CHM 101- INTRODUCTION TO GENERAL CHEMISTRY

THE ATOMIC NATURE OF MATTER

At the beginning of human history, people started to think about what matter is and how matter behave. At about 400BC, the Greek philosopher Democritus argued that if one divided a piece of matter such as iron into smaller and smaller pieces, one would eventually reach an invisibly small particle of iron that could not be sub-divided further but still retain the properties of iron. Democritus gave these ultimate particles the name “atomos” (which literally means “indivisible” in greek). He believed that all matter was made up of various arrangements of atoms. The idea was a philosophical one and was not proved by any experimental evidence.

It was until 1808 that idea of atomic nature of matter was seriously revived by John Dalton. He suggested that matter was made up of different tiny particles that would remain unchanged during chemical reactions. The theory was supported by experimental evidence.

DALTON’S ATOMIC THEORY

The main points of Dalton’s atomic theory are as follows:

1. All matter is made up of tiny, indivisible particles called atoms.
2. Atoms can neither be created nor destroyed.
3. Atoms of the same element are identical and have the same mass and chemical properties.
4. Atoms of different elements are different, having different masses and different chemical properties
5. Atoms of different elements combine together to form compound, the number of atoms combined bearing a simple whole number ratio to each other.

The atomic theory was verified by three fundamental laws of chemistry, they are:

1. The laws of conservation of mass
2. The laws of definite proportion
3. The law of multiple proportion

ATOM

An **atom** is the basic unit that makes up all matter. There are many different types of atoms, each with its own name, mass and size. These different types of atoms are called chemical elements. Examples of elements are hydrogen and gold

Atoms are very small, but the exact size changes depending on the element.

Atoms range from 0.1 to 0.5 nanometers in width. One nanometer is around 100,000 times smaller than the width of a human hair. This makes atoms impossible to see without special tools.

Equations must be used to see the way they work and how they interact with other atoms.

Atoms come together to make molecules or particles: for example, two hydrogen (H_2) atoms and one oxygen (O_2) atom combine to make a watermolecule, a form of a chemical reaction.

Atoms themselves are made up of three kinds of smaller particles, called protons, neutrons and electrons. The protons and neutrons are in the middle of the atom. They are called the nucleus. The nucleus is surrounded by a cloud of electrons with a negative charge which are bound to the nucleus by an electromagnetic force.

Protons and neutrons are made up of even smaller particles called quarks. Electrons are elementary or fundamental particles; they cannot be split into smaller parts.

The number of protons, neutrons and electrons an atom has determines what element it is. Hydrogen, for example, has one proton, no neutrons and one electron; the element sulfur has 16 protons, 16 neutrons and 16 electrons.

Atoms move faster when in gas form (as they are free to move) than liquid and solid matter. In solid materials the atoms are tightly next to each other so they vibrate, but are not able to move (there is no room) as atoms in liquids do.

Structure and Part

Parts

The complex atom is made up of three main particles; the proton, the neutron and the electron. The isotope of Hydrogen Hydrogen-1 has no neutrons, and a positive hydrogen ion has no electrons. These are the only known exceptions, all other atoms have at least one proton, neutron and electron each.

Electrons are by far the smallest of the three, their mass and size is too small to be measured using current technology. They have a negative charge. Protons and neutrons are similar sizes. Protons are positively charged and neutrons have no charge. Most atoms have a neutral charge; because the number of protons (positive) and electrons (negative) are the same, the charges balance out to zero. However in ions (different number of electrons) this is not always the case and they can have a positive or a negative charge. Protons and Neutrons are made out ofquarks, of two types; up quarks and down quarks. A proton is made of two up quarks and one down quark and a neutron is made of two down quarks and one up quark.

Nucleus

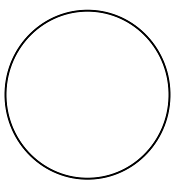
The nucleus is in the middle of an atom. It is made up of protons and neutrons. Usually in nature, two things with the same charge repel or shoot away from each other. So for a long time it was a mystery to scientists how the positively charged protons in the nucleus stayed together. They solved this by finding a particle called a Gluon. Its name comes from the word *glue* as Gluons act like atomic glue, sticking the protons together using the *strong nuclear force*. It is this force which also holds the quarks together that make up the protons and neutrons

Electrons

Electrons orbit or go around the nucleus. They are called the atom's *electron cloud*. They are attracted towards the nucleus because of the electromagnetic force. Electrons have a negative charge and the nucleus always has a positive charge, so they attract each other. Around the nucleus some electrons are further out than others. These are called *electron shells*. In most atoms the first shell has two electrons, and all after that have eight. Exceptions are rare, but they do happen and are difficult to predict.[12] The further away the electron is from the nucleus, the weaker the pull of the nucleus on it. This is why bigger atoms, with more electrons, react more easily with other atoms. The electromagnetism of the nucleus is not enough to hold onto their electrons and they lose them to the strong attraction of smaller atoms

Sub-atomic particle	Symbols	Location in atom	Charge (c)	Relative charge	Mass (g)	Approx.relative mass (amu)
Proton	P	Nucleus	1.6×10^{-19}	+1	1.7×10^{-24}	1
Neutron	N	Nucleus	0	0	1.7×10^{-24}	1
Electron	*E*	Surrouding the Nucleus	1.6×10^{-19}	-1	9.1×10^{-28}	0

- Amu- atomic mass unit



ATOMIC NUMBER, MASS NUMBER AND ISOTOPE

Atomic number (Z) of an element is the number of protons contained in the nucleus of the atom. It is a basic property and determines the identity of an element. Because atoms are electrically neural, the atomic number is the same as the number of electrons surrounding the nucleus.

Atomic number = **number of protons** = **no of electrons**

Mass number (A) is the sum of the of protons and neutrons in the nucleus of an atom

Mass number = **no. of protons** + **no. of neutrons**

The numbers of protons and neutrons are integers, hence the mass number is a whole number.

Atom	***symbol***	***Number of protons***	***Number of electrons***	***Number of neutrons***	***Atomic number***	***Mass number***
Hydrogen	H	1	1	0	1	(1+ 0) = 0
Oxygen	O	8	8	8	8	(8+ 8) = 16
Argon	Ar	18	18	22	18	(18 +22) =40
Radon	Rn	86	86	136	86	(86+136)=222

ISOTOPES

Are atoms of the same element with the same number of protons (same atomic number) but different number of neutrons (different mass number). Isotopes of the same element have exactly the same chemical properties although their physical properties differ

$^{A}_{Z}X$

X= symbol

A= mass number

Z= atomic number

The two isotopes of Chlorine are

$^{35}_{17}Cl$ $^{37}_{17}Cl$

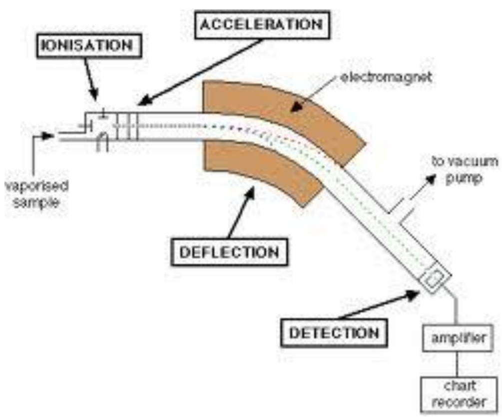
Element	***Isotope***	***Atomic Number***	***No. of Protons***	***No. of Neutrons***	***Natural abundance (%)***
Hydrogen	$^{1}_{1}H$ $^{2}_{1}H$	1 1	1 1	0 1	99.98 0.02
Helium	$^{3}_{2}He$ $^{4}_{2}He$	2 2	2 2	1 2	0.0001 99.999
Lithium	$^{6}_{3}Li$ $^{7}_{3}Li$	3 3	3 3	3 4	7.42 92.58
Beryllium	$^{9}_{4}Be$	4	4	5	100.00
Boron	$^{10}_{5}B$ $^{11}_{5}B$	5 5	5 5	5 6	19.6 80.4
Carbon	$^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$	6 6 6	6 6 6	6 7 8	98.89 1.11 Trace
Nitrogen	$^{14}_{7}N$ $^{15}_{7}N$	7 7	7 7	7 8	99.63 0.37

