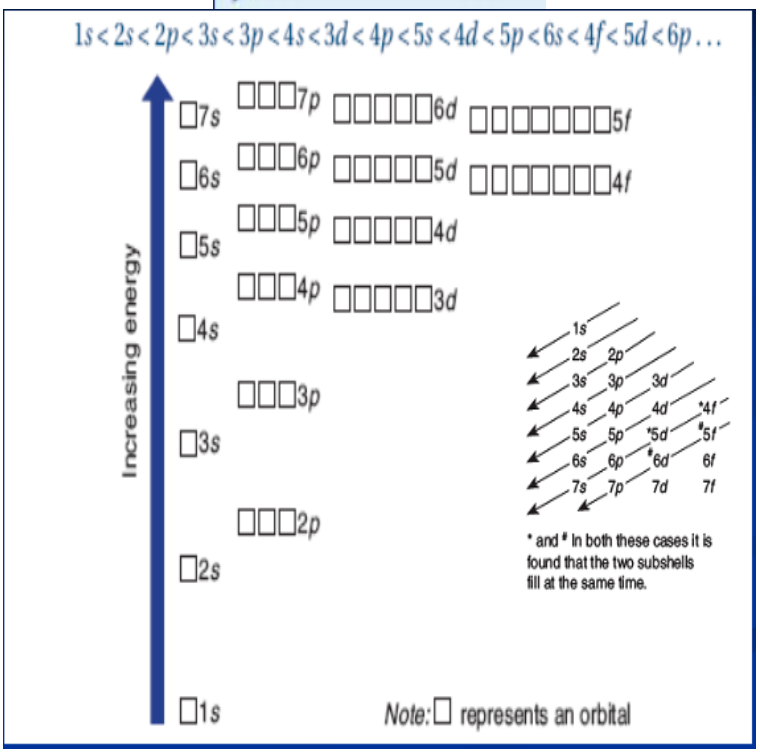
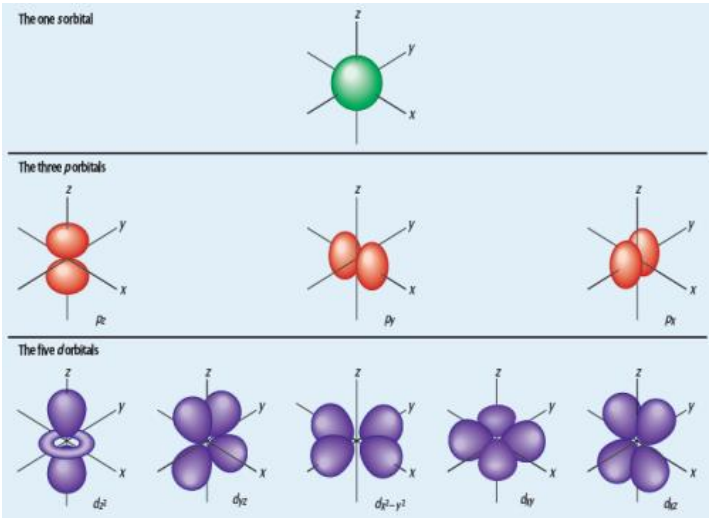

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p . . .
Increasing energy
7s
7p
6d
5f
6s
6p
5d
4f
5s
5p
4d
4s
4p
3d
3s
3p
2s
2p
1s
Note: represents an orbital
1s
2s 2p
3s 3p 3d
4s 4p 4d *4f
5s 5p *5d #5f
6s 6p *6d 6f
7s 7p 7d 7f
* and # In both these cases it is found that the two subshells fill at the same time.



Modern atomic theory

It was based on three hypotheses

- The dual nature of the electron
- Principle of uncertainty (Heisenberg)
- Wave equation (Schrodinger)

The following is a brief explanation of these three concepts to show us the modern view of the atomic structure.

