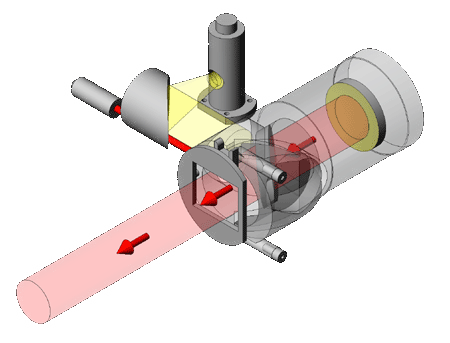
# CHAPTER

1 BASIC CONCEPTS



Animation 1.1: Spectrometer

Source & Credit: [gascell](http://www.gascell.com/htmls/primer.htm)

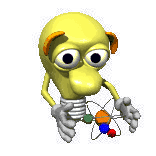
## 1.1 ATOM

Long time ago, it was thought that matter is made up of simple, indivisible particles. Greek philosophers thought that, matter could be divided into smaller and smaller particles to reach a basic unit, which could not be further sub-divided. Democritus (460-370 B.C.) called these particles atomos, derived from the word “atomos” means indivisible. However, the ideas of Greek philosophers were not based on experimental evidences.

In the late 17th century, the quantitative study of the composition of pure substances disclosed that a few elements were the components of many diferent substances. It was also investigated that how, elements combined to form compounds and how compounds could be broken down into their constituent elements.

In 1808, an English school teacher, John Dalton, recognized that the law of conservation of matter and the law of deinite proportions could be explained by the existence of atoms. He developed an atomic theory; the main postulate of which is that all matter is composed of atoms of diferent elements, which difer in their properties.

**Atom is the smallest particle of an element, which can take part in a chemical reaction.** For example, He and Ne, etc. have atoms, which have independent existence while atoms of hydrogen, nitrogen and oxygen do not exist independently.



Animation 1.2: [Atom](http://www.123gifs.eu/atoms/)

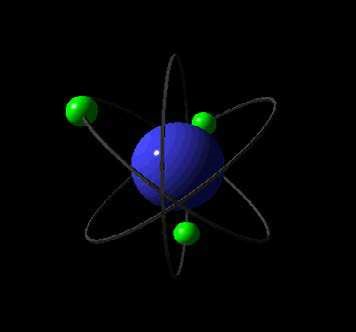
Source & Credit[: 123gifs](http://www.123gifs.eu/atoms/)

The modern researches have clearly shown that an atom is further composed of subatomic particles like electron, proton, neutron, hypron, neutrino, anti-neutrino, etc. More than 100 such particles are thought to exist in an atom. However, electron, proton and neutron are regarded as the fundamental particles of atoms.

A Swedish chemist J. Berzelius- (1779-1848) determined the atomic masses of elements. A number of his values are close to the modern values of atomic masses. Berzelius also developed the system of giving element a symbol.

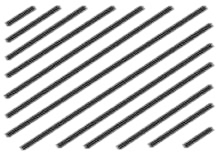
**1.1.1 Evidence of Atoms**

It is not possible actually to see the atoms but the nearest possibility to its direct evidence is by using an electron microscope. A clear and accurate image of an object that is smaller than the wavelength of visible light, cannot be obtained. Thus an ordinary optical microscope can measure the size of an object upto or above 500 nm (lnm = 10-9m). However, objects of the size of an atom can be observed in an electron microscope. It uses beams of electrons instead of visible light, because wavelength of electron is much shorter than that of visible light.



Animation 1.3[:Made of Atom](http://imgur.com/gallery/Qjsc9)

Source & Credit: [imgur](http://imgur.com/gallery/Qjsc9)

Fig. (1.1) shows electron microscopic photograph of a piece of a graphite which has been magniied about 15 millions times. The bright band in the igure are layers of carbon atoms.

In the 20th century, X-ray work has shown that the diameter of atoms are of the order 2x10-10 m which is 0.2 nm. Masses of atoms range from 10-27 to 10-25 kg. They are often expressed in atomic mass units (amu) when 1 amu is = 1.661 x 10-27 kg. The students can have an idea about the amazingly small size of an atom from

the fact that a full stop may have two million atoms present in it. Fig (1.1) Electron microscopic

photograph of graphite

**1.1.2 Molecule**

**A molecule is the smallest particle of a pure substance which can exist independently.** It may contain one or more atoms. The number of atoms present in a molecule determines its atomicity. Thus molecules can be monoatomic, diatomic and triatomic, etc., if they contain one, two and three atoms respectively. Molecules of elements may contain one, two or more same type of atoms. For example, He, Cl2, O3, P4, S8. On the other hand, molecules of compounds consist of diferent kind of atoms. For example, HCl, NH3, H2SO4,C6H12O6.

The sizes of molecules are deinitely bigger than atoms. They depend upon the number of atoms present in them and their shapes. Some molecules are so big that they are called macromolecules. Haemoglobin is such a macromolecule found in blood. It helps to carry oxygen from our lungs to all parts of our body. Each molecule of haemoglobin is made up of nearly 10,000 atoms and it is 68,000 times heavier than a hydrogen atom.

**1.1.3 Ion**

**Ions are those species which carry either positive or negative charge.** Whenever an atom of an element loses one or more electrons, positive ions are formed.

A suicient amount of energy is to be provided to a neutral atom to ionize it.

1. → A + e+ -

This A+ is called a cation. A cation may carry +1, +2, +3, etc.charge or charges. The number of charges present on an ion depends upon the number of electrons lost by the atom. Anyhow, energy is always required to do so. Hence the formation of the positive ions is an endothermic process. The most common positive ions are formed by the metal atoms such as Na+, K+, Ca2+, Mg2+ , Al3+, Fe3+, Sn4+, etc. The chapter on chemical bonding will enable us to understand the feasibilities of their formation. When a neutral atom picks up one or more electrons, a negative ion is produced, which is called an anion.