Oxygen	$^{16}_{8}O$ $^{17}_{8}O$ $^{18}_{8}O$	8 8 8	8 8 8	8 9 10	99.76 0.04 0.20
Fluorine*	$^{19}_{9}F$	9	9	10	100.0
Sulphur	$^{32}_{16}S$ $^{33}_{16}S$ $^{34}_{16}S$ $^{36}_{16}S$	16 16 16 16	16 16 16 16	16 17 18 20	95.0 0.76 4.23 0.01
Chlorine	$^{35}_{15}Cl$ $^{37}_{17}Cl$	17 17	17 17	18 20	75.77 24.23
Copper	$^{63}_{29}Cu$ $^{65}_{29}Cu$	29 29	29 29	34 36	69.09 30.91
Uranium	$^{234}_{92}U$ $^{235}_{92}U$ $^{238}_{92}U$	92 92 92	92 92 92	142 143 146	0.01 0.72 99.27

*Note: Fluorine has no isotope

MASS SPECTROMETER

Relative masses of isotopes and molecules are determined by a very sophisticated apparatus called the mass spectrometer. It is the most accurate instrument currently available for comparing the masses of atoms nd molecules. It can differentiate between isotopes of the same element with only a very small mass difference of several arbitrary units (each arbitrary unit = 1.674 x 10^{-24}g); this can hardly be done by even the most sensitive weighing device. It measures the mass charge (*m/e*) ratio of ionized particles.

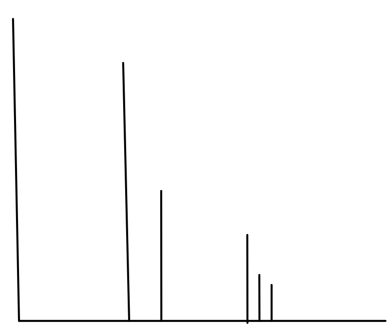


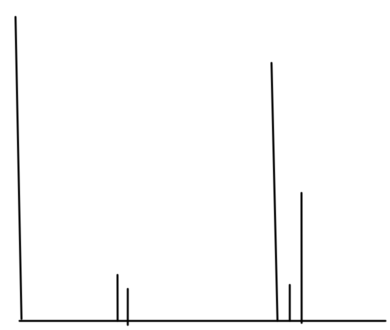
In the mass spectrometer, there are six important parts. Their names and functions are as follows

1. **Vaporization chamber-** the material to be analysed can be an element or a compound. A sample of the material under investigation is injected into the chamber and heated until vaporizes and changes to the gaseous form.
2. **Ionization chamber:** there is a heated filament in this chamber giving high-energy electrons. The vaporized sample (may be atoms or molecules) is then bombarded by the electrons. Atoms or molecules of the sample are ionized here. Positive ions formed as electrons are knocked off.

   $$X(g) + \underset{\text{Fast}}{e} \longrightarrow X^{+}(g) + \underset{\text{Slow}}{2e}$$

3. **Accelerating electric field-** the positive ions formed in the ionization chamber are accelerated here. They move towards the magnetic field with a very high speed.
4. **Deflecting magnetic field-** the positive ions are deflected here along a circular path. The lighter the positive ions, the greater is the deflection, and the smaller the radius of the circular path taken by the ions. The higher the charge of the positive ions, the greater the deflection. Ions with a high ratio of m/e are deflected less than those with low m/e ratio. By varying the accelerating electric field or deflecting magnetic field, ions of any m/e ratio can be targeted into the ion detector.
5. **Ion detector-** the detector converts the signal of different m/e ratio received into electric current. The magnitude of the current is proportional to the number of ions.
6. **Recorder-** ions detected in the detected in the detector are traced out by the recorder showing the m/e ratio of theions and their corresponding intensity. The position of the peak gives the m/e ratio of an ion. The height of the peak gives the **relative abundance** of the ion.





On the mass spectrum of an element, different peaks correspond to various isotopes of the element. Fig shows the mass spectrum of chlorine and chloromethane. In the mass spectrum of a compound, the peak with the highest m/e ratio will most likely correspond to the 'molecular ion', i.e. the molecule which has lost only a single electron

Table below illustrates the corresponding "molecular ions" or ions of the peaks in a) Cl and b) chloromethane mass spectra. In most spectra, the values of m/e ratio can be converted to the relative masses of the ions if the charges on the ions are known

Table 1: the corresponding " molecular ions" or ions of the peaks

(a) Chlorine Cl_2 b) Chloromethane CH_3Cl

m/e ratio	***Ions***
35	$^{35}Cl^+$
37	$^{37}Cl^+$
70	$^{35}Cl - {}^{35}Cl^+$
72	$^{35}Cl - {}^{37}Cl^+$
74	$^{37}Cl - {}^{37}Cl^+$

m/e ratio	***Ions***
35	$^{35}Cl^+$
37	$^{37}Cl^+$
50	$^{12}CH - {}^{35}Cl^+$
51	$^{12}CH_3 - {}^{37}Cl^+$
52	$^{37}Cl - {}^{37}Cl^+$

RELATIVE ISOTOPIC, ATOMIC AND MOLECULAR MASSES

Relative Isotopic Mass

Relative isotopic mass of a particular isotope of an element is the relative mass of one atom of that isotope on the $^{12}C = 12.000$ scale

Relative masses of all other atoms are compared with this scale as a reference standard to give their relative isotopic mass. The relative masses of a proton and a neutron are both taken as 1 and that of an electron is taken as 0.

The position of the peak in the mass spectrum gives the relative mass of the isotope

Relative Atomic Mass

Relative atomic mass of an element is the weighted average of the relative isotopic masses of the natural isotopes on the ^{12}C= 12.000 0 scale.

THE MOLE CONCEPT

The mole and the avogadro's constant

Chemists are always interested in counting the number of particles (including atoms, molecules and ions) in substances. As these particles are incredibly small, it is obviously impossible to count the numbers. Even to obtain the smallest amount of carbon visible to the naked eye, billions of carbon atoms would be required. Consequently, chemists now use a unit called the mole for counting particles.

The mole is defined as the amount of substances containing the same number of particles as the number of carbon atoms in exactly 12g of carbon-12. The particles may be atoms, molecules, ions or electrons.

One mole = 6.02×10^{23} particles

1 mole of C atoms = 6.02×10^{23} C atoms

2 moles of Na^+ ion = $2 \times 6.02 \times 10^{23}$ Na^+ ion.

1 mole of NaCl formula unit = $6.02 \times 6.02 \times 10^{23}$ NaCl formula unit

3 moles of CO_2 molecules = $3 \times 6.02 \times 10^{23}$ CO_2 molecules

The number 6.02×10^{23} is called the Avogadro constant as this was experimentally determined by Amadeo Avagadro.