The dual nature of the electron Experiments have shown that the electron has a dual nature in the sense that it is a material particle with wave properties. The movement of any particle, such as the electron, nucleus or part, is accompanied by a wave movement called the physical waves which different from electromagnetic waves in: is not separated from the moving body and the speed is not equal to the speed of light

The uncertainty principle of "Heisenberg"

Heisenberg has used quantum mechanics to an important principle: The simultaneous determination of the location and velocity of the electron simultaneously is virtually impossible, but speaking in the language of probability is the closest to the right, where we can probably say very much or very little the presence of the electron in this place.

Wave equation for Schrodinger

Schrödinger, based on the ideas of Planck, Einstein, De Brawley, and Heisenberg, developed the wave equation and solved this equation: The creation of permitted energy levels and the determination of the vacuum zones around the nucleus where the probability of the electron being as large as possible is present, and therefore the expression of the electronic cloud is the acceptable form of the orbital.

Electronic cloud: The region where the electron is likely to exist in all directions and dimensions around the nucleus.

The Pauli Exclusion Principle

To explain certain aspects of atomic spectra, Wolfgang Pauli determined that no 2 electrons can have all 4 quantum numbers alike. This is called the Pauli Exclusion Principle.

The **Pauli Exclusion Principle suggests that only two electrons with opposite spins can occupy an atomic orbital**. Stated another way, **no two electrons have the same 4**

quantum numbers *n, l, m, s*. Pauli's exclusion principle can be stated in some other ways, but the idea is that energy states have limited room to accommodate electrons.
A state accepts two electrons of different spins.

In full orbitals (orbitals containing 2 electrons of opposite spin) one electron must be spin up, and the other spin down, and the electrons are said to be paired.

Electronic Configurations For Atoms With More Than One Electron

The Schroedinger wave equation was developed initially for hydrogen, an atom with only one electron. In such a case, all orbitals in each energy level have the same energy and are called **Degenerate**

In atoms with more than one electron, the electrons repel each other, also the effective nuclear charge varies with the atomic number and the inner shell electrons screen the outer ones.
As a result, the orbital energies are shifted somewhat as shown in the figure below.

Hund's Rule

Hund's rule suggests that electrons prefer parallel spins in separate orbitals of subshells. This rule guides us in assigning electrons to different states in each subshell of the atmic orbitals. **In other words, electrons fill each and all orbitals in the subshell before they pair up with opposite spins.**

Pauli exclusion principle and Hund's rule guide in figuring the electron configurations for all elements.

Task *

1. Spin-orbit coupling splits all states into two except the s state. Why ?

2. Explain why the effective radius of helium atom is less than that of a hydrogen atom.

Formative Evaluation *

1. Determine the shortest and longest wavelengths of Lyman series of hydrogen.
2. The study of atomic spectra was a kind of an industry towards the end of 19th century and at the beginning of the 20th century. Discuss
3. The longest wavelength in the Lyman series for hydrogen is 1215 Å. Calculate the Rydberg constant.
4. Electrons of energy 12.2 eV are fired at hydrogen atoms in a gas discharge tube. Determine the wavelengths of the lines that can be emitted by the hydrogen.
5. Determine the magnetic moment of an electron moving in a circular orbit of radius r aout a proton.
6. Use the results from quantum mechanics to calculate the magnetic moments that are possible for n = 3.

7. Determine the normal Zeeman splitting of the cadmium red line of 6438 Å when the atoms are placed in a magnetic field of 0.009 T.
8. Express **L.S**. in terms of J, L, and S. Given L = 1 and S = ½, calculate the possible values of **L.S**.
9. A beam of electrons enters a uniform magnetic field B = 1.2 T. Find the energy difference between electrons whose spin are parallel and antiparallel to the magnetic field.

PART YWO

ATOMIC SPECTRUM:

In this lesson, we'll discuss the meaning of the term 'atomic spectrum' and distinguish between the two main types of atomic spectra: absorption and emission spectra. We'll also describe how each forms.

Rainbows are one of the most beautiful things on Earth, containing a spread of the colors that our eyes can see. That's why we use the expression 'every color of the rainbow.' But did you know this expression isn't technically true -- that the rainbow is actually missing a few colors? If you could look extremely closely at a rainbow and analyze its light, you would find dark 'gaps' in the otherwise continuous spectrum of light. This is because pure, white light can be shined through a prism and split into every color our eyes can see. But the Sun doesn't send us pure, white light. It sends us an atomic spectrum.

