

Chemistry

SOME BASIC CONCEPTS OF CHEMISTRY

Significant Figures

Significant figures are the meaningful digits in a measured or calculated quantity. It includes all those digits that are known with certainty plus one more which is uncertain or estimated.

Greater the number of significant figures in a measurement, smaller the uncertainty.

Rules for Significant Figures

1. All non – zero digits are significant.
For example : In 185185 significant figure is 33.
2. Zero, proceeding to 1st non – zero digit are not significant. Such zero indicates the position of decimal point. For example : 0.03 – significant figure 1.
3. Zero between non – zero digits are significant.
For example : 2.005 – significant figure . 1001 – significant figure 4
4. All zeros placed to the right of a number are significant. For example, 16.0 has three significant figures, while 16.00 has four significant figures. Zeros at the end of a number without decimal point are ambiguous.

Rounding off Significant Figures

1. The rounding off procedure is applied to retain the required number of significant figures.
2. If the digit coming after the desired number of significant figures happens to be more than 5, the preceding significant figure is increased by one, 5.318 is rounded off to 5.32
3. If the digit involved is less than 5, it is neglected and the preceding significant figure remains unchanged, 4.312 is rounded off to 4.31.
4. If the digit happens to be 5, the last mentioned or preceding significant figure is increased by one only in case it happens to be odd. In case of even figure, the preceding digit remains unchanged. 8.375 is rounded off to 8.38 while 8.365 is rounded off to 8.36.

Calculations Involving Significant Figures

1. In addition or subtraction, the final result should be reported to the same number of decimal places as that of the term with the least number of decimal places, e.g.,

$$\begin{array}{r}
 2.512 \\
 2.2 \\
 5.23 \\
 \hline
 9.942 \Rightarrow 9.9
 \end{array}$$

(Reported sum should have only one decimal point.)

addition and subtraction significant figures examples

$$1. \quad 10.1 + 1.52 + 2.301 = 4.921 \quad 10.1 + 1.52 + 2.301 = 4.921$$

Answer 4.94.9

$$2. \quad 2.52 + 1.11 + 2.222 = 5.852 \quad 2.52 + 1.11 + 2.222 = 5.852$$

Answer 5.855.85

$$3. \quad 5.4 - 2.1 \quad 5.4 - 2.1 \\ = 3.3$$

2. In multiplication and division, the result is reported to the same number of significant figures as least precise term or the term with least number of significant figures, e.g.

multiplying and dividing significant figures examples

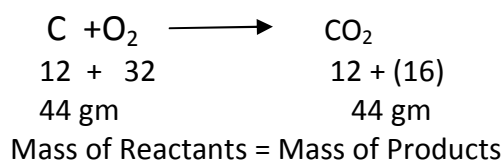
$$\begin{array}{ll}
 1. \quad 2.210 \times 0.011 & 2. \quad 1.01 \times 0.02 \\
 = 0.024312 & = 0.0202 \\
 = 0.024 & = 0.02
 \end{array}$$

Laws of Chemical Combinations

The combination of elements to form compounds is governed by the following six basic laws:

1. Law of conservation of mass (Lavoisier, 1774)

This law states that during any physical or chemical change, the total mass of the products is equal to the total mass of reactants. It does not hold good for nuclear reactions. Example

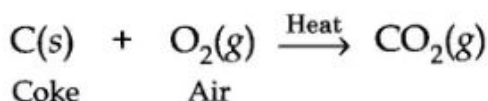


Mass of Reactants = Mass of Products

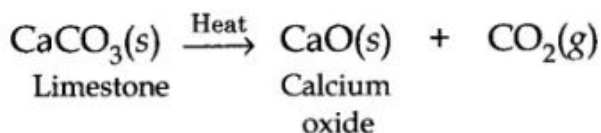
2. Law of definite proportions (Proust, 1799)

According to this law, a chemical compound obtained by different sources always contains same percentage of each constituent element. For eg. Carbondioxide can be formed by number of ways:

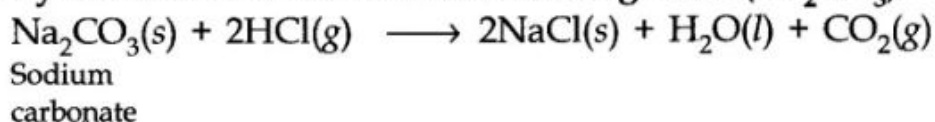
(a) **By burning coke in air**



(b) **By the decomposition of limestone (CaCO_3) on heating**



(c) **By the action of dil HCl on washing soda (Na_2CO_3)**

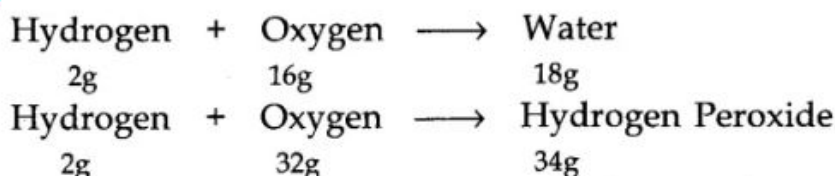


In all the three samples of CO_2 carbon and oxygen are in the ratio 3 : 8 by weight.

3. Law of multiple proportions (Dalton, 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, e.g., in NH_3 and N_2H_4 , fixed mass of nitrogen requires hydrogen in the ratio 3 : 2.

For example,



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

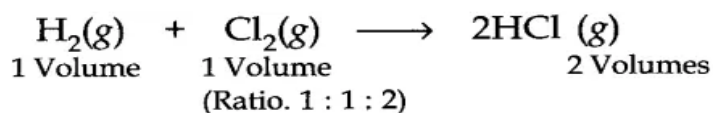
4. Law of reciprocal proportions (Richter, 1792)

According to this law, when two elements (say A and B) combine separately with the same weight of a third element (say C), the ratio in which they do so is the same or simple multiple of the ratio in which they (A and B) combine with each other.

Law of definite proportions, law of multiple proportions and law of reciprocal proportions do not hold good when same compound is obtained by using different isotopes of the same element. e.g H_2O and D_2O

5. Gay Lussac's law of gaseous volumes

It states that under similar conditions of temperature and pressure, whenever gases react together, the volumes of the reacting gases as well as products (if gases) bear a simple whole number ratio. For eg. Combination between hydrogen and chlorine-



6. Avogadro's hypothesis

It states that equal volumes of all gases under the same conditions of temperature and pressure contain the same number of molecules. for eg. If we consider 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.

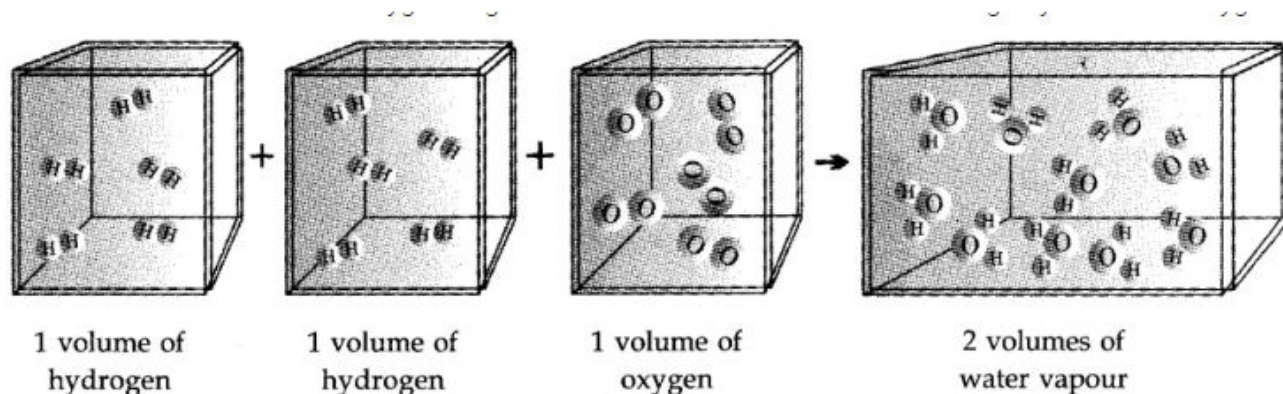


Fig. 1.4 Two volumes of hydrogen react with one volume of oxygen to give two volumes of water vapour.

Dalton's Atomic Theory (1803)

This theory was based on laws of chemical combinations. It's basic postulates are

1. All substances are made up of tiny, indivisible particles, called atoms.
2. In each element, the atoms are all alike and have the same mass. The atoms of different elements differ in mass.
3. Atoms can neither be created nor destroyed during any physical or chemical change.
4. Compounds or molecules result from combination of atoms in some simple numerical ratio.