In conclusion,

Number of moles = Number of particles/ Avogadro constant

The Mole and the Molar mass

It is easy to count 12 eggs but it seems impossible to count a mole (602 000 000 000 0000 …) of carbon atoms. Carbon particles are incredibly small and the number is so large. Fortunately, we can count indirectly by weighing. Hence, the mass of a mole depends on the identity of the substance being weighted.

The **molar mass** is the mass of one mole of a substance. The molar mass of a substance has the same numerical value as the relative atomic mass or formula mass* of the substance is but it is expressed in gram per mole

Note: relative atomic mass or formula mass carry no units

Table 2: The relations between relative atomic mass/ formula mass and molar mass

Substance	***Relative Atomic mass/ formula mass***	***Molar mass (g mol^{-1})***
C	12.0	12.0
S	32.1	32.1
Cl_2	71.0	71.0
NH_3	17.0	17.0
Al_2O_3	102.0	102.0

As the molar mass of a substance can be easily found by looking up the relative atomic masses of elements already provided, the number of moles a substance can be easily found by simple calculations using the following equation.

Number of moles = Mass of the substance / molar mass of the substance

As one mole of any substances contains 6.02×10^{23} structural units, the number of particles in a known mass of substance can be found by:

Number of particles = Mass/ molar mass x Avogadro constant

Examples

1. *What is the mass of 0.2 mole of calcium carbonate?*

(R.a.m : C= 12.0, O = 16.0, Ca = 40.1)*

Solution

The Chemical formula of calcium carbonate is $CaCO_3$

The molar mass of $CaCO_3$ = 40.1 + 12 + 3x 16.0

= 100.1 gmol^{-1}

Mass the $CaCO_3$ = number moles x molar mass

= 0.2 x 100.1

= 20. 02 g

2. *Calculate the number of gold atoms in a 20g gold coin*

   (R.a.m: Au =197.0)

Solution

Number of gold atoms in a 20g gold coin

= 20/ 197 x 6.02 x 10^{23}

= 6.11 x 10^{22}

3. *Given that the molar mass of water is 18 g mol^{-}*
   a. *What is the mass of 4 moles of water molecules*
   b. *How many molecules are there?*
   c. *How many atoms are there?*

   (R.a.m. : H= 1, O= 16)

   Solution

   a. Mass of water = number of molecules x molar mass
      = 4 x (1.0 x 2 + 16)
      = 72g
   b. There are 4 moles of water molecules
      Number of molecules = number of moles x avogadro constant
      = 4 x 6.02 x 10^{23}
      = 2.408 x 10^{24}
   c. 1 water molecule has 3 atoms
      1 mole of water molecules has 3 moles of atoms of atoms
      Thus, 4 moles of water molecules has 12 moles of atoms
      Number of atoms = 12 x 6.02 x 10^{23}
      = 7.224 x 10^{24}
4. *A magnesium chloride solution contains 10g of magnesium chloride*
   a. *Calculate the no of moles of magnesium chloride in the solution*
   b. *Calculate the number of magnesium ions in the solution*
   c. *Calculate the number of chloride ions in the solution*
   d. *Calculate the total number of ions in the solution*
      (R.a.m. : Mg= 24.3, Cl =35.5)

   Solution

   a. The chemical formula of magnesium chloride is $MgCl_2$
      Molar mass of magnesium chloride = 24.3 + 35.5 x 2
      = 95.3 g mol^{-1}
      Number of moles of magnesium chloride = 10/ 95.3
      = 0.105 mol
   b. 1 mole of $MgCl_2$ contains 1 mole of Mg^{2+} and 2 moles of Cl^-
      Then, 0.105 mole of $MgCl_2$, contains 0.105 mole of Mg^{2+}
      Number of Mg^{2+} ions = number of moles of Mg^{2+} x 6.02 x 10^{23}
   c. 0.105 mole of $MgCl_2$ contains 0.21 mole of Cl^-
      Number of Cl^- ions = number of moles of Cl^- x 6.02 x 10^{23}

$= 0.21 \times 6.02 \times 10^{23}$

$= 1.264 \times 10^{23}$

d. Total number of ions = $6.321 \times 10^{23} + 1.264 \times 10^{23}$

$= 1.896 \times 10^{23}$

5. *What is the mass of a carbon dioxide molecule?*
   (R.a.m. : C= 12, O= 16)

   Solution

   The chemical formula of carbon dioxide is CO_2

   The molar mass of carbon dioxide = 12.0 + 16.0 x 2

   $= 44.0\ gmol^{-1}$

   Number of moles = mass / molar mass = number of molecules/ Avogadro constant

   Mass of a mole/ 44 = 1/ 6.02×10^{23}

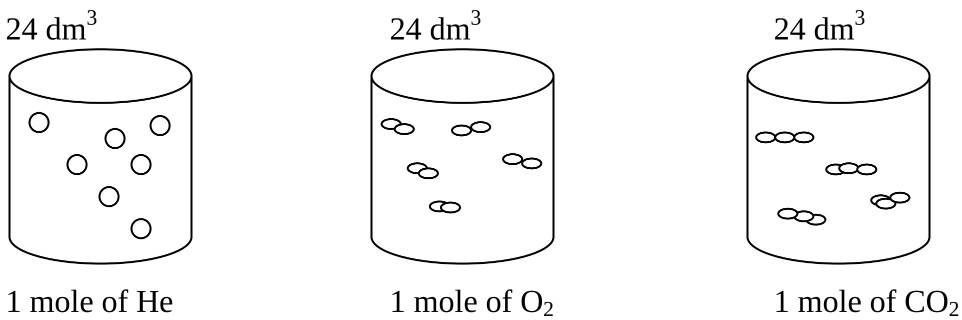
   Mass of a carbon dioxide molecule = 44/ 6.02×10^{-23}

   $= 7.31 \times 10^{-23}$ g.

MOLAR VOLUME OF GASES

A gas unlike solid or a liquid, does not have a fixed shape or a fixed volume. Both the shape and volume of gas change with the container in which it is stored. A gas usually has a very low density because the molecules of a given mas of gas are far apart. So, most of the volume in a gas is empty space. The molecules in a gas are in rapid and random motion. They exert very little or no attractive forces on each other.

When one mole of Hydrogen is introduced into a container by a frictionless piston at 25^oC and 1 atmospheric pressure (room temperature and pressure, R.T.P), the gas occupies a volume of 24.0 dm^3. As the volume occupied by one mole of gas is called the **molar volume** of the gas, the molar volume of hydrogen at R.T.P is $24.0 dm^3\ mol^{-1}$. The results are the same whether one mole of helium gas, one mole of oxygen gas or one mole carbon dioxide gas is introduced into the container. In fact the molar volume of any gas at R.T.P is $24.0\ dm^3\ mol^{-1}$.



Avogadro's law

Avogadro's law is an important law applied to all gases. It states that equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

All the experimental results suggest that the molar volumes of different gases are the same at a given temperature and pressure. Therefore, the number of moles of a given volume of gas can be found by the following formula.