1. +e− → B-

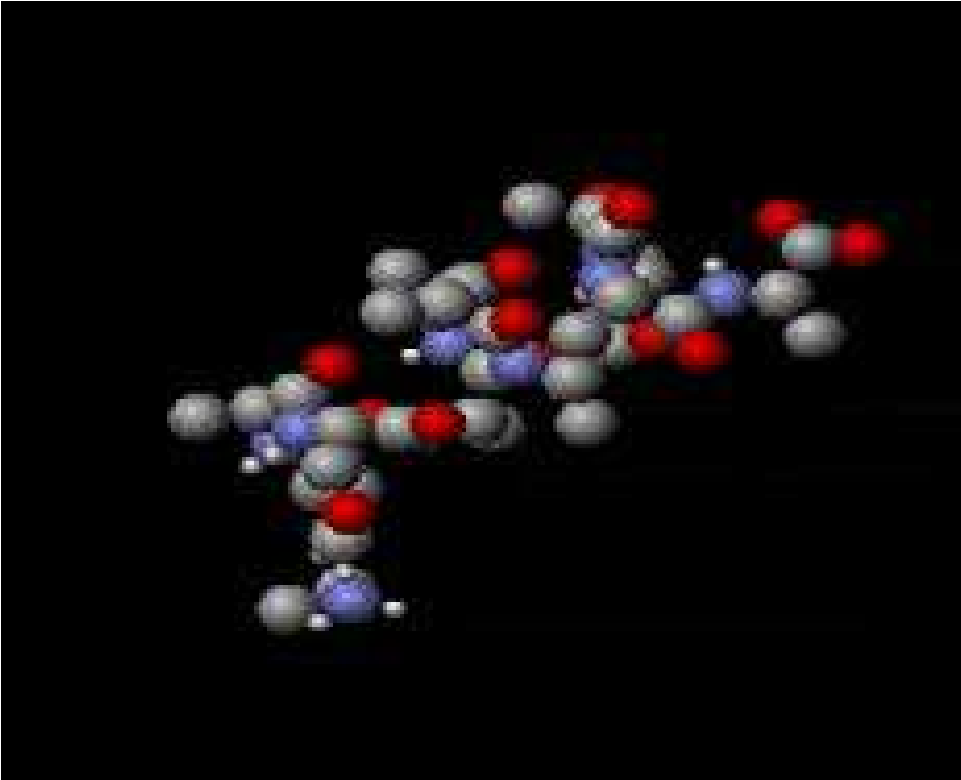
Energy is usually released when an electron is added to the isolated neutral atom, Therefore, the formation of an uninegative ion is an exothermic process.The most common negative ions are

*F*−,Cl ,− *Br*−,*S*2−etc.

The cations and anions possess altogether diferent properties from their corresponding neutral atoms. There are many examples of negative ions which consist of group of atoms like OH-, CO32-,

SO42-, PO43-, MnO41-, Cr2O72-  etc. The positive ions having group of atoms are less common e.g. NH4+ and some carbocations in organic chemistry.

**1.1.4 Molecular Ion**

When an atom loses or gains an electron, it forms an ion. Similarly, a molecule may also lose or gain an electron to form a molecular ion, e.g., CH4+, CO+, N2+ Cationic molecular ions are more abundant than anionic ones. These ions can be generated by passing high energy electron beam or α-particles or X-rays through a gas. The break down of molecular ions obtained from the natural products can give important information about their structure.

Animation 1.4: [Molecules](https://commons.wikimedia.org/wiki/File:Thermally_Agitated_Molecule.gif)

Source & credit: [wikimedia](https://commons.wikimedia.org/wiki/File:Thermally_Agitated_Molecule.gif)

## 1.2 RELATIVE ATOMIC MASS

**Relative atomic mass is the mass of an atom of an element as compared to the mass of an atom of carbon taken as 12.**

The unit used to express the relative atomic mass is called atomic mass unit (amu) and it is 1/12 th of the mass of one carbon atom, On carbon -12 scale, the relative atomic mass of 126*C* is 12.0000

amu and the relative atomic mass of 11 *H* is 1.008 amu. The masses of the atoms are extremely small. We-don’t have any balance to weigh such an extremely small mass, that is why we use the relative atomic mass unit scale.

The relative atomic masses of some elements are given in the following Table (1.1).

### Table (1.1) Relative atomic masses of a few elements

|  |  |  |  |
| --- | --- | --- | --- |
| **Element** | **Relative Atomic Mass (amu)** | **Element** | **Relative Atomic Mass (amu)** |
| H  O  Ne | 1.008  15.9994  20.1797 | Cl  Cu  U | 35.453  63.546  238.0289 |

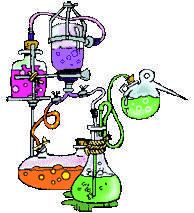
These element have atomic masses in fractions and will be explained in the following article on isotopes.

## 1.3 ISOTOPES

In Dalton’s atomic theory, all the atoms of an element were considered alike in all the properties including their masses. Later on, it was discovered that **atoms of the same element can possess diferent masses but same atomic numbers. Such atoms of an element are called isotopes**. So isotopes are diferent kind of atoms of the same element having same atomic number, but diferent atomic masses. The isotopes of an element possess same chemical properties and same position in the periodic table. This phenomenon of isotopy was irst discovered by Soddy. Isotopes have same number of protons and electrons but they difer in the number of neutrons present in their nuclei.

Carbon has three isotopes written as 126C , 136C , 146C and expressed as C-12, C-13 and C-14. Each of these have 6-protrons and 6 electrons. However, these isotopes have 6, 7 and 8 neutrons respectively.

Similarly, hydrogen has three isotopes 11H , 21H , 31Hcalled protium, deuterium and tritium. Oxygen has three, nickel has ive, calcium has six, palladium has six, cadmium has nine and tin has eleven isotopes.



Animation 1.5: Basic Concepts

Source & Credit: [pixshark](http://pixshark.com/science-moving-animations.htm)

**1.3.1 Relative Abundance of Isotopes**

The isotopes of all the elements have their own natural abundance. The properties of a particular element, which are mentioned in the literature, mostly correspond to the most abundant isotope of that element. The relative abundance of the isotopes of elements can be determined by mass spectrometry.

Table (1.2) shows the natural abundance of some common isotopes.

