1. Details of Module and its structure

Module Detail	
Subject Name	Chemistry
Course Name	Chemistry 01 (Class XI, Semester 01)
Module Name/Title	Some Basic Concepts of Chemistry: Part 3
Module Id	kech_10103
Pre-requisites	Atom, Molecule, Matter, Different types of properties of Matter
Objectives	After going through this module you will be able to: 1. Explain the laws of chemical combination 2. Explain Dalton's atomic theory 3. Explain the term Avogadro number 4. Define atomic mass unit(u) 5. Differentiate between atomic mass, average atomic mass, molecular mass, formula mass
Keywords	Law of conservation of mass, law of definite proportion, law of multiple proportion, Gay Lussac's law, Dalton's atomic theory, Avogadro number, atomic mass unit, average atomic mass, molecular mass, formula mass

2. Development Team

Role	Name	Affiliation
National MOOC Coordinator (NMC)	Prof. Amarendra P. Behera	CIET, NCERT, New Delhi
Program Coordinator	Dr. Mohd. Mamur Ali	CIET, NCERT, New Delhi
Course Coordinator (CC) / PI	Prof. R. K. Parashar	DESM, NCERT, New Delhi
Course Co-Coordinator / Co-PI	Dr. Aerum Khan	CIET, NCERT, New Delhi
Subject Matter Expert (SME)	Dr. Komal S. Khatri	G. B. Pant Institute of Polytechnic, Okhla-II, New Delhi
Review Team	Dr. Alka Mehrotra Dr. Aerum Khan	DESM, NCERT, New Delhi CIET, NCERT, New Delhi

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1. Introduction

In Module-1 you have learnt that chemistry deals with the composition, structure and properties of matter. These aspects can be best described and understood in terms of basic constituents of matter - atoms and molecules. It is interesting to note that physics and biology made fair progress without knowledge of atoms but progress in chemistry began after formulation of atomic hypothesis. Laws of chemical combination provided us the first scientific evidence for existence of atoms; on the basis of which John Dalton proposed the 'Atomic Theory of Matter'. You will learn about this theory later in this module.

You know from your earlier classes that scientists recognised the difference between elements and compounds and became interested in finding out how and why elements combine and what happens when they combine. Antoine L. Lavoisier laid the foundation of chemical sciences by establishing two important laws of chemical combination.

2. Laws of Chemical Combinations

Now it is well known that chemical reactions involve the process of either combination of elements to form compounds or dissociation of compounds into its elementary particles. All these processes are governed by the following five basic laws:

1. Law of Conservation of Mass

2. Law of Definite Proportions

3. Law of Multiple Proportions

4. Avogadro Law

5. Law of Combining Volumes or Gay Lussac's Law of Gaseous Volumes

Out of these five laws, first four laws are related to masses while the fifth law i.e. Gay Lussac's law of gaseous volume deals with the volumes of the reacting gases.

2.1 Law of Conservation of Mass

It was put forth by a French chemist, Antoine Lavoisier (Fig. 1) in 1789. He performed careful experimental studies for combustion reactions and reached to the conclusion that in all physical and chemical changes or reactions, the total mass of the reactants is equal to that of the products. Hence, according to this law there is no net change in mass during the process. The law of conservation of mass is stated as ―Matter can neither be created nor destroyed‖.

Fig. 1 Antoine Lavoisier (1743 - 1794)

(Source: https://upload.wikimedia.org/wikipedia/commons/7/78/Antoine_laurent_lavoisier.jpg)

There are many processes in our day-to-day life which illustrate this law. For example, change of ice cube in water. If we take an ice cube in a flask and properly corked it, upon heating this ice cube is converted into water. On weighting the flask containing water, it is found that there is no change in mass of water upon physical change. The law may be easily verified in the laboratory. A reaction between $CuSO_4$ (Copper Sulphate) solution and $BaCl_2$ (Barium Chloride) solution may be carried out and the contents of the reaction flask may be weighed before and after the reaction as explained in the following activity:

1. Take 10 ml dilute $CuSO_4$ solution in the conical flask.

2. Hang on ignition tube containing $BaCl_2$ solution in the flask carefully, taking care that the two solution do not mix.