Number of moles of a gas = volume of the gas/ molar volume

It can be established that an equal number of moles of gases occupy equal volume at the same temperature and pressure. However, gas volumes vary with temperature and pressure. Fig summarizes the interconversions involving number of moles.

$$\text{Mass} \xleftrightarrow{\text{Molar Mass}} \text{Moles} \xleftrightarrow{\text{Avogadro's Number}} \text{Atoms}$$

Moles (mol) x Molar Mass (g/mol) = Mass (g)

Moles (mol) x Avogadro's Number (atoms/mol) = atoms

$$\frac{\text{Mass (g)}}{\text{Molar Mass (g/mol)}} = \text{Moles (mol)}$$

$$\frac{\text{Atoms}}{\text{Avogadro's Number (atoms/mol)}} = \text{Moles (mol)}$$

Fig Interconversions involving number of moles

Examples:

1. *Find the volume occupied by 3.55g of chlorine gas at room temperature and pressure. (Molar volume of gas at R.T.P = 24.0 dm^3 mol^{-1}, R.a.m. : cl = 35.5)*
   Solution:
   Molar mass of chlorine gas (Cl_2) = 35.5 x 2 = 71.0 g mol^{-1}
   Number of moles of Cl_2 = 3.55/ 71.0
   = 0.05 mol
   Volume of Cl_2 = Number of moles of Cl_2 = number of moles of Cl_2 x molar volume
   = 0.05 x 24.0
   = 1.3 dm^3
2. *Find the number of molecules in 4.48 cm^3 of carbon dioxide gas at standard tempersture and pressure.*
   (Molar volume of gas at S.T.P. = 22.4 dm^3 mol^{-1} avogadro constant = 6.02 x 10^{23})
   Solution
   Molar volume of carbon dioxide gas at S.T.P = 22.4 dm^3 mol^{-1}
   Number of moles of CO_2 = 4.48/ 22400
   = 2 x 10^{-4} mol
   Number of molecules = 2 x 10^{-4} x 6.02 x 10^{23}
   = 1.204 x 10^{20}
3. *The molar volume of nitrogen gas is found to be 24.0 dm^3 mol^{-1} at room temperature and pressure. Find the density of nitrogen gas. (R.a.m. : N= 14.0)*
   Solution

Molar mass of nitrogen gas (N_2) = 14.0 + 14.0
= 28.0 g mol^{-1}

Density = mass/ volume

Density of nitrogen gas = 28.0/ 24.0
= 1.167 g dm^{-3}

4. *1.6g of a gas occupies 1.2 dm^3 at room temperature and pressure. What is the relative molecular mass of the gas?*
   (Molar volume of gas at R.T.P. = 24.0 dm^3 mol^{-1})

   Solution

   Number of moles of the gas = 1.2/ 24.0
   = 0.05 mol

   Molar mass of the gas = 1.6/ 0.05
   = 32g mol^{-1}

   Relative molecular mass of the gas = 32 (no unit)

CHEMICAL EQUATIONS AND STOICHIOMETRY

Formulae of compounds

The chemical formula of a compound tells us how many different types of elements are present in it. It cannot be determined *experimentally*. By utilizing the mole concept, we can find out the number of its constituent atoms or ions present in a formula unit of the compound.

There are three type of formulae commonly used in chemistry, they are:

1. The empirical formula
2. The molecular formula
3. The structural formula

The **empirical formula** of a compound is the formula which shows the simplest *whole number ratio* of the atoms or ions present in the compound. This formula does not represent the actual number of atoms or ions present in the compound.

The **Molecular formula** of a compound is the formula which shows the *actual number* of each kind of atoms present in a molecule of a compound

The **Structural formula** of a compound is the formula which shows *its component particles are joined together* within the compound. Table illustrates the different types of formulae of some compounds.

Compound	***Empirical formula***	***Molecular formula***	***Structural Formula***
Carbon dioxide	CO_2	CO_2	
Water	H_2O	H_2O	
Methane	CH_4	CH_4	

Ethane	CH_3	C_2H_6	
Ethene	CH_2	C_2H_4	
Ethanoic acid	CH_2O	CH_3COOH	
Glucose	CH_2O	$C_6H_{12}O_6$	
Sulphuric (VI) acid	H_2SO_4	H_2SO_4	
Sodium Fluoride*	NaF	Not applicable	
Ammonium Chloride*	NH_4Cl	Not applicable	

Determination of Empirical formulae

From Combustion Data

The Empirical formula of a compound states the simplest ratio of its component atoms. Infact, this ratio can be determined from the combustion data of the compound.

Examples:

A hydrocarbon is burnt completely in excess oxygen. It is found that 1.00g of hydrocarbon gives 2.93g carbon dioxide and 1.80g water. Find the empirical formula of the hydrocarbon.

(R.a.m. : H= 1.0, C = 1.0, O = 16.0)

Solution

The relative molecular mass of CO_2 = 12.0 + 2 x 16.0

= 44.0

Mass of carbon in 2.93g CO_2 = 2.93 x 12.0/ 44.0

= 0.80g

The relative molecular mass of H_2O = 2 x 1.0 + 16.0

= 18.0

Mass of hydrogen in 1.80g H_2O = 1.80 x 2.0 / 18.0

= 0.20 g

Let the empirical formula of the hydrocarbon be C_xH_y , the simplest ratio of x and y can be determined by using the following table:

	Carbon	***Hydrogen***

Mass(g)	*0.80*	*0.20*
Number of moles	*0.8/12.0 =0.0667*	*0.2/ 1.0 = 0.20*
Relative number of moles	*0.0667/ 0.0667 = 1*	*0.2/ 0.0667 = 3*
Simple mole ratio	*1*	*3*

Therefore, the empirical formula of the hydrocarbon is CH_3

Example

A compound X is known to contain carbon, hydrogen and Oxygen only. When it is burnt completely in excess oxygen, carbon dioxide and water are given out as the only products. It is found that 0.46 g of X gives 0.88g carbon dioxide and 0.54 g water. Find the empirical formula of compound X

Solution

Mass of compound X= 0.46 g

Mass of carbon in X = 0.88 x 12.0/ 44.0

= 0.24 g

Mass of hydrogen in X = 0.54 x 2.0/18.0

= 0.06 g

Mass of Oxygen in X = 0.46 – 0.24 – 0.06

= 0.16g

Let the empirical formula of X be $C_x H_yO_z$.

	Carbon	**Hydrogen**	**Oxygen**
Mass (g)	0.24	0.06	0.16
Number of moles	0.24/ 12.0 = 0.02	0.06/ 1.0 = 0.06	0.16/ 16.0 = 0.01
Relative number of moles	0.02/ 0.01 = 2	0.06/0.01 =6	0.01/ 0.01 = 1
Simplest mole ratio	2	6	1

Therefore, the empirical formula of X is C_2H_6O.

From Composition by Mass

Besides the combustion data, the empirical formula can also be determined from the composition by mass of compound.

Example:

Compound A contains carbon and Hydrogen atoms only. It is found that the compound contains 75% carbon by mass. Determine its empirical formula

(R.a.m. : H=1.0, C= 12.0)

Solution

Let the empirical formula of the hydrocarbon be C_xH_y and the mass of the compound be 100g.

Mass of carbon in the compound = 75g

Mass of Hydrogen in the compound = 100 – 75 = 25 g

	Carbon	**Hydrogen**
Mass	75	25
Number of moles	75/12.0 = 6.25	25/1.0 = 25
Relative number of moles	6.25/ 6.25 =1	25/ 6.25 = 4
Simplest mole ratio	1	4

Therefore, the empirical formula of the hydrocarbon is CH_4

The percentages by mass of Phosphorus and chlorine in a sample of Phosphorus Chloride are 25.55% and 77.45% respectively. Find the empirical formula of the chloride. Find the empirical formula of the chloride (R.a.M, P= 31, Cl= 35.5)

Solution

Let the mass of phosphorus chloride be 100g

Mass of Phosphorus in the compound = 22.55 g

Mass of chlorine in the Compound = 77.45 g

	Phosphorus	**Chlorine**
Mass (g)	22.55	77.45
Number of moles	22.5/ 31.0 = 0.727	77.45/ 35.5 = 2.182
Relative number of moles	0.727/ 0.727 =1	2.182/ 0.727 = 3
Simplest mole ratio	1	3

Therefore, the empirical formula of the phosphorus Chloride is PCl_3

Determination of Molecular Formulae

The Molecular formula of a covalent compound shows the actual number of atoms that each element found in the molecule of the compound contains. It is a simple *whole number multiple* of the empirical formula of the compound. For instance the empirical formula of a compound is CH_2, the molecular formula of the compound may be CH_2, C_2H_4, C_3H_6, C_4H_8, etc.