**Table (1.2) Natural abundance of some common isotopes.**

|  |  |  |  |
| --- | --- | --- | --- |
| **Element** | **Isotope** | **Abundance (%)** | **Mass (amu)** |
| Hydrogen  Carbon  Nitrogen  Oxygen  Sulphur  Chlorine  Bromine | 1 2  H, H  12 13 C, C  14 15 N, N  16 17 18  O, O, O  32 33 34 36  S, S, S, S  36 37  Cl, Cl 79 81  Br, Br | 99.985, 0.015  98.893, 1.107  99.634, 0.366  99.759, 0.037,0.204  95.0, 0.76, 4.22, 0.014  75.53, 24.47  50.54, 49.49 | 1.007825, 2.01410  12.0000, 13.00335  14.00307 15.00011  15.99491, 16.99914, 17.9916  31.97207, 32.97146, 33.96786, 35.96709  34.96885, 36.96590  78.918, 80.916 |

We know at present above 280 diferent isotopes occur in nature. They include 40 radioactive isotopes as well. Besides these about 300 unstable radioactive isotopes have been produced through artiicial disintegration. The distribution of isotopes among the elements is varied and complex as it is evident from the Table (1.2). The elements like arsenic, luorine, iodine and gold, etc have only a single isotope. They are called mono-isotopic elements.

In general, the elements of odd atomic number almost never possess more than two stable isotopes. The elements of even atomic number usually have larger number of isotopes and isotopes whose mass numbers are multiples of four are particularly abundant. For example, l6O, 24Mg, 28Si, 40Ca and 56Fe form nearly 50% of the earth’s crust. Out of 280 isotopes that occur in nature, 154 have even mass number and even atomic number.

**1.3.2 Determination of Relative Atomic Masses of Isotopes by Mass Spectrometry**

**Mass spectrometer is an instrument which is used to measure the exact masses of diferent isotopes of an element.** In this technique, a substance is irst volatilized and then ionized with the help of high energy beam of electrons. The gaseous positive ions, thus formed, are separated on the basis of their mass to charge ratio (m/e) and then recorded in the form of peaks. Actually mass spectrum is the plot of data in such a way that (m/e) is plotted as abscissa (x-axis) and the relative number of ions as ordinate (y-axis).

First of all, Aston’s mass spectrograph was designed to identify the isotopes of an element on the basis of their atomic masses. There is another instrument called Dempster’s mass spectrometer. This was designed for the identiication of elements which were available in solid state.

The substance whose analysis for the separation of isotopes is required, is converted into the vapour state. The pressure of these vapours is kept very low, that is, 10-6 to 10-7 torr. These vapours are allowed to enter the ionization chamber where fast moving electrons arc thrown upon them. The atoms of isotopic element present in the form of vapours, are ionized. These positively charged ions of isotopes of an element have diferent masses depending upon the nature of the isotopes present in them.

The positive ion of each isotope has its own (m/e) value. When a potential diference (E) of 500-2000 volts is applied between perforated accelerating plates, then these positive ions are strongly attracted towards the negative plate. In this way, the ions are accelerated.

These ions are then allowed to pass through a strong magnetic ield of strength (H), which will separate them on the basis of their (m/e) values. Actually, the magnetic ield makes the ions to move in a circular path. The ions of deinite m/e value will move in the form of groups one after the other and fall on the electrometer.

The mathematical relationship for (m/e) is

m/e = H r/2E2

Where H is the strength of magnetic ield, E is the strength of electrical ield, r is the radius of circular path. If E is increased, by keeping H constant then radius will increase and positive ion of a particular m/e will fall at a diferent place as compared to the irst place. This can also be done by changing the magnetic ield. Each ion sets up a minute electrical current.

Electrometer is also called an ion collector and develops the electrical current. The strength of the current thus measured gives the relative abundance of ions of a deinite m/e value.

Similarly, the ions of other isotopes having diferent masses are made to fall on the collector and the current strength is measured. The current strength in each case gives the relative abundance of each of the isotopes. The same experiment is performed with C-12 isotope and the current strength is compared.

This comparison allows us to measure the exact mass number of the isotope Fig. (1.2), shows the separation of isotopes of Ne. Smaller the (m/e) of an isotope, smaller the radius of curvature produced by the magnetic ield according to above equation.

In modern spectrographs, each ion strikes a detector, the ionic current is ampliied and is fed to the recorder. The recorder makes a graph showing the relative abundance of isotopes plotted against the mass number.

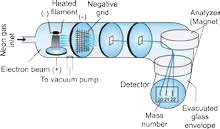
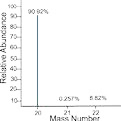


Fig (1.3) Computer plotted graph

Fig(1.2) Diagram of a simple Mass for the isotopes of neon Spectrometer

The above Fig (1.3) shows a computer plotted graph for the isotopes of neon.

The separation of isotopes can be done by the methods based on their properties. Some important methods are as gaseous difusion, thermal difusion, distillation, ultracentrifuge, electromagnetic separation and laser separation.

**1.3.3 Average Atomic Masses**

Table (1.1) of atomic masses of elements shows many examples of fractional values. Actually the atomic masses depend upon the number of possible isotopes and their natural abundance.

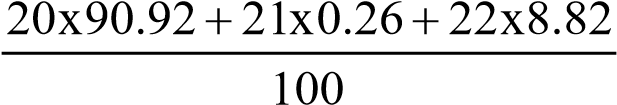
Following solved example will throw light on this aspect.

**Example (1):**

A sample of neon is found to consist of 2010 Ne , 1021 Ne and 2210 Ne in the percentages of 90.92%, 0.26%, 8.82% respectively. Calculate the fractional atomic mass of neon.

**Solution:**

The overall atomic mass of neon, which is an ordinary isotopic mixture, is the average of the determined atomic masses of individual isotopes. Hence

Average atomic mass = = 20.18*Answer*

Hence the average atomic mass of neon is 20.18 amu

It is important to realize that no individual neon atom in the sample has a mass of 20.18 amu. For most laboratory purposes, however, we consider the sample to consist of atoms with this average mass.

**1.4 ANALYSIS OF A COMPOUND - EMPIRICAL AND MOLECULAR**

## FORMULAS

Before we go into the details of empirical and molecular formulas of a compound, we should be interested to know the percentage of each element in the compound. For this purpose all the elements present in the compound are irst identiied.

This is called qualitative analysis. After that the compound is subjected to quantitative analysis in which the mass of each element in a sample of the compound is determined. From this we determine the percentage by mass of each element. The percentage of an element in a compound is the number of grams of that element present in 100 grams of the compound.

Mass of the element in the compound

Percentage of an element = x 100 mass of the compound

**Example (2):**

8.657 g of a compound were decomposed into its elements and gave 5.217 g of carbon, 0.962 g of hydrogen, 2.478 g of oxygen. Calculate the percentage composition of the compound under study.