3. Put a cork on the mouth of the flask to secure the content of the flask.

4. Weigh the reaction set up.

5. Tilt, swirl and mix the reactants in the flask and ignition tube.

6. Formation of a white precipitate of $BaSO_4$ indicates that a chemical reaction has taken place.

7. Weigh the reaction set up again .

This law formed the basis for several later developments in chemistry. In fact, this was the result of exact measurement of masses of reactants and products, and carefully planned experiments performed by Lavoisier.

2.2 Law of Definite Proportions

This law was given by, a French chemist, Joseph Proust (Fig. 2). The law states that –a given compound always contains exactly the same proportion of elements by weight‖. Proust worked with two samples of cupric carbonate — one of which was of natural origin and the other was synthetic.

Fig. 2 Joseph Proust (1754 - 1826)

(Source: http://dic.academic.ru/pictures/wiki/files/106/joseph-louis_proust.jpg)

He found that the composition of elements present in it was same for both the samples as shown below:

	% of Copper	% of Oxygen	% of Carbon
Natural Sample	51.35	9.74	38.91
Synthetic Sample	51.35	9.74	38.91

Thus, irrespective of its source, a given compound always contains same elements in the same proportion. The validity of this law has been confirmed by various experiments. It is sometimes also referred to as **Law of Definite Composition**.

Example-1:

Pure water obtained from whatever source like well, river, sea or lake of any country is always made up of hydrogen and oxygen atoms combined together in the same fixed ratio of 1:8 by mass.

Example-2:

Carbon dioxide contains carbon and oxygen combined together in a fixed ratio of 12:32 irrespective of process of its formation, i.e., heating of limestone ($CaCO_3$), burning of coal in air or by heating sodium hydrogen carbonate ($NaHCO_3$).

It is important to note that ratio of elements present in different samples of compounds may be different in case compounds in the samples contain different isotopes of elements. For example, in a molecule of CO_2 containing isotope C-12 two elements are combined together in a ratio of 12:32 by mass. If the molecule of CO_2 contains C-14 isotope, the ratio of carbon and oxygen is 14:32 by mass. Also, it is possible that same ratio of elements is present in different compounds. For example, ethanol (C_2H_5OH) and acetone (CH_3OCH_3) are two different compounds having same molecular formula of C_2H_6O with same ratio of (C:H:O) i.e., 24:6:16.

Problem: 6.448g of lead combine directly with 1.002g oxygen to form lead peroxide. Lead peroxide is also produced by heating lead nitrate and it was found that the percentage of oxygen present in lead peroxide is 13.38%. Use these data to illustrate the law of definite proportion.

Solution:

In first experiment, mass of lead peroxide formed = (6.488 + 1.002) = 7.490g

7.490g of lead peroxide contains 1.002g of oxygen.

Therefore,100g of lead peroxide will contain (1.002/7.490)×100=13.38g oxygen or 13.38% oxygen. In the second experiment, it is given that percentage of oxygen in lead peroxide is 13.38%.

The, percentage of oxygen in both types of samples of lead peroxide is same; hence this illustrates the law of definite proportions.

2.3 Law of Multiple Proportions

This law was given by by Dalton (Fig. 3) in 1803. According to this law, –if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers‖.

Fig. 3 John Dalton

(Source:
https://upload.wikimedia.org/wikipedia/commons/d/d4/John_Dalton_by_Charles_Turner.jpg)

For example, hydrogen combines with oxygen to form two compounds, namely, water and hydrogen peroxide.

Hydrogen + Oxygen → Water
2g 16g 18g
Hydrogen + Oxygen → Hydrogen Peroxide
2g 32g 34g

Here, the masses of oxygen (i.e. 16g and 32g) which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e. 16:32 or 1: 2.

Another example for the law of multiple proportions is compounds containing nitrogen and oxygen. The two elements nitrogen and oxygen combine together to form a variety of compounds shown in Table 1. In each of these compounds, a fixed mass of nitrogen i.e. 14 parts combines with multiple mass of oxygen to give variety of compounds. These masses follow a simple ratio of 1:2:3:4:5 to one another.

Table 1: Nitrogen and Oxygen combine together to form a variety of compounds

	Nitrogen Oxide	Name	Mass of N	Mass of O	Fixed mass of N, i.e. 14 parts combine with mass of oxygen	Simple ratio of Oxygen
1	N_2O	Nitrous Oxide	28	16	8	1
2	NO	Nitric Oxide	14	16	16	2
3	N_2O_3	Nitrogen Trioxide	28	48	24	3
4	N_2O_4	Nitrogen Tetraoxide	28	64	32	4
5	N_2O_5	Nitrogen Pentaoxide	28	80	40	5

Problem: Percentage of carbon present in two oxides of carbon is given below, show that, these oxides of carbon illustrate the law of multiple proportions.