Molecular formula = (Empirical formula)$_n$

From Composition by mass

Knowing the *empirical formula* and the *relative molecular mass* of the compound, the molecular formula of the compound can be worked out.

Examples

1. A hydrocarbon is burnt completely in excess oxygen. It is found that 5.0g of the hydrocarbon gives 14.6 g carbon dioxide and 9.0g of water. Knowing that the relative molecular mass of hydrocarbon is 30.0, determine the molecular formula of the compound.

(R.a.m, : H=1, C= 12, O=16)

Solution

Let the empirical formula of the hydrocarbon be C_xH_y

Mass of the carbon in the compound = 14.6g x 12/ 44

= 4.0 g

Mass of hydrogen in the compound = 9.0 x 2.0/18

= 1g

	Carbon	**Hydrogen**
Mass (g)	4.0	1.0
Number of moles	4.0/12.0 = 0.333	1/1 = 1
Relative number of moles	0.333/ 0.333 = 1	1/ 0.333 =3
Simplest mole ratio	1	3

The empirical formula of the hydrocarbon is CH_3.

Let the molecular formula of the hydrocarbon be $(CH_3)_n$.

Relative molecular mass of $(CH_3)_n$= 30.0

n (12.0 + 1.0 x 3) = 30.0

n = 2

Therefore, the molecular formula of the hydrocarbon is either $(CH_3)_2$ or C_2H_6

2. Compound X is known to contain 44.44% of carbon, 6.18% of hydrogen and 49.38% of oxygen by mass. A typical analysis found that it has a relative molecular mass of 162.0. find its molecular formula (R.a.m. : H= 1, C= 12, O= 16)

Solution

Let the empirical formula of the compound X be $C_xH_yO_z$ and the mass of the compound be 100g

Mass of the carbon in the compound = 44.44g

Mass of the compound = 6.18 g

Mass of oxygen in the compound = 49.38g

	Carbon	***Hydrogen***	***Oxygen***
Mass (g)	44.44	6.18	49.38
Number of moles	44.44/12.0 = 3.70	6.18/1.0 =6.18	49.38/16 =3.08
Relative number of moles	3.70/ 3.08 = 1.2	6.18/ 3.08 = 2	3.08/3.08 =1
Simplest mole ratio *	6	10	5

*Note: the ratio of 6 : 10 : 5 is obtained by multiplying the relative number of moles by 5

The empirical formula of compound X is $C_6H_{10}O_5$

Let the molecular formula of the compound X be $(C_6H_{10}O_5)_n$

Relative molecular mass $(C_6H_{10}O_5)_n$

n (12 x 6 + 1 x 10 + 16 x 5) = 162.0

n = 1

therefore, the molecular formula of compound X is $C_6H_{10}O_5$

Water of crystallization Derived from Composition by Mass

Examples of salts that contain water of crystallization

Hydrated salt	***Anhydrous salt***
$CUSO_4.5H_2O$ – blue crystals	Anhydrous $CUSO_4$- white powder
Na_2CO_3. $10H_2O$ – colourless crystal	Anhydrous Na_2CO_3 – white crystals
$CoCl_2$. $2H_2O$ – Pink crystals	Anhydrous $CoCl_2$ – blue crystals

Examples

The chemical formula of hydrated copper (II) sulphate is known to be $CUSO_4$. xH_2O. it is found that the percentage of the water by mass in the compound is 36%. Find x (R.a.m : H= 1, O = 16, S = 32.1, Cu = 63.5)

Solution

Formula mass of $CUSO_4.xH_2O$ = 63.5 + 32.1 + 16 x 4 + (1 x 2 + 16)$_x$

= 159.6 + 18x

Mass of water of crystallization = 18x g

18x/ 159.6 + 18x = 36/ 100

1800x = 5745.6 + 648x

1152x = 5745.6

x =4.99 = 5

therefore the chemical formula of hydrated copper (II) sulphate is $CUSO_4.5H_2O$

Composition By Mass Derived from Formula

Knowing the *molecular formula* of a compound, the composition by mass of each of the constituent elements can easily be worked out.

Examples

The chemical formula of ethanoic acid is CH_3COOH. Calculate its composition by mass of carbon, hydrogen and oxygen.

(R.a.m : C = 12, H = 1, O = 16)

Solution

Molecular mass of CH_3COOH = 12 x 2 + 1 x 4 +16 x 2

= 60

% by mass of C = 12 x 2/ 60 x 100%

= 40%

% by mass of H = 1 x 4 / 60 x 100%

= 6.67%

% by mass of O = 16 x 2 / 60 x 100%

= 53.33%

The percentages by mass of Carbon, hydrogen and oxygen are 40%, 6.67% and 53.33% respectively.

CHEMICAL EQUATIONS

When chemicals react with each other, new substances are formed. A chemical reaction can be summarized by means of an equation which states the reactants and the products of the reaction. A **word equation** summarizes the experimental results in words

Magnesium + hydrochloric acid $\longrightarrow$ magnesium chloride + hydrogen

However, such equation does not tell the relative amount of reactants and products in the reaction. The best way to summarize it is by a **Chemical equation**

$Mg\ (s) + 2HCl\ (aq) \longrightarrow MgCl_2(aq) + H_2\ (g)$

A chemical equation is a statement, in formulae, which shows the relative number of particles involved in a chemical reaction

In the case of the above equation, it tells us that 1 atom of Mg reacts with 2 molecules of HCl to give 1 $MgCl_2$ formula unit and 1 H_2 molecule. The number appearing before the chemical formulae in the equation are called **Stoichiometric coefficients**

Stiochiometry refers to the relative number of moles of substances involved in a chemical reaction

Calculations Based on Equations

Calculations involving mole ratio of reacting masses

A balanced chemical equation gives us all the information about a chemical reaction. If the quantities of the reactants are known, we can work out the amounts of the products formed and vice- versa.

Examples

Calculate the mass of copper formed when 12.45 g of copper (II) oxide is completely reduced by hydrogen. (R. a. m. : H= 1, O= 16, Cu = 63.5)

Solution

$CuO\ (s) + H_2\ (g) \longrightarrow Cu\ (s) + H_2O\ (l)$

As the mole ratio of Cu: CuO is 1:1, the number of moles of Cu formed is the same as the number of moles of CuO reduced.