**Solution:**

Applying the formula

Mass of carbon 5.217*g*

Percentage of carbon = x 100 = x 100 = 60.28 Answer

Mass of the compound 8.657*g*

Mass of hydrogen 0.962*g*

Percentage of hydrogen = x 100 = x 100 = 11.11 Answer

Mass of the compound 8.657*g*

Mass of oxygen 2.478*g*

Percentage of oxygen = x 100 = x 100 = 28.62 Answer

Mass of the compound 8.657*g*

The above results tell us that in one hundred grams of the given compound, there are 60.26 grams of carbon, 11.11 grams of hydrogen and 28.62 grams of oxygen.

Percentage composition of a compound can also be determined theoretically if we know the formula mass of the compound. The following equation can be used for this purpose.

Mass of the element in one mole of the compound

Percentage of an element = x 100 Formula mass of the compound

**1.4.1 Empirical Formula**

It is the simplest formula that gives the small whole number ratio between the atoms of diferent elements present in a compound. In an empirical formula of a compound, A**x**B**y**, there are x atoms of an element A and y atoms of an element B.

The empirical formula of glucose (C6H12O6) is CH2O and that of benzene (C6H6) is CH.

Empirical formula of a compound can be calculated following the steps mentioned below:

1. Determination of the percentage composition.
2. Finding the number of gram atoms of each element. For this purpose divide the mass of each element (% of an element) by its atomic mass.
3. Determination of the atomic ratio of each element. To get this, divide the number of moles of each element (gram atoms) by the smallest number of moles.
4. If the atomic ratio is simple whole number, it gives the empirical formula, otherwise multiply with a suitable digit to get the whole number atomic ratio.

**Example (3):**

Ascorbic acid (vitamin C) contains 40.92% carbon, 4.58% hydrogen and 54.5% of oxygen by mass. What is the empirical formula of the ascorbic acid?

**Solution:**

From the percentages of these elements, we believe that in 100 grams of ascorbic acid, there are 40.92 grams of carbon, 4.58 grams of hydrogen and 54.5 grams of oxygen.

Divide these masses of the elements (or percentages) by their atomic masses to get the number of gram atoms.

4.58g

No. of gram atoms of hydrogen = −1 = 4.54 gram atoms

1.008 gmol

54.5g

No. of gram atoms of oxygen = −1 = 3.406 gram atoms

16 gmol

40.92g

No. of gram atoms of carbon = −1 = 3.41 gram atoms

12.0 gmol

Atomic ratio is obtained by dividing the gram atoms with 3.406, which is the smallest number.

3.41 4.54 3.406

C:H:O = : :

3.406 3.406 3.406

C:H:O = 1 : 1.33 : 1

To convert them into whole numbers, multiply with three

C:H:O = 3(1 : 1.33 : 1) = 3 : 4 : 3 Answer

This whole number ratio gives us the subscripts for the empirical formula of the ascorbic acid i.e.,C3H4O3.

**1.4.2 Empirical Formula from Combustion Analysis**

Those organic compounds which simply consist of carbon, hydrogen and oxygen can be analyzed by combustion. The sole products will be CO2 and H2O. These two products of combustion are separately collected.

**Combustion Analysis**

A weighed sample of the organic compound is placed in the combustion tube. This combustion tube is itted in a furnance. Oxygen is supplied to burn the compound. Hydrogen is converted to H2O and carbon is converted to CO2. These gases are absorbed in Mg (ClO4)2 and 50% KOH respectively. (Fig 1.4). The diference in the masses of these absorbers gives us the amounts of H2O and CO2 produced. The amount of oxygen is determined by the method of diference.



Fig(1.4) Combustion analysis

Following formulas are used to get the percentages of carbon, hydrogen and oxygen, respectively.

Mass of CO2 12.00

% of carbon = x x 100

Mass of organic compound 44.00

Mass of H O2 2.016

% of hydrogen = x x 100 Mass of organic compound 18

The percentage of oxygen is obtained by the method of diference.

% of oxygen = 100 - (% of carbon + % of hydrogen) .

**Example (4):**

A sample of liquid consisting of carbon, , hydrogen and oxygen was subjected to combustion analysis. 0.5439 g of the compound gave 1.039 g of CO2, 0.6369. g of H2O.

Determine the empirical formula of the compound.

**Solution:**

|  |  |  |  |
| --- | --- | --- | --- |
| Mass of organic Compound | | = 0.5439 g | |
| Mass of carbon dioxide | | = 1.039g | |
| Mass of water | | = 0.6369 g | |
| **Element** | **%** | | **No. of Gram atoms** | | **Atomic Ratio** | **Empirical formula** |
| **C**  **H**  **O** | 1.039g 12.00  x x100  0.543*g* 44.00  =52.108  0.6369g 2.016  x x100  0.5439*g* 18  =13.11  100 (52.108 13.11)− +  =34.77 | | = 4.34  13.11  = 13.01  1.008  = 2.17 | | 4.34  = 2  2.17  = 6  2.17  =1  2.17 | C H O2 6 |

**1.4.3 Molecular Formula**

That formula of a substance which is based on the actual molecule is called molecular formula. It gives the total number of atoms of diferent elements present in the molecule of a compound. For example, molecular formula of benzene is C6H6 while that of glucose is C6H12O6.

The empirical formulas of benzene and glucose are CH and CH2O respectively, so for these compounds the molecular formulas are the simple multiple of empirical formulas. Hence Molecular formula = n (Empirical formula)

Where ‘n’ is a simple integer. Those compounds whose empirical and molecular formulae are the same are numerous. For example, H2O, CO2, NH3 and C12H22O11 have same empirical and molecular formulas. Their simple multiple ‘n’ is unity. Actually the value of ‘n’ is the ratio of molecular mass and empirical formula mass of a substance.

Molecular mass

n =

Empirical formula mass

**Example (5):**

The combustion analysis of an organic compound shows it to contain 65.44% carbon, 5.50% hydrogen and 29.06% or oxygen. What is the empirical formula of the compound? If the molecular mass of this compound is 110.15 g.mol-1. Calculate the molecular formula of the compound.

**Solution:**

First of all divide the percentage of each element by its atomic mass to get the number of gram atoms or moles.

|  |  |  |
| --- | --- | --- |
|  |  | 65.44 g of C  No of gram atoms of C = = 5.45 gram atoms of C  12 g / mol  5.50 g of H  No of gram atoms of hydrogen = = 5.46 gram atoms of H  1.008 g / mol  29.06 g of O  No of gram atoms of oxygen = = 1.82 gram atoms of O  16.00 g / mol |
| Molar ratio: |  | C : H : O |
| 5.45 : 4.46 : 1.82 | |
|  |  | |

Divide the number of gram atoms by the smallest number i.e 1.82

C : H : O

5.45 5.46 1.82

: :

1.82 1.82 1.82

3 : 3 : 1

Carbon, hydrogen and oxygen are present in the given organic compound in the ratio of 3:3:1. So the empirical formula is C3H3O.