Percentage of Carbon in first oxide	Percentage of Carbon in second oxide
42.9	27.3

Solution:

Percentage composition of the carbon and oxygen in two oxides can be calculated as given in the following Table 2.

Table 2: Percentage composition of carbon and oxygen in two oxides

Element present in the compound	First Oxide	Second Oxide
Carbon (given)	42.9 %	27.3%
Oxygen (by difference)	57.1%	72.7%

Thus in 100 g of first oxide 42.9g of carbon and 57.1g oxygen have combined. Therefore 1g carbon has combined with 57.1/42.9g oxygen or 1.33g oxygen

In 100 g of second oxide 27.3g of carbon and 72.7g oxygen have combined. Therefore 1g of carbon has combined with 72.7/27.3g oxygen or 2.66g oxygen. The ratio of masses of oxygen combined with the1g (fixed mass) of carbon is 1.33:2.66 or 1:2.

This is simple whole number ratio. Hence, these oxides of carbon illustrate the law of multiple proportion.

2.4. Gay Lussac's Law of Gaseous Volumes

This law was given by Gay Lussac (Fig. 4) in 1808. He observed that when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume provided all gases are at same temperature and pressure.

Fig. 4 Joseph Louis Gay Lussac

(Source: https://upload.wikimedia.org/wikipedia/commons/thumb/2/2f/Gaylussac.jpg/170px-Gaylussac.jpg)

Thus, 100 ml of hydrogen combine with 50 mL of oxygen to give 100 mL of water vapour.

Hydrogen + Oxygen → Water

100 ml 50 ml 100 ml

Thus, the volumes of hydrogen and oxygen which combine together (i.e. 100 ml and 50 ml) bear a simple ratio of 2:1.

Gay-Lussac‘s discovery of integer ratio in volume relationship is actually the law of definite proportions by volume. The law of definite proportions, stated earlier, was with respect to mass. The Gay-Lussac‘s law was explained properly by the work of Avogadro in 1811.

Problem: How much volume of oxygen will be required for complete combustion of 40 ml of acetylene (C_2H_2).What volume of carbon dioxide will be formed? Chemical equation for the reaction between acetylene and oxygen is given below.

$2C_2H_2\,(g) + 5O_2\,(g) \rightarrow 4CO_2\,(g) + 2H_2O$

Solution: According to given chemical equation 2 volumes of acetylene combine with 5 volumes of Oxygen as shown below.

2 C_2H_2 (g) +	**5 O_2 (g)** →	**4CO_2 (g) + 2 H_2O**
2 volume	5volume	4 volume
1 Volume	(5/2) volume	4/2volumes of **CO_2** (g)
40 mL	(5/2) × 40 mL	(4/2) ×40 mL
40 mL	100 mL	80 mL

So for complete combustion of 40 mL of acetylene, 100 mL of oxygen will be required and 80 mL of CO_2 will be produced.

2.5. Avogadro Law

In 1811, Avogadro (Fig. 5) proposed that –equal volumes of all gases at the same temperature and pressure should contain equal number of molecules‖. Avogadro made a distinction between atoms and molecules which is quite understandable in the present times.

Fig. 5: Lorenzo Romano Amedeo Carlo Avagadro di Quarequa edi Carreto

(Source: https://whewellsghost.files.wordpress.com/2015/08/avogadro_amedeo-251x300.jpg)

If we consider again the reaction of hydrogen and oxygen to produce water, we see that two volumes of hydrogen combine with one volume of oxygen to give two volumes of water without leaving any unreacted oxygen.

In fact, Avogadro could explain the above result by considering the molecules to be polyatomic. If hydrogen and oxygen were considered as diatomic as recognised now, then the above results are easily understandable. Fig. 6 depicts this. Note that each box contains equal number of molecules.