Number of moles of CuO reduced = 12.45 / 63.5 + 16

= 0.157

Number of moles of Cu formed = 0.157

i.e Mass of Cu/ 63.5 = 0.157

Mass of Cu = 9.97g

Therefore, the mass of copper formed in the reaction is 9.97 g

Sodium hydrogencarbonate decomposes according to the following equation

$$2NaHCO_3\ (s) \longrightarrow Na_2\,CO_3(s) + CO_2\ (g) + H_2O(l)$$

In order to obtain 240cm^3 of CO_2 at room temperature and pressure, what is the minimum amount of sodium hydrogencarbonate required? (R.a.m.: H= 1, C = 12 , O= 16, Na = 23 molar volume of gas at RTP = 24dm^3 mol^{-1})

Solution

Mole ratio of $NaHCO_3 : CO_2 = 2 : 1$

Number of moles of CO_2 formed = 240 / 24000 = 0.01 mol

Number of moles of $NaHCO_3$ required = 0.01 x 2 = 0.02 mol

Mass of $NaHCO_3$ required = 0.02 x (23 + 1 + 12 + 16 x 3)

= 0.02 x 84.0

= 1.68g

Therefore, the minimum mass of sodium hydrogencarbonate required is 1.68g

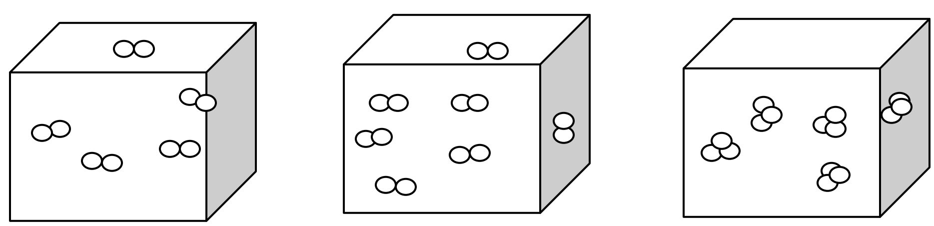
Calculations involving volume ratio of gases

Besides mass analysis, the formulae of gaseous compounds can also be obtained from volumes involved in the gaseous reactions. Consider the equation for the reaction in the **Haber process,**

$$N_2(g) + 3H_2(g) \longleftrightarrow 2NH_3(g)$$

Nitrogen reacts with hydrogen in a mole ratio of 1: 3. When volumes of the reacting gases are measured, it is found that nitrogen reacts with hydrogen in a volume ratio of 1: 3 too. These relationships between mole ratio and volume ratio of reacting gases are stated in Avogadro's law.

Avogadro's law states that an equal number of moles of any gas occupy the same volume under the same conditions of pressure and temperature.



Nitrogen (N_2) Hydrogen (H_2) Ammonia (NH_3)

Equal volumes of all gases contain the same number of the molecules, at the same temperature and pressure

	$N_2(g)$	$+ 3H_2(g) \longleftrightarrow$	$2NH_3(g)$
	1 molecule :	3 molecules :	2 molecules
i.e.	1 mole :	3 moles :	2 moles
mole ratio	1 :	3 :	2
	24 dm^3 :	72 dm^3 :	48 dm^3
i.e	1 volume :	3 volumes :	2 volumes
volume ratio	1 :	3 :	2

Example

Calculate the volume of carbon dioxide formed when 20 cm^3 of ethane and 70 cm^3 of oxygen are exploded, assuming all volumes are measured at room temperature and pressure.

Solution

$2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(l)$

2moles : 7moles : 4 moles : 6 moles (from equation)

2 volumes : 7 volumes : 4 volumes : - (by avogadro's)

$20cm^3$ $70\ cm^3$ $x\ cm^3$ –

It can be judged from the equation that the mole ratio of $CO_2 : C_2H_6$ is 4 : 2, the volume ratio of $CO_2 : C_2H_6$ should also be 4: 2

x/ 20 = 4/2

$x = 40\ cm^3$

therefore, the volume of CO_2 (g) formed is $40\ cm^3$

Example

10 cm^3 of gaseous hydrocarbon was mixed with 80 cm^3 of oxygen which was in excess. The mixture was exploded and then cooled. The volume left was 70 cm^3. Upon passing the resulting gaseous mixture through concentrated sodium hydroxide solution (to absorb carbon dioxide), the volume of the residual gas became 50cm^3. Find the molecular formula of the hydrocarbon.

Solution

Let the molecular formula of the hydrocarbon be C_xH_y.

Volume of hydrocarbon reacted = $10\ cm^3$

Volume of O_2(g) left unreacted = $50cm^3$

Volume of O_2 (g) reacted = 80 – 50 = $30\ cm^3$

Volume of CO_2(g) formed = 70 – 50 = $20\ cm^3$

$C_xH_y + (x + y/4)\ O_2 \longrightarrow xCO_2 + y/2\ H_2O$

1 mole : x + y/4 mole : x mole

$10\ cm^3$ $30\ cm^3$ $20\ cm^3$

Volume of CO_2 : volume of $C_x H_y$ = 20 : 10

= x : 1

x = 2

Volume of O_2 : volume of C_xH_y = 30 : 10

= (x + y/4) : 4

x + y/4 = 3

as x =2, 2 + y/4 = 3

$y = 4$

therefore, the molecular formula of the hydrocarbon is C_2H_4.

Simple titration

Volumetric analysis is a method used by chemists to determine the quantity or concentration of a substance by measuring the *volumes* some solutions. **Titration** is an important step in volumetric analysis.

The following relationship is important in calculation involving titrations

Number of moles = Molarity* x Volume

(mol/dm^{-3}**)** **(**dm^3**)**

***Note:** Molarity (or molar concentration) of a solution is the number of moles of the solution

Acid- Base Titration

1. Titration with indicator

The general equation for acid base titration is:

$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

Examples

25cm^3 of sodium hydroxide solution were titrated against 0.067 M of sulphuric acid using methyl orange as indicator. The indicator changed from yellow to red when 22.5 cm^3 of H_2SO_4 had been added. Calculate the concentration of sodium hydroxide solution.

Solution

$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$

Mole ratio of NaoH (aq) : H_2SO_4(aq) = 2 : 1

½ x number of moles of NaOH = number of moles of H_2SO_4

Number of moles of H_2SO_4 (aq) = 0.067 x 22.5/ 1000

= 1.508 x 10^{-3} mol

Number of moles of NaOH(aq) = 2 x 1.508 x 10^{-3} mol

= 3.016 x 10^{-3} mol

i.e Concentration of NaOH (aq) = 3.016 x 10^{-3} x 1000/ 25

= 0.121 M

2.52 g of a pure dibasic acid of formula mass 126.0 is dissolved in water and made up to 250 cm^3 in a volumetric flask. 25.0 cm^3 of this solution is found to neutralize 28.5 cm^{3}of a sodium hydroxide solution

a. what is the concentration of the acid solution?

b. if the dibasic acid is represented by H_2X, write an equation for the reaction between the acid and sodium hydroxide.

c. what is the concentration of the sodium hydroxide solution?

Solution

a. Number of moles of acid = 2.52 / 126 = 0.02 mol

   Molarity of acid solution = 0.02 / 250 x 10^{-3} = 0.08 M

b. H_2X (*aq*) + 2NaOH (*aq*) $\longrightarrow$ Na_2X(*aq*) + $2H_2O$ (*l*)

c. Number of moles of H_2X = ½ x Molarity of NaOH x 28.5/ 1000
   Molarity of NaOH = 0.14 M
   Therefore, the molarity of sodium hydroxide solution is 0.14 M

Titration without an indicator

Apart from using indicators, the end point of titration can also be detected by the following; the changes in pH value, temperature or electrical conductivity during the course of a reaction.

By change in pH value

A pH meter is a device that gives direct reading of the pH value from an electrochemical cell. It is highly sensitive to the concentration of hydrogen ion.

A pH meter

Example

0.186g of a sample of hydrated sodium carbonate $Na_2CO_3.nH_2O$ was dissolved in $100cm^3$ of distilled water in a conical flask. 0.10 M hydrochloric acid was added from the burrette, $2cm^3$at a time. The pH value of the solution was measured by a pH meter. The result was recorded and shown in the following figure

Calculate the value of n in $Na_2CO_3.nH_2O$

(R.a.m.: H= 1, C = 12, O = 16, Na = 23)

Solution

There is a sudden drop in the pH value of the solution (from 8 to 3) with the end point at 30 cm^3.