An atomic spectrum is a spectrum that has been shined through or originates from a material (usually a gas) and contains patterns that are characteristic of the elements present in the material. When we analyze the Sun's spectrum, for example, we can Ògure out what elements are present in the Sun.

Atomic spectroscopy

Atomic spectroscopy is the study of the electromagnetic radiation absorbed and emitted by atoms. Since unique elements have characteristic (signature) spectra, atomic spectroscopy, specifically the electromagnetic spectrum or mass spectrum, is applied for determination of elemental compositions. It can be divided by atomization source or by the type of spectroscopy used. In the latter case, the main division is between optical and mass spectrometry. Mass spectrometry generally gives significantly better analytical performance, but is also significantly more complex. This complexity translates into higher purchase costs, higher operational costs, more operator training, and a greater number of components that can potentially fail. Because optical spectroscopy is often less expensive and has performance adequate for many tasks, it is far more common Atomic absorption spectrometers are one of the most commonly sold and used analytical devices.

Optical spectroscopy

Electrons exist in energy levels (i.e. atomic orbitals) within an atom. Atomic orbitals are quantized, meaning they exist as defined values instead of being continuous (see: atomic orbitals). Electrons may move between orbitals, but in doing so they must absorb or emit energy equal to the energy difference between their atom's specific quantized orbital energy levels. In optical spectroscopy, energy absorbed to move an electron to a higher energy level (higher orbital) and/or the energy emitted as the electron moves to a lower energy level is absorbed or emitted in the form of photons (light particles). Because each element has a unique number of electrons, an atom will absorb/release energy in a pattern unique to its elemental identity (e.g. Ca, Na, etc.) and thus will absorb/emit photons in a correspondingly unique pattern. The type of atoms present in a sample, or the amount of atoms present in a sample can be deduced from measuring these changes in light wavelength and light intensity.

Optical spectroscopy is further divided into atomic absorption spectroscopy and atomic emission spectroscopy. In atomic absorption spectroscopy, light of a predetermined wavelength is passed through a collection of atoms. If the wavelength of the source light has energy corresponding to the energy difference between two energy levels in the atoms, a portion of the light will be absorbed. The difference between the intensity of the light emitted from the source (e.g., lamp) and the light collected by the detector yields an absorbance value. This absorbance value can then be used to determine the

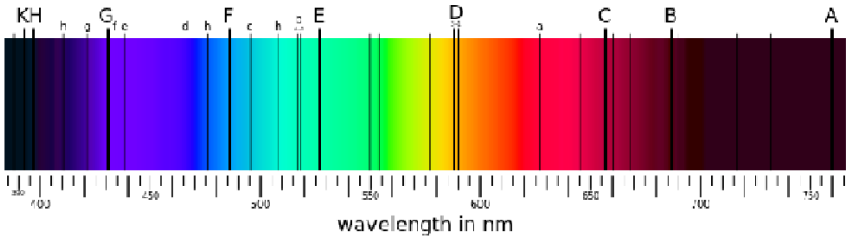
concentration of a given element (or atoms) within the sample. The relationship between the concentration of atoms, the distance the light travels through the collection of atoms, and the portion of the light absorbed is given by the Beer–Lambert law. In atomic emission spectroscopy, the intensity of the emitted light is directly proportional to the concentration of atoms.

Mass spectrometry

Atomic mass spectrometry is similar to other types of mass spectrometry in that it consists of an ion source, a mass analyzer, and a detector. Atoms' identities are determined by their mass-to-charge ratio (via the mass analyzer) and their concentrations are determined by the number of ions detected. Although considerable research has gone into customizing mass spectrometers for atomic ion sources, it is the ion source that differs most from other forms of mass spectrometry. These ion sources must also atomize samples, or an atomization step must take place before ionization. Atomic ion sources are generally modifications of atomic optical spectroscopy atom sources.