ATOMIC MASS

The mass of one atom is called as atomic mass. Atomic mass of an element is defined as the average relative mass of an atom of an element as compared to the mass of an atom of carbon -12 taken as 12. It is also known as unified mass and denoted by u

$$1 \text{ amu} = 1.660056 \times 10^{-24} \text{ g}$$

Atomic Mass of the elements are expressed in terms of amu.

$$\text{Atomic Mass of element in amu} = \text{Mass of atom in grams} / 1.660056 \times 10^{-24}$$

$$\text{Atomic mass} = \frac{\text{mass of an atom}}{1/12 \text{ mass of a carbon atom } (^{12}\text{C})}$$

MOLECULAR MASS

Mass of molecule is the sum of masses of elements present in molecule

Example

$$\begin{aligned} \text{Molecular Mass of CO}_2 &= \text{Atomic mass of C} + 2 \times \text{atomic mass of Oxygen} \\ &= 12 + 2 \times 16 = 44 \text{ amu or } 44 \text{ u} \end{aligned}$$

$$\begin{aligned} \text{Molecular Mass of H}_2\text{O} &= \text{Atomic mass of O} + 2 \times \text{atomic mass of Hydrogen} \\ &= 16 + 2 \times 1 = 18 \text{ amu} \end{aligned}$$

Mole Concept

A mole is defined as the amount of substance which contains same number of elementary particles (atoms, molecules or ions) as the number of atoms present in 12 g of carbon (C-12).

$$1 \text{ mol} = 6.023 \times 10^{23} \text{ atoms} = \text{one gram-atom} = \text{gram atomic mass}$$

$$1 \text{ mol} = 6.023 \times 10^{23} \text{ molecules} = \text{gram molecular mass}$$

1. Calculate the no. of atoms present in 23 gm Sodium.

Solution

$$N_A \text{ atoms} = \text{G. A. M.}$$

$$6.022 \times 10^{23} \text{ atoms} = 23 \text{ gm}$$

$$23 \text{ gm} = 6.022 \times 10^{23} \text{ atoms}$$

$$1 \text{ gm} = (1/23) \times 6.022 \times 10^{23} \text{ atoms}$$

$$2.3 \text{ gm} = (2.3/23) \times 6.022 \times 10^{23} \text{ atoms}$$

$$= 6.022 \times 10^{22} \text{ atoms}$$

2. Calculate the no. of moles present in 6 gm Carbon.

Solution

For Carbon
G. A. M. = 1 mole
12 gm = 1 mole
1 gm = $1/12$ mole
For 6 gm = $6 \times (1/12)$ moles
= $1/2$ moles.

3. Give the no. of Molecules present in 0.18 gm of water.

Solution

Water Molecular mass = 18
so 18 gm = 1 moles
 $0.18 \text{ g} = .18 \times (1/18)$ moles
= 0.01 mole
= 6.022×10^{21} molecules

Limiting Reagents/Reactants

The reactants or reagent which totally get consumed during the reaction and limits the reaction is termed or Limiting Reagent.

Chemical Equation

The method of representing a chemical reaction with the help of symbols and formula of the substances involved in it is known as a chemical equation.

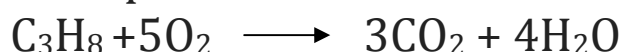
The substances which combine or react are known as reactants.

The new substances produced in a reaction are known as products.

A chemical equation is a short-hand method of representing a chemical reaction.

A balanced chemical equation has an equal number of atoms of different elements in the reactants and products.

Example of Balanced Chemical equation



C_3H_8 and O_2 are reactant and CO_2 and H_2O are product.

The coefficient 1 in C_3H_8 and 5 in O_2 are called stoichiometric coefficients.

Empirical Formula and Molecular Formula—

An **empirical formula** represents the simplest whole number ratio of various atoms present in a compound. E.g. CH is the empirical formula of benzene.

The **molecular formula** shows the exact number of different types of atoms present in a molecule of a compound. E.g. C₆H₆ is the molecular formula of benzene.

Relationship between empirical and molecular formulae

The two formulas are related as Molecular formula = n x empirical formula

$$n = \frac{\text{Molecular mass}}{\text{empirical formula mass}}$$