$Na_2CO_3.\ nH_2O\ (s) + 2HCl\ (aq) \longrightarrow 2NaCl(aq) + CO_2(g) + (n + 1)H_2O(l)$

Number of moles of $Na_2CO_3.\ nH_2O$ x ½ x No. of moles of HCl

i.e. 0.186/ 23 x 2 + 12 +16 x 3 + 18n = ½ x 0.1 x 30 x 10^{-3}

therefore, the formula of hydrated sodium carbonate is $Na_2CO_3.H_2O$

By change in temperature

As neutralization is an exothermic reaction, the end point of a titration can be detected by the temperature change of the reaction mixture. This type of titration is also known as **thermometric titration.**

Example

$5cm^3$of 0.50M sulphuric acid was added to $25cm^3$ of potassium hydroxide solution. The mixture was stirred after that and the highest temperature was recorded. The experiment was repeated with different volumes of sulphuric acid. The set up of the experiment and the results were as follows.

Volume of H_2SO_4 Added	*Temperature*
0	20.0
5	21.8
10	23.4
15	25.0
20	26.5
25	25.2
30	24.0

a. *Draw a graph of temperature against volume of sulphuric acid added.*
b. *Calculate the molarity of the potassium hydroxide solution.*
c. *Explain why the graph rises to a maximum and then falls*

b. $2KOH(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$

2 moles : 1 moles

From the graph, $20cm^3$ of H_2SO_4 is the end point of the titration.

Number of moles of H_2SO_4 = 0.5 x 20/ 1000

= 0.01 mol

From the equation,

Mole ratio of KOH(*aq*) : H_2SO_4(*aq*) = 2:1

Number of moles of KOH(*aq*) = 2 x 0.01

= 0.02 mol

Concentration of KOH (*aq*) = 0.02 x 1000/ 25 = 0.8M

d. Neutralization is an exothermic reaction. When increasing amounts of sulphuric acid are added and reacts with potassium hydroxide, the temperature rises. The temperature reaches to a maximum value at which the end point of the titration is reached. After that, any excess sulphuric acid added will cool down the reacting solution, causing the temperature to drop

THE ELECTRONIC STRUCTURE OF ATOMS

By 1911, it was known that an atom is composed of a nucleus of protons and neutrons surrounded by a cloud of rapidly moving electrons. But how are theses electrons arranged in the space around the nucleus?

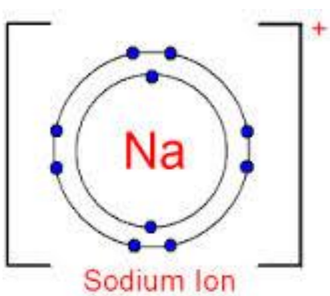
In 1913, Niels Bohr suggested a model for the electronic structure of atoms. He proposed that electrons in an atom are moving around the nucleus in certain **orbits,** like the orbits of the planets around the sun.

The electron orbits are now called **electron shells** which exist in certain regions surrounding the central nucleus. Table below lists the electron shells of atom and the maximum number of electrons which they can hold.

The electron shells

Shell Number, n	***Maximum number of Electrons (= $2n^2$)***
1	2
2	8
3	18
4	32

The arrangement of electrons in the various shells is known as the **Electronic Configuration** of the atoms



Arrangement of electrons in a sodium atom

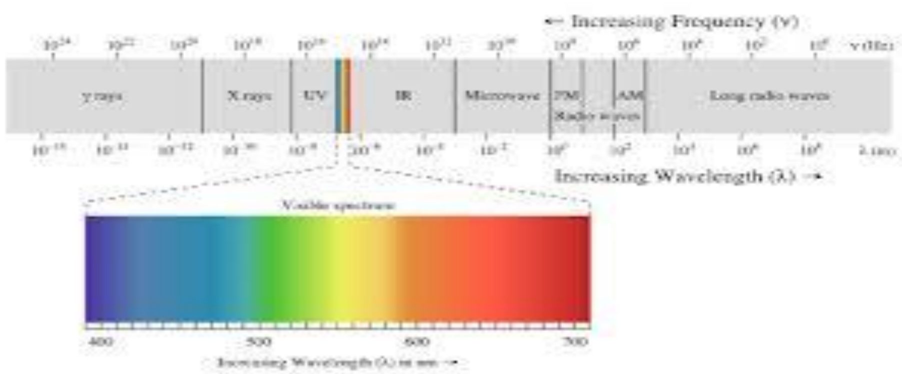
The Electromagnetic Spectrum

Bohr's explanation of the electronic configuration of atoms was based upon experimental evidence concerning the nature of light and the emission of light by different elements.

It should be noted that visible light spectrum only represents a narrow part of the **electromagnetic spectrum** which ranges from 10^{-13}m to 10^{6} m in wavelength. Electromagnetic waves radiate at the same speed as light *c= (3 x 10⁸m s⁻¹)*, as light. The frequency, v of a radiation corresponds to a spectral line of wavelength, by the following relationship:

v = c/ wavelength

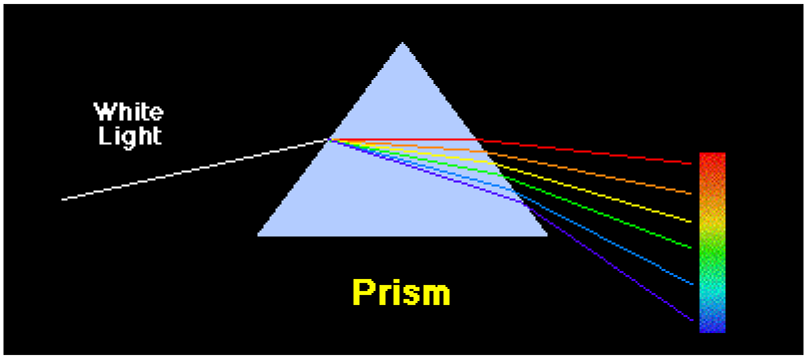
to analyse the spectrum of a substance, a special instrument called a spectrometer is required.



The electromagnetic spectrum; from the shorter wavelength cosmic rays to a longer electric waves.

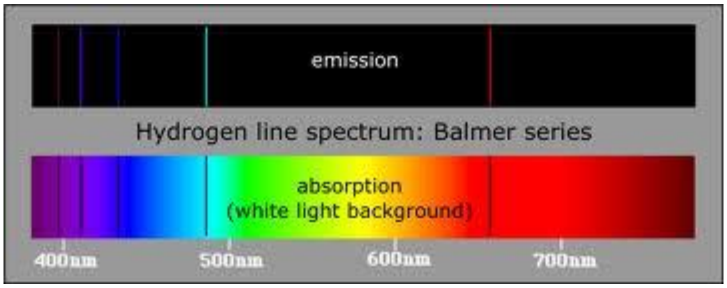
Continuous and line spectra

When white light passes through a prism and is allowed to fall onto a screen, a rainbow of different colours is formed. The spectrum of colours is composed of visible light of all wavelengths – violet, blue, green, yellow, orange and red. The colours merge smoothly into one another in an unbroken band. This describes why we see rainbow in the sky when sun reappears after a shower, the falling rain drops act as prisms and disperse the sunlight. The different colour produced by a prism or by raindrops represent different amount of radiant energy. In the spectrum, no separate line can be found, this is called a **Continuous spectrum.**

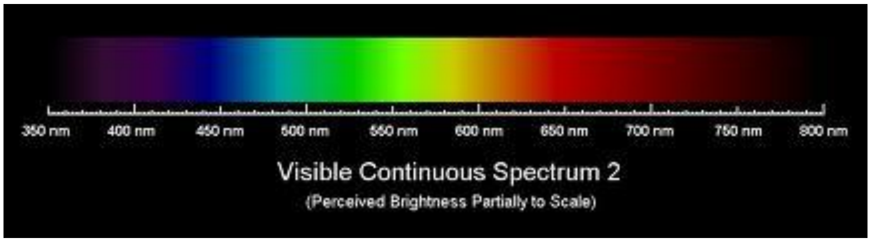


Continuous spectrum of a white light,

On the contrary, when light is from a gas discharge tube containing Hydrogen is allowed to pass through a prism and analysed, some coloured lines are observed. This spectrum differ from the continuous spectrum of light. The pattern recorded comprises only a few lines, each which corresponds to discrete wavelength in the visible region. This hydrogen spectrum is called an **emission line spectrum.** Hence, spectra can be classified into continuous or line spectra depending on their appearances.