Absorption Spectrum

To analyze the Sun's light, we shine it through a spectrometer, which is a device that separates light by energy and color. When we do that, we create an image of the Sun's spectrum that looks like this:



This is an absorption spectrum similar to the Sun. The black lines show where light is absorbed by the elements in the outer layers of the Sun. Red light is the lowest energy and blue light is the highest energy, just like how a red ember on a fire is not as hot as a yellow fame, and blue fames are the hottest of all!

The black bars (gaps) in the Sun's spectrum are known as absorption lines, and they're caused by the gases in the outer layers of the Sun absorbing some of the light.

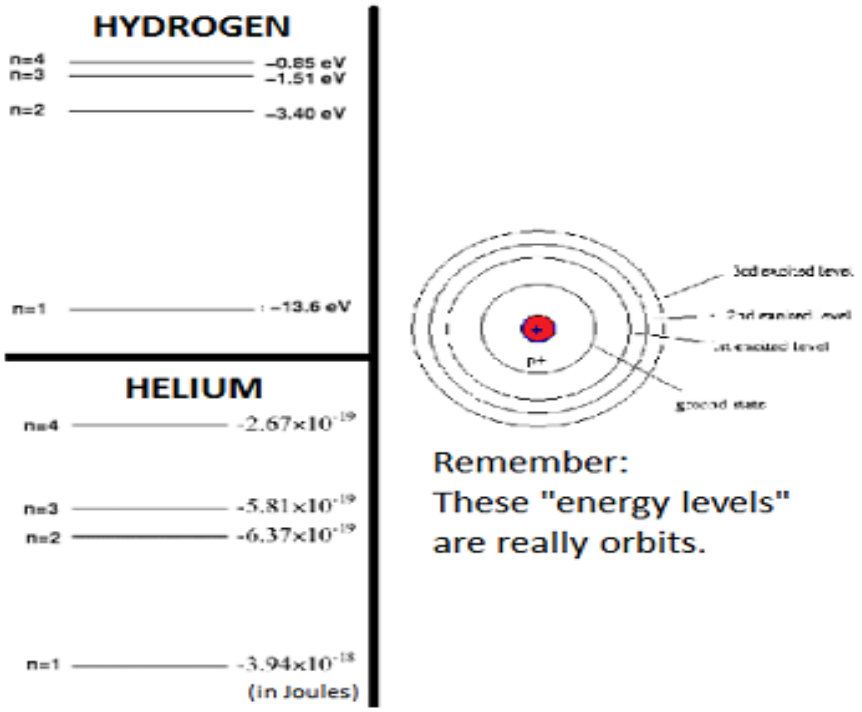
The Sun contains many elements: hydrogen, helium, carbon, and smaller amounts of heavy elements. When the light from the Sun shines through these elements, the atoms absorb the energy, but they only absorb light that is just the right color to match the energy they need. This gives us those gaps in the Sun's spectrum. And by looking at the gaps, we can see what the Sun is made of.

Emission Spectrum

An emission spectrum is the opposite of an absorption spectrum. Instead of getting light with a few colors missing, in an emission spectrum, those are the only colors we get. To create an absorption spectrum, we had to shine light through a gas. But to create an emission spectrum, we heat up a gas instead. The atoms in the gas will absorb this energy, but only for a little while. Heating the gas causes the atoms to be jiggled up and energetic... they have too much energy. Eventually, this energy is re-released (or *emitted*) as light. The color of the light that is emitted is different for every element, so we can look at the emission spectrum of any given gas to figure out what elements are in the gas we're heating.

Atoms and Energy Levels

So far, we've learned about how atoms can absorb the light of particular colors or, when heated up, emit light of those same colors. But how does it do that? And why is it only certain colors? Atoms contain electrons orbiting around the nucleus, and those electrons sit inside energy levels (also known as shells). Energy levels are particular orbits, or particular amounts of energy, that electrons are allowed to have. Some examples are shown here:



As you can see, we can draw them as circular orbits (like on the right), or represent them as straight lines, like floors of a hotel (like on the left).

The electrons in an atom can absorb energy and jump from a lower energy level to a higher one but only if they get exactly the right amount of energy to make the jump. If they jump too far, or not far enough, they'll miss! And if they're going to miss, then they stay where they are.