In order to determine the molecular formula, irst calculate the empirical formula mass.

Empirical formula mass **= 12 x 3 + 1.008 x 3 + 16 x 1 = 55.05 g/mol**

Molar mass of the compound **= 110.15g.mol-1**

Molar mass of the compound 110.15

n = = = 2 Empirical formula mass 55.05

Molecular formula **= n (empirical formula)**

**= 2 (C3H3O) =C6H6O2 Answer**

There are many possible structures for this molecular formula.

## 1.5 CONCEPT OF MOLE

We know that atom is an extremely small particle. The mass of an individual atom is extremely small quantity. It is not possible to weigh individual atoms or even small number of atoms directly. That is why, we use the atomic mass unit (amu) to express the atomic masses.

For the sake of convenience, the atomic mass may be given in any unit of measurement i.e. grams, kg, pounds, and so on.

When the substance at our disposal is an element then the atomic mass of that element expressed in grams is called one gram atom. It is also called one gram mole or simply a mole of that element.

Mass of an element in grams

Number of gram atoms or moles of an element =

Molar mass of an element

**For example**

1 gram atom of hydrogen = 1.008 g

1 gram atom of carbon = 12.000 g

and 1 gram atom of uranium = 238.0 g

It means that one gram atom of diferent elements have diferent masses in them. One mole of carbon is 12 g, while 1 mole of magnesium is 24g. It also shows that one atom of magnesium is twice as heavy as an atom of carbon.

The molecular mass of a substance expressed in grams is called gram molecule or gram mole or simply the mole of a substance.

Mass of molecular substance in grams

Number of gram molecules or moles of a molecular substance = Molar mass of the substance

**For example**

1 gram molecule of water = 18.0 g

1 gram molecule of H2SO4 = 98.0 g and 1 gram molecule of sucrose = 342.0 g

It means that one gram molecules of diferent molecular substances have diferent masses.

**The formula unit mass of an ionic compound expressed in grams is called gram formula of the substance.** Since ionic compounds do not exist in molecular form therefore the sum of atomic masses of individual ions gives the formula mass. The gram formula is also referred to as gram mole or simply a mole.

Mass of the ionic substance in grams

Number of gram formulas or moles of a substance =

Formula mass of the ionic substance

1 gram formula of NaCl = 58.50 g

1 gram formula of Na2CO3 = 106 g

1gram formula of AgNO3 = 170g

It may also be mentioned here that ionic mass of an ionic species expressed in grams is called one gram ion or one mole of ions.

Mass of the ionic species in grams

Number of gram ions or moles of an species = Formula mass of the ionic species

**For example**

1 g ion of OH- = 17g

1 g ion of SO42-=96g

1 g ion of CO32- =60g

So, t**he atomic mass, molecular mass, formula mass or ionic mass of the substance expressed in gram is called molar mass of the substance.**

**Example (6):**

Calculate the gram atoms (moles) in (a) 0.1 g of sodium.

(b) 0.1 kg of silicon.

**Solution**

1. No. of gram atoms = Mass of element in gram

Molar mass

Mass of sodium = 0.1 g

Molar mass = 23 g/mol

0.1g

Number of gram atoms of sodium = −1 = 0.0043 mol 23 gmol

= 4.3 x 10 mol -3  Answer

1. First of all convert the mass of silicon into grams.

Mass of silicon = 0.1 kg = 0.1 x 1000 = 100 g

Molar mass = 28.086 gmol-1

Number of gram atoms of silicon = 100 g −1 = 3.56 moles Answer

28.086 gmol

**Example (7):**

Calculate the mass of 10-3 moles of MgSO4.

**Solution:**

MgSO4 is an ionic compound. We will consider its formula mass in place of molecular mass. Number of gram formula or mole of a substance = Mass of the ionic substance

Formula mass of the ionic substance

Formula mass of MgSO4 = 24 + 96 = 120 gmol-1

Number of moles of MgS04 = 10-3moles

Applying the formula

10-3 = Mass of MgSO−1 4

120 gmol

Mass of MgSO4 = 10-3 moles x 120gmol-1  = 120 x 10-3 = 0.12 g Answer

### 1.5.1 Avogadro's Number

Avogadro's number is the number of atoms, molecules and ions in one gram atom of an element, one gram molecule of a compound and one gram ion of a substance, respectively. To understand Avogadro's number let us consider the following quantities of substances.

1.008 g of hydrogen = 1 mole of hydrogen = 6.02 x 1023 atoms of H

23 g of sodium = 1 mole of Na = 6.02 x1O23 atoms of Na

238 g of uranium = 1 mole of U = 6.02 x 1023 atoms of U This number, 6.02 x 1023 is the number of atoms in one mole of the element. It is interesting to know that diferent masses of elements have the same number of atoms. An atom of sodium is 23 times heavier than an atom of hydrogen. In order to have equal number of atoms sodium should be taken 23 times greater in mass than hydrogen. Magnesium atom is twice heavier than carbon; i.e. 10g of Mg and 5 g of C contain the same number of atoms.

18 g of H2O =1 mole of water =6.02 x 1023 molecules of water 180 g of glucose = 1 mole of glucose = 6.02 x 1023 molecules of glucose

342 g of sucrose = 1 mole of sucrose = 6.02 x 1023 molecules of sucrose Hence, one mole of diferent compounds has diferent masses but has the same number of

molecules.

When we take into consideration the ions, then

96 g of SO42- = 1 mole of SO42- = 6.02 x 1023 ions of SO42-

62 g of NO3-  = 1 mole of NO3- = 6.02 x 1023 ions of NO3-

From the above discussion, we reach the conclusion that the number 6.02 X 1023 is equal to one mole of a substance. This number is called Avogadro’s number and it is denoted by NA.