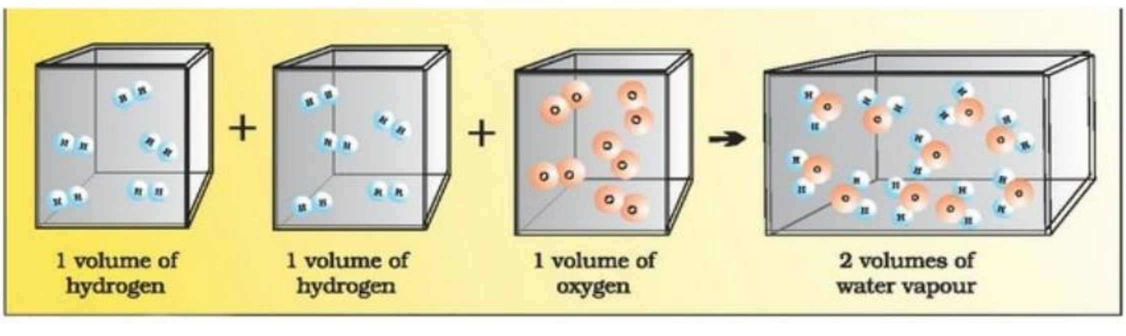

Fig. 6: Two volumes of hydrogen react with one volume of oxygen to give two volumes of water vapour

(Source: Chapter 1, page no. 12, XI Textbook, NCERT)

Avogadro‘s proposal was published in the French Journal de Physidue. In spite of being correct, it did not gain much support at that time. Dalton and others believed that atoms of the same kind cannot combine and molecules of oxygen or hydrogen containing two atoms did not exist. After about 50 years, in 1860, first international conference on chemistry was held in Karlsruhe, Germany to resolve various ideas. At the meeting, Stanislao Cannizaro presented a sketch of a course of chemical philosophy which emphasised the importance of Avogadro‘s work.

3. Dalton’s Atomic Theory

Although the origin of idea that matter is composed of small indivisible particles called ‗a-tomio‘ (meaning — indivisible), dates back to the time of Democritus, a Greek Philosopher (460 — 370 BC), it again started emerging after the several experimental studies which led to the Laws mentioned above. John Dalton an English school teacher in Manchester, was the first scientist who suggested that the laws of chemical combination point to the existence of atoms. In 1808, Dalton published ‗A New System of Chemical Philosophy‘. In this theory, he proposed the following:

1. Matter is made up of very small indivisible particles known as atoms.

2. All the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.

3. Compounds are formed when atoms of different elements combine in a fixed ratio.

4. Chemical reactions involve reorganisation of atoms. Atoms are neither created nor destroyed in a chemical reaction.

Dalton‘s theory also could explain the laws of chemical combination. The theory explains that there are as many distinct type of atoms as the number of element. However, Dalton‘s theory could not explain the law of gaseous volumes. It could not provide the reason for the combining of atoms. It raised many other questions .such as what is the shape of an atom? what is inside the atom? How are the atoms arranged in a molecule, which were answered later by other scientists.

4. Atomic and Molecular Masses

After having some idea about the terms atoms and molecules, it is appropriate here to understand what we mean by atomic and molecular masses.

Atomic Mass: The atomic mass or the mass of an atom is actually very small because atoms are extremely small. Today, we have sophisticated techniques e.g., mass spectrometry for determining the atomic masses fairly accurately. But, in the nineteenth century, scientists could determine mass of one atom relative to another by experimental means. Hydrogen, being lightest atom was arbitrarily assigned a mass of 1 (without any units) and other elements were assigned masses relative to it.

However, as agreed upon in 1961,the present system of atomic masses considers carbon - 12 as the standard. Here, Carbon - 12 is one of the isotopes of carbon and can be represented as 12 C. In this system, 12 C is assigned a mass of exactly 12 atomic mass unit (amu) and masses of all other atoms are given relative to this standard. One atomic mass unit (amu) is defined as a mass exactly equal to one twelfth the mass of one carbon - 12 atom.

And, 1 amu = 1.66056×10^{-24} g

Mass of an atom of hydrogen = 1.6736×10^{-24} g

Thus, in terms of amu, the mass of hydrogen atom = $(1.6736 \times 10^{-24}\ g)/(1.66056 \times 10^{-24}\ g)$

= 1.0078 amu

= 1.0080 amu

Similarly, the mass of oxygen - 16 (16 O) atom would be 15.995 amu. Today, _amu‘ has been

replaced by _u‘ which is known as unified mass.

When we use atomic masses of elements in calculations, we actually use average atomic masses of elements which are explained below.

Average Atomic Mass: Many naturally occurring elements exist as more than one isotope. When we take into account the existence of these isotopes and their relative abundance in nature (per cent occurrence), the average atomic mass of that element can be computed.