The line spectrum of hydrogen



The continuous spectrum in the visible region

The line spectrum emitted by the atomic hydrogen in the visible region

The Emission Spectrum of Atomic Hydrogen.

When Hydrogen gas at low pressure is subjected to a high potential diference, it glows with reddish- pink light. If the emitted light is analysed by viewing through a spectrometer, several groups or series of discrete lines are observed.

Atomic Hydrogen spectrum

Wavelengths of the spectral series for atomic hydrogen

N	***1***	***2***	***3***	***4***
	Lyman (Near UV)	Balmer (visible)	Paschen (Infra-red)	Brackett (Far Infra- red)
2	121.6	-	-	-
3	102.6	656.3	-	-
4	97.3	486.1	1875.1	-
5	95.0	434.0	1281.8	4050

The spectrum is composed of a series of lines dispersed in various regions of the electromagnetic spectrum. These series are named after their discoverers, such as Lyman, Balmer

Interpretation of the Atomic Hydrogen Spectrum

In each series of lines, it can be observed that the intensity of the spectral lines and distances between two successive lines decrease as the frequency increases (wavelength decreases) until the lines converge to form a **Continuum**

Bohr conceived that the electrons in an atom are moving constantly around the nucleus in circular orbits in the same way that the planets revolve around the sun. According to this model, each orbit is associated with a definite amount of energy. Energy is needed to excite an electron from a lower orbit (i.e lower energy level) to a higher orbit (i.e higher energy level).

From the characteristics of the emission spectrum of atomic hydrogen, Bohr suggested a simple model for the structure of an atom known as **quantum theory.** He put forward three basic assumptions in this model.

1. An electron in an atom can only exist in certain states characterized by definite energy levels. The energy of an electron can only change by some definite whole number multiple of a unit called **quantum,** e.g 1,2,3… etc Quanta cannot have non- intergral values such as 1.5 or 2.2. This means that the energy of an electron cannot change continuously. Instead the energy of an electron is quantized.
2. As the energy of an electron is quantized, the radius of the orbit should also be quantized. Different orbits represent different energy levels. The energy level with the lowest energy

is the one close to the nucleus and that with the higher energy is one further from the nucleus.

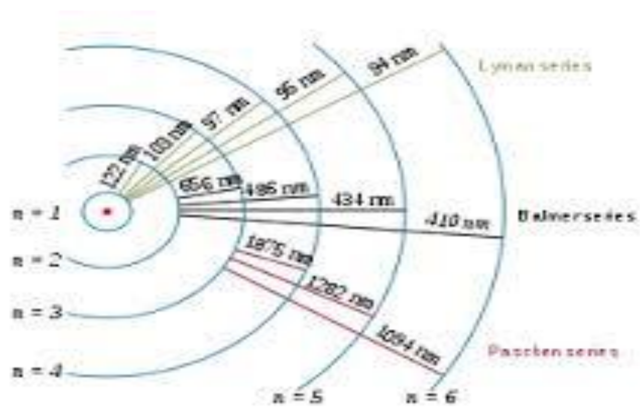
3. When an electron moves from one orbit to another, it must emit or absorb a definite amount of energy to bring it to that orbit. When the electron jumps from a higher energy state E_2 to a lower energy state E_1, energy is emitted. The energy emitted is related to the frequency of the light recorded in the spectrum by the following equation.

$$\Delta E = E_2 - E_1 = hv$$

where ΔE is the difference in energy between two energy levels

h is the planck's constant having a value of 6.626×10^{-34}Js

and v is the frequency of the light emitted.



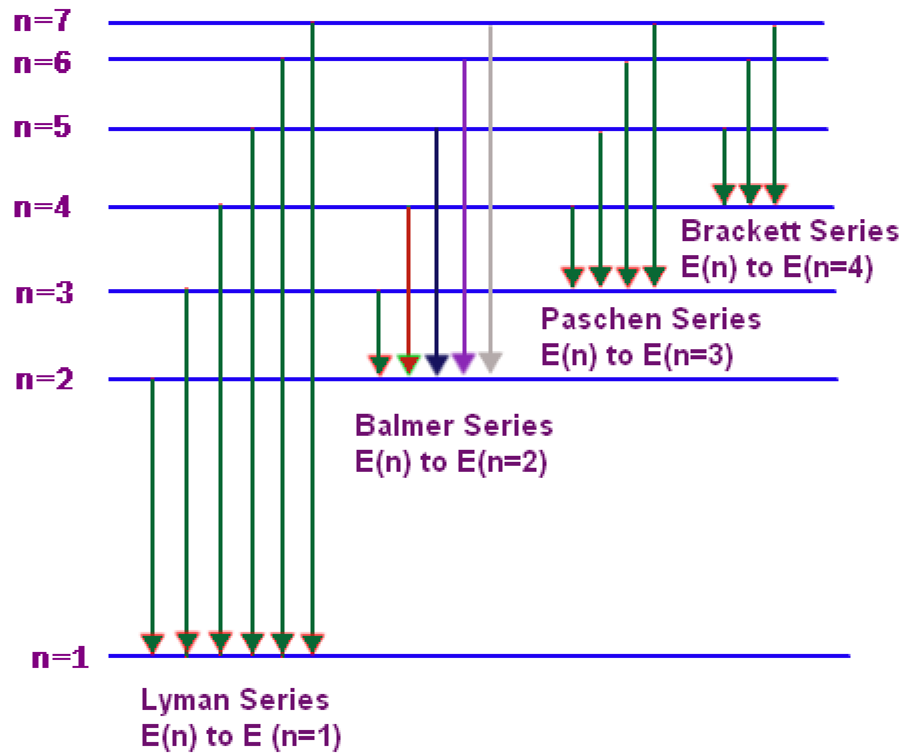
Energy levels of an atom of hydrogen showing the possible transitions of electrons between different energy levels in its emission spectrum.

Similarly, energy has to be absorbed before the electron can jump from a lower energy state (E_2) to a higher energy state (E_3). The amount of energy absorbed is also quantized to the frequency absorbed in the spectrum by the same relationship.

$$\Delta E = E_3 - E_2 = hv^1$$

In brief, when sufficient energy is supplied to an atom, it is possible to promote or excite an electron from a lower energy to a higher energy. This process is called **excitation.** Since electron is unstable at a higher energy, it will fall back to a lower energy level. Excess energy is emitted as radiation.

Bohr gave each energy level a quantum number. He assigned quantum number 1 to the energy level of the lowest energy content. The next energy level has quantum number 2 and so on.



Relationship between electron transition in an atomic hydrogen and its line spectrum.