Following relationships between amounts of substances in terms of their masses and the number of particles present in them, are useful

Mass of the element x NA

1. Number of atoms of an element =

Atomic mass

Mass of the compound x NA

1. Number of molecules of a compound =

Molecular mass

Mass of the ion x NA

1. Number of ions of an ionic species = Ionic mass

When we have compounds of known mass we can calculate the number of atoms from their formulas. In 18 g of water there are present 6.02 x 1023 molecules of H2O, 2 x 6.02 x 1023 atoms of hydrogen and 6.02 x 1023 atoms of oxygen. Similarly, in 98g of H2SO4, it has twice the Avogadro’s number of hydrogen atoms, four times the Avogadro’s number of oxygen atoms and the Avogadro’s number of sulphur atoms.

Some substances ionize in suitable solvents to yield cations and anions. The number of such ions, their masses, number of positive and negative charges can be easily calculated from the known amount of the substance dissolved. Let us dissolve 9.8 g of H2SO4 in suicient quantity of H2O to get it completely ionized. It has 0.1 moles of H2SO4. It will yield 0.2 mole or 0.2 x 6.02 x 1023 H+ and 0.1 moles or 0.1 x 6.02 x 1023 SO42- etc. Total positive charges will be 0.2 x 6.02 x 1023 and the total negative charges will be 0.2 x 6.02 x 1023 (because each SO42-, has two negative charges). The total mass of H+ is (0.2 x 1.008)g and that of SO42- is (0.1 x 96) g.

**Example (8):**

How many molecules of water are there in 10.0 g of ice? Also calculate the number of atoms of hydrogen and oxygen separately, the total number of atoms and the covalent bonds present in the sample.

**Solution:**

Mass of ice (water) = 10.0 g Molar mass of water =18gmol-1

Mass of water in gram

Number of molecules of water = −1 x Avogadro's number Molar mass of water in g mol

10 23

= -1 x 6.02 x 10

18 g mol

|  |
| --- |
| 3.31 *x* 1023 |

Number of molecules of water = 0.55 x 6.02 x 1023 = Answer

|  |  |  |
| --- | --- | --- |
| One molecule of water contain hydrogen atoms |  | = 2 |
| 3.31 x 1023 molecules of water contain hydrogen atoms |  | = 2 x 3.31 x 1023 |

|  |
| --- |
| 6.68 *x* 1023 |

= Answer

One molecule of water contains oxygen atom = 1

|  |
| --- |
| 3.31 *x* 1023 |

3.31 x 1023 molecules of water contain oxygen atoms = Answer

One molecule of water contains number of covalent bonds =2

3.31 x 1023 molecules of water contain number of covalent bonds = 2 x 3.31 x 1023

|  |
| --- |
| 6.68 *x* 1023 |

= Answer

|  |
| --- |
| 9.99 *x* 1023 |

Total number of atoms of hydrogen and oxygen = 6.68 x 1023 + 3.31 x 1023  = Answer

**Example (9):**

10.0 g of H3PO4 has been dissolved in excess of water to dissociate it completely into ions.

Calculate,

1. Number of molecules in 10.0 g of H3PO4.
2. Number of positive and negative ions in case of complete dissociation in water.c) Masses of individual ions.

d) Number of positive and negative charges dispersed in the solution.

**Solution:**

|  |  |  |  |
| --- | --- | --- | --- |
| (a) | Mass of H3PO4 |  | =10 g |
|  | Molar mass of H3PO4 |  | =3 + 31 + 64 = 98 |
|  | No. of molecules of H3P04 |  | = Mass of H PO  = 3 4  x 6.02 x 1023  Molar mass of H PO3 4 |

10 23  = −1 x 6.02 x 10

98 g *mol*

= 0.102 x 6.02 x 1023

= 0.614 x 1023

|  |
| --- |
| 6.14 x 1022 |

= Answer

(b) H3PO4 dissolves in water and ionizes as follows

H3PO4 ฀ 3H+ +PO43According to the balanced chemical equation

H3PO4 : H+

1 : 3

6.14 x 1022 : 3 x 6.14 x 1022 6.14 x 1022 : 1.842 x 1023

Hence, the number of H+ will be 1.842 x 1023

H3PO4 : PO43-

1 : 1

6.14 x 1022 : 6.14 x 1022

|  |
| --- |
| 6.14 x 1022 |

Hence, the number of PO43- will be Answer

(c) In order to calculate the mass of the ions, use the formulas

+ Total mass of H+ 23

Number of H = + x 6.02 x 10

Ionic mass of H

23 Total mass of H+ 23

1.842 x 10 = x 6.02 x 10

1.008

Total mass of H = + 1.842 x 10 x 23 231.008 = 0.308 g

6.02 x 10

3- Total mass of PO43- 23

No. of PO4 = 3- x 6.02 x 10 molecules

Ionic mass of PO4

22 Total mass of PO43- 23

6.14 x 10 = x 6.02 x 10

95

3- 6.14 x 1022 x 95

Total mass of PO4 = 23 = 9.689g Answer 6.02 x 10

(d) One molecule of H3PO4 gives three positive charges in the solution

6.14 x 1022 molecules of H3PO4 will give =3 x 6.14 x 1022 = 1.842 x 10 positive charges23 Answer

Number of positive and negative charges are always equal. So the number of negative charges dispersed in the solution = 1.842 x 1023

**1.5.2 Molar Volume**

One mole of any gas at standard temperature and pressure (STP) occupies a volume of 22.414 dm3. This volume of 22.414 dm3 is called molar volume and it is true only when the gas is ideal (the idea of the ideality of the gas is mentioned in chapter three).

With the help of this information, we can convert the mass of a gas at STP into its volume and vice versa.

Hence we can say that

2.016 g of H2 = 1 mole of H2 = 6.02 x 1023 molecules of H2 = 22.414 dm3 of H2 at S.T.P 16g of CH4 = 1 mole of CH4 = 6.02 x 1023 molecules of CH4= 22.414 dm3 of CH4 at S.T.P.

It is very interesting to know from the above data that 22.414 dm3 of each gas has a diferent mass but the same number of molecules. The reason is that the masses and the sizes of the molecules don't afect the volumes. Normally, it is known that in the gaseous state the distance between molecules is 300 times greater than their diameters.

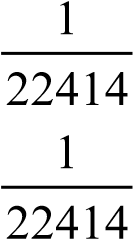
**Example (10):**

A well known ideal gas is enclosed in a container having volume 500 cm3 at S.T.P. Its mass comes out to be 0.72g.What is the molar mass of this gas.

**Solution:**

We can calculate the number of moles of the ideal gas at S.T.P from the given volume.