For example, carbon has the following three isotopes with relative abundances and masses as in (Table 3).

Table 3: Isotopes of Carbon relative abundance and masses

Isotopes	Relative Abundance (%)	Atomic Mass
^{12}C	98.892	12
^{13}C	1.108	13.00335
^{14}C	2×10^{-10}	14.00317

From the above data, the average atomic mass of carbon will come out to be:

(0.98892) (12 u) + (0.01108) (13.00335 u) + (2 × 10 -12) (14.00317 u) = 12.011 u

Similarly, average atomic masses for other elements can be calculated. In the periodic table of elements, the atomic masses mentioned for different elements actually represent their average atomic masses.

Problem: Chlorine has two isotopes with atomic masses 34.97 u and 36.97 u respectively. The relative abundance of the two isotopes are 0.755 and 0.245 respectively. Calculate the average mass of chlorine.

Solution: Average Atomic mass = (34.97 × 0.755) + (36.97 × 0.245) = 35.46

Molecular Mass: Molecular mass is the sum of atomic masses of the elements present in a molecule. It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together. For example, a molecule of methane contains one carbon atom and four hydrogen atoms. The molecular mass can be obtained as follows:

Molecular mass of methane, (CH_4) = (1× atomic mass of carbon atom) + (4 × atomic mass of hydrogen atom)

= (12.011 u) + 4(1.008 u)

= 16.043 u

Similarly, a water molecule contains 2 atoms of hydrogen and one atom of oxygen. Hence, Molecular mass of water = 2 (1.008 u) + 16.00 u = 18.02u

Problem: Calculate the molecular mass of a glucose molecule ($C_6H_{12}O_6$).

Solution: Molecular mass of glucose = 6 × (12.011 u) + 12 × (1.008 u) + 6 × (16.00 u)

= 180.162 u

Formula Mass: Some substances, especially ionic compounds such as sodium chloride (NaCl), potassium nitrate (KNO_3) etc. do not contain discrete molecules as their constituent units. In such compounds, positive (sodium or potassium etc.) and negative (chloride or nitrate etc.) entities are arranged in a three-dimensional structure. Arrangement of sodium and chlorides ions in sodium chloride (NaCl) is shown in Fig. 4.

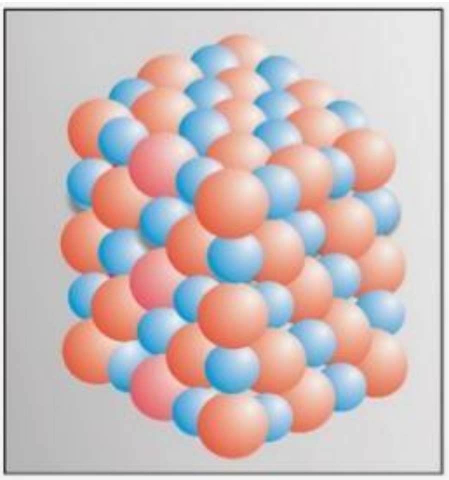

Fig. 4 Packing of Na^+ and Cl^- ions in sodium chloride.

(Source: Chapter 1, page no. 14, XI Textbook, NCERT)

It may be noted that in sodium chloride, one Na^+ ion is surrounded by six Cl^- ions and vice-versa. The formula such as NaCl is used to calculate the formula mass instead of molecular mass as in the solid state sodium chloride does not exist as a single entity. Thus, formula mass of sodium chloride

= atomic mass of sodium + atomic mass of chlorine

= 23.0 u + 35.5 u

= 58.5 u

5. Summary

In this module, we have studied that combination of different atoms is governed by basic laws of chemical combination. These laws are -- the Law of Conservation of Mass, the Law of Definite Proportions, the Law of Multiple Proportions, the Gay Lussac's Law of Gaseous Volumes and the Avogadro Law. All these laws led to the Dalton's atomic theory which states that atoms are building blocks of matter. The atomic mass of an element is expressed relative to the mass ^{12}C isotope of carbon which has an exact mass of 12 u. Usually, the value of atomic masses used for an element is the average atomic mass obtained by taking into account the natural abundance of different isotopes of that element. The molecular mass of a molecule is obtained by taking sum of the atomic masses of different atoms present in a molecule. For ionic compounds formula mass is obtained instead of molecular mass.