22.414 dm or 22.414 cm of the ideal gas at S.T.P = 1 mole 3 3

 1 cm of the ideal gas at S.T.P3  = mole

500 cm of the ideal gas at S.T.P3  = x 500

= 0.0223 moles

We know that

Mass of the gas

Number of moles of the gas =

Molar of the gas

Mass of the gas

Molar mass of the gas =

Number of moles of the gas

|  |
| --- |
| 32 g mol-1 |

0.72 g

Molar mass of the gas = = Answer

0.0223 mole

## 1.6 STOICHIOMETRY

With the knowledge of atomic mass, molecular mass, the mole, the Avogadro’s number and the molar volume, we can make use of the chemical equations in a much better way and can get many useful information from them.

Chemical equations have certain limitations as well. They do not tell about the conditions and the rate of reaction. Chemical equation can even be written to describe a chemical change that does not occur. So, when stoichiometeric calculations are performed, we have to assume the following conditions.

1. All the reactants are completely converted into the products.
2. No side reaction occurs.

**Stoichiometry is a branch of chemistry which tells us the quantitative relationship between reactants and products in a balanced chemical equation.**

While doing calculations, the law of conservation of mass and the law of deinite proportions are obeyed.

The following type of relationships can be studied with the help of a balanced chemical equation.

1. **Mass-mass Relationship**

If we are given the mass of one substance, we can calculate the mass of the other substances involved in the chemical reaction.

1. **Mass-mole Relationship or Mole-mass Relationship**

If we are given the mass of one substance, we can calculate the moles of other substance and viceversa.

1. **Mass-volume Relationship**

If we are given the mass of one substance, we can calculate the volume of the other substances and vice-versa.Similarly, mole-mole calculations can also be performed.

**Example (11):**

Calculate the number of grams of K2SO4 and water produced when 14 g of KOH are reacted with excess of H2SO4. Also calculate the number of molecules of water produced.

**Solution:**

For doing such calculations, irst of all convert the given mass of KOH into moles and then compare these moles with those of K2SO4 with the help of the balanced chemical equation.

Mass of KOH = 14.0 g

Molar mass of KOH = 39 + 16 + 1 = 56 g/mol

14.0 g

Number of moles of KOH = -1 = 0.25 56 g mol

**Equation:** 2KOH(aq) + H2SO4(aq) K2SO4(aq) + 2H2O(l)

To get the number of moles of K2SO4, compare the moles of KOH with those of K2SO4.

KOH : K SO2 4

2 : 1



1 :

0.25 : 0.125

So, 0.125 moles of K2SO4 is being produced from 0.25 moles of KOH

Molar mass of K2SO4 = 2 x 39 + 96

= 174 g/mol

Mass of K2SO4 produced = No. of moles x molar mass

= 0.125 moles x 174 g mol-1 =21.75g

To get the number of moles of H2O, compare the moles of KOH with those of water

*KOH* : H O2

2 : 2

1 : 1

0.25 : 0.25

So, the number of moles of water produced is 0.25 from 0.25 moles of KOH

Mass of water produced = 0.25 moles x 18 g mol-1

= 4.50 g

Number of molecules of water = No. of moles x 6.02 x 1023

= 0.25 moles x 6.02 x 1023 molecules per mole

= 1.50 x 1023 molecules Answer

**Example (12):**

Mg metal reacts with HCl to give hydrogen gas. What is the minimum volume of HCl solution (27% by weight) required to produce 12.1g of H2. The density of HCl solution is 1.14g/cm3.

Mg (s) + 2HCl (aq)  MgCl2(aq) + H2(g) **Solution:**

Mass of H2 produced = 12.1 g

Molar mass of H2 = 2.016 g mol-1

Moles of H = 2 Molar mass of HMass of H 2 2 = 2.01612.1 g molg -1=6.0 moles

To calculate the number of moles of HCl, compare the moles of H2 with those of HCl

*H* 2 : HCl

1 : 2

6 : 12

So, 12 moles of HCl are being consumed to produce 6 moles of H2.

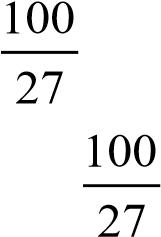
Mass of HCl =Moles of HCl x Molar mass of HCl

= 12 moles x 36.5 g mol-1  = 438 grams

We know that HCl solution is 27% by weight, it means that

27 g of HCl are present in HCl solution = 100 g

1 g is present in HCl solution =

 438 g are present in HCl solution = x 438 = 1622.2 g

Density of HCl solution = 1.14 g/cm3

Mass of HCl solution

Volume of HCl =

Density of HCl

1622.2*g* 3

= −3 = 1423 cm Answer

1.14*gcm*

## 1.7 LIMITING REACTANT

Having completely understood the theory of stoichiometry of the chemical reactions, we shift towards the real stoichiometric calculations. Real in the sense that we deal with such calculations very commonly in chemistry. Often, in experimental work, one or more reactants is/are deliberately used in excess quantity. The quantity exceeds the amount required by the reaction’s stoichiometry. This is done, to ensure that all of the other expensive reactant is completely used up in the chemical reaction. Sometimes, this strategy is employed to make reactions occur faster. For example, we know that a large quantity of oxygen in a chemical reaction makes things burn more rapidly. In this way excess of oxygen is left behind at the end of reaction and the other reactant is consumed earlier. This reactant which is consumed earlier is called a limiting reactant. In this way, the amount of product that forms is limited by the reactant that is completely used. Once this reactant is consumed, the reaction stops and no additional product is formed. **Hence the limiting reactant is a reactant that controls the amount of the product formed in a chemical reaction due to its smaller amount.**

The concept of limiting reactant is analogous to the relationship between the number of “kababs” and the “slices” to prepare “sandwiches”. If we have 30 “kababs” and ive breads “having 58 slices”, then we can only prepare 29 “sandwiches”. One “kabab” will be extra (excess reactant) and “slices” will be the limiting reactant. It is a practical problem that we can not purchase exactly sixty “slices” for 30 “kababs” to prepare 30 “sandwiches”.

Consider the reaction between hydrogen and oxygen to form water.

2H2(g) + O2(g) 2H2O(l)

When we take 2 moles of hydrogen (4g) and allow it to react with 2 moles of oxygen (64g), then we will get only 2 moles (36 g) of water. Actually, we will get 2 moles (36g) of water because 2 moles (4g) of hydrogen react with 1 mole (32 g) of oxygen according to the balanced equation. Since less hydrogen is present as compared to oxygen, so hydrogen is a limiting reactant. If we would have reacted 4 moles (8g) of hydrogen with 2 moles (64 g) of oxygen, we would have obtained 4 moles (72 g) of water.

### Identiication of Limiting Reactant

To identify a limiting reactant, the following three steps are performed.

1. Calculate the number of moles from the given amount of reactant.
2. Find out the number of moles of product with the help of a balanced chemical equation.
3. Identify the reactant which produces the least amount of product as limiting reactant.

Following numerical problem will make the idea clear.

**Example (13):**

NH3 gas can be prepared by heating together two solids NH4Cl and Ca(OH)2. If a mixture containing 100 g of each solid is heated then

1. Calculate the number of grams of NH3 produced.
2. Calculate the excess amount of reagent left unreacted.

2NH4Cl (s) + Ca (OH)2 (s) CaCl2(s) + 2NH3 (g) + 2H2O(l)

**Solution:**

1. Convert the given amounts of both reactants into their number of moles.

Mass of NH Cl = 100g4  Molar mass of NH C1 = 53.5g mol4 -1

100*g*

Mass of NH Cl = 4 -1 =1.87

53.5g mol

Mass of Ca(OH) = 1002 *g*

Molar mass of Ca(OH) = 74 g mol2 -1

100*g*

Moles of Ca(OH)2 = -1 =1.35

74 g mol

Compare the number of moles of NH4Cl with those of NH3

NH4Cl : NH3

2 : 2

1 : 1

1.87 : 1.87

Similarly compare the number of moles of Ca(OH)2 with those of NH3.

Ca(OH)2 : NH3

1 : 2

1.35 : 2.70

Since the number of moles of NH3 produced by l00g or 1.87 moles of NH4Cl are less, so NH4Cl is the limiting reactant. The other reactant, Ca(OH)2 is present in excess. Hence Mass of NH3 produced = 1.87 moles x 17 g mol-1

= 31.79 g Answer

1. Amount of the reagent present in excess

Let us calculate the number of moles of Ca(OH)2 which will completely react with 1.87 moles of NH4Cl with the help of equation. For this purpose, compare NH4Cl and Ca(OH)2

NH4Cl : Ca (OH)2

2 : 1

1 : 

1.87 : 0.935

Hence the number of moles of Ca(OH)2 which completely react with 1.87 moles of NH4Cl is 0.935 moles.

No. of moles of Ca(OH)2 taken =1.35

No. of moles of Ca(OH)2 used = 0.935

No. of moles of Ca(OH)2 left behind = 1.35 - 0.935 = 0.415

Mass of Ca(OH)2 left unreacted (excess) = 0.415x74 = 30.71 g  Answer

It means that we should have mixed 100 g of NH4Cl with 69.3 g (100 - 30.71) of Ca(OH)2 to get 1.87 moles of NH3.

## 1.8 YIELD

The amount of the products obtained in a chemical reaction is called the actual yield of that reaction. The amount of the products calculated from the balanced chemical equation represents the theoretical yield. The theoretical yield is the maximum amount of the product that can be produced by a given amount of a reactant, according to balanced chemical equation.

In most chemical reactions the amount of the product obtained is less than the theoretical yield. There are various reasons for that. A practically inexperienced worker has many shortcomings and cannot get the expected yield. The processes like iltration, separation by distillation, separation by a separating funnel, washing, drying and crystallization if not properly carried out, decrease the actual yield. Some of the reactants might take part in a competing side reaction and reduce the amount of the desired product. So in most of the reactions the actual yield is less than the theoretical yield.

A chemist is usually interested in the eiciency of a reaction. The eiciency of a reaction is expressed by comparing the actual and theoretical yields in the form of percentage (%) yield.

Actual yield

% yield = x 100

Theoretical yield

**Example (14):**

When lime stone (CaCO3) is roasted, quicklime (CaO) is produced according to the following equation. The actual yield of CaO is 2.5 kg, when 4.5 kg of lime stone is roasted. What is the percentage yield of this reaction.

CaCO (s) 3 → CaO(s) + CO (s)2

**Solution:**

Mass of limestone roasted = 4.5 kg = 4500 g

Mass of quick lime produced (actual yield) = 2.5 kg = 2500 g

Molar mass of CaCO3 = 100 g mol-1

Molar mass of CaO = 56 g mol-1

According to the balanced chemical equation

100 g of CaCO3 should give CaO = 56 g

1g of CaCO3 should give CaO = 56 / 100

4500 g of CaCO3 should give CaO = 56 / 100 x 4500

= 2520 g

Theoretical yield of CaO = 2520 g

Actual yield of CaO = 2500 g

Actual yield 2500*g*

% yield = x 100 = x 100

Theoretical yield 2520*g*

= 99.2 % Answer

## KEY POINTS

1. Atoms are the building blocks of matter. Atoms can combine to form molecules. Covalent compounds mostly exist in the form of molecules. Atoms and molecules can either gain or lose electrons, forming charged particles called ions. Metals tend to lose electrons, becoming positively charged ions. Non-metals tend to gain electrons forming negatively charged ions. When X-rays or α -particles are passed through molecules in a gaseous state, they are converted into molecular ions.
2. The atomic mass of an element is determined with reference to the mass of carbon as a standard element and is expressed in amu. The fractional atomic masses can be calculated from the relative abundance of isotopes. The separation and identiication of isotopes can be carried out by mass spectrograph.
3. The composition of a substance is given by its chemical formula. A molecular substance can be represented by its empirical or a molecular formula. The empirical and molecular formula are related through a simple integer.
4. Combustion analysis is one of the techniques to determine the empirical formula and then the molecular formula of a substance by knowing its molar mass.
5. A mole of any substance is the Avogadro’s number of atoms or molecules or formula units of that substance.
6. The study of quantitative relationship between the reactants and the products in a balanced chemical equation is known as stoichiometry. The mole concept can be used to calculate the relative quantities of reactants and products in a balanced chemical equation.
7. The concept of molar volume of gases helps to relate solids and liquids with gases in a quantitative manner.
8. A limiting reactant is completely consumed in a reaction and controls the quantity of products formed.
9. The theoretical yield of a reaction is the quantity of the products calculated with the help of a balanced chemical equation. The actual yield of a reaction is always less than the theoretical yield. The eiciency of a chemical reaction can be checked by calculating its percentage yield.