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 2
Module Id	kech_10102
Pre-requisites	Atom, Molecule, Matter, Different types of properties of Matter
Objectives	After going through this module you will be able to: ☐ Explain the need for common system of measurement ☐ Name base physical quantities and write their symbol and write name of SI units used for their measurement ☐ Use scientific notations and perform simple mathematical operations on numbers ☐ Differentiate between terms precision and accuracy ☐ Determine significant figures ☐ Convert units of measurement of one system to the units of measurement in another system
Keywords	International System of Units, SI units, Base Physical Quantity, Mass, Volume, Density, Temperature, Uncertainty, Scientific Notation, Accuracy, Precision, Dimensional Analysis.

2. DevelopmentTeam

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1. Properties of Matter and their Measurement

In your earlier classes you have learnt that every substance has unique characteristic properties. These properties can be classified as physical properties and chemical properties. Color, odour, melting point, boiling point, density, volume, etc., are examples of physical properties and combustibility, composition, reaction towards acids or bases etc. are examples of chemical properties. Chemists describe, interpret and predict the behaviour of substances on the basis of their physical and chemical properties, which are determined by careful measurements and experimentation. Therefore, quantitative measurement of properties is required in scientific investigations. Earlier two different systems of measurement were being used in different parts of the world namely the English System and the Metric System. The metric system which originated in France in late eighteenth century was more convenient as it was based on the decimal system. Later on, a need of a common standard system was felt by the community of scientists. Such a system was established in 1960 and named as the International System of Units (SI).

1.1. The International System of Units (SI): The International System of Units (in French Le Systeme International d'Unités – abbreviated as SI) was established by the 11th General Conference on Weights and Measures (CGPM from *Conference Generale des Poids et Measures*). The CGPM is an inter governmental treaty organization created by a diplomatic treaty known as Metre Convention which was signed in Paris in 1875. The SI system has seven base units and they are listed in Table-1. These units pertain to the seven fundamental scientific quantities.

Table 1 Base physical quantities and their units

Base Physical Quantity	Symbol For Quantity	Name of the SI Unit	Symbol for SI Unit
Length	L	meter	m
Mass	M	kilogram	kg
Time	T	second	s
Electric Current	I	ampere	A
Thermodynamic Temperature	T	kelvin	K
Amount of substance	N	mole	mol
Luminous Intensity	I_v	candela	cd

Units for other physical quantities such as speed, volume, density etc. can be derived from these quantities.

The definitions of the SI base units are given in Table-2:

Table- 2: Definitions of SI Base Units

Unit of length	metre	The *metre* is the length of the path travelled by light in vacuum during a time interval of 1/299 792 458 of a second.
Unit of mass	kilogram	The *kilogram* is the unit of mass; it is equal to the mass of the international prototype of the kilogram.
Unit of time	second	The *second* is the duration of 9 192 631 770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the caesium-133 atom.
Unit of electric current	ampere	The *ampere* is that constant current which, if maintained in two straight parallel conductors of infinite length, of negligible circular cross-section, and placed 1 metre apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per metre of length.
Unit of thermodynamic temperature	kelvin	The *kelvin*, unit of thermodynamic temperature temperature, is the fraction 1/273.16 of the thermodynamic temperature of the triple point of water.
Unit of amount of substance	mole	1. The *mole* is the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12; its symbol is “mol.” 2. When the mole is used, the elementary entities must be specified and these may be atoms, molecules, ions, electrons, other particles, or specified groups of such particles.
Unit of luminous intensity	candela	The *candela* is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian.

1.2. Mass and Weight: Mass of a substance is the amount of matter present in it while weight is the force exerted by gravity on an object. The mass of a substance is constant whereas its weight may vary from one place to another due to change in gravity. The mass of a substance can be determined very accurately in the laboratory by using an analytical balance (Fig.1).

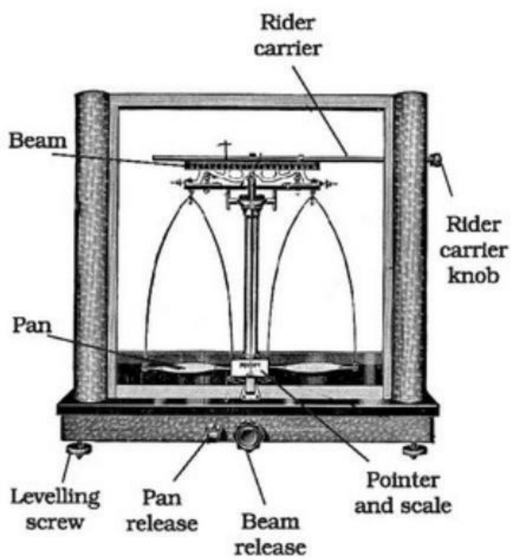

Fig. 1: Analytical balance

The SI unit of mass is kilogram. However, its fraction which is named as gram (1 kg = 1000 g), is used in laboratories because small amounts of chemicals are used in chemical reactions (Table 3).

Table 3: Prefixes used in the SI System

Multiple	Prefix	Symbol
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	A
10^{-15}	femto	F
10^{-12}	pico	P

10^{-9}	nano	N
10^{-6}	micro	μ
10^{-3}	milli	M
10^{-2}	centi	C
10^{-1}	deci	D
10	deca	Da
10^{2}	hecto	H
10^{3}	kilo	K
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta	P
10^{18}	exa	E
10^{21}	zeta	Z

1.3. Volume: Volume of a substance is the amount of space occupied by that substance. It has units of $(length)^3$ and in SI system, it is m^3. As in chemistry laboratories, generally smaller volumes are used. Hence, volume is often denoted in cm^3 or dm^3 units.

A common unit, litre (L) which is not an SI unit, is used for measurement of volume of liquids.

1 L = 1000 mL or $1 dm^3 = 1000 cm^3$

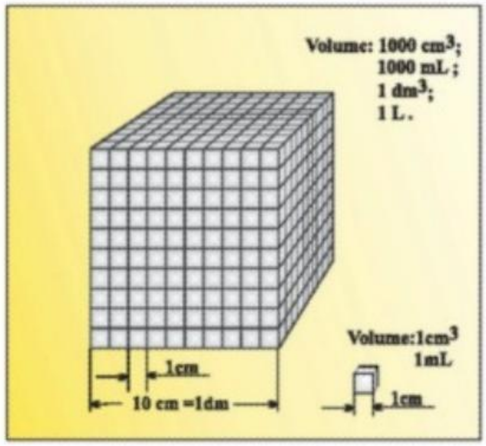

Fig 2. Different units used to express volume
(Source: Chapter 1, page no. 7, XI textbook, NCERT)

Fig. 2 helps to visualize these relations. In the laboratory, volume of liquids or solutions can be measured by graduated cylinder, burette, pipette etc. A volumetric flask is used to prepare a known volume of a solution. These measuring devices are shown in Fig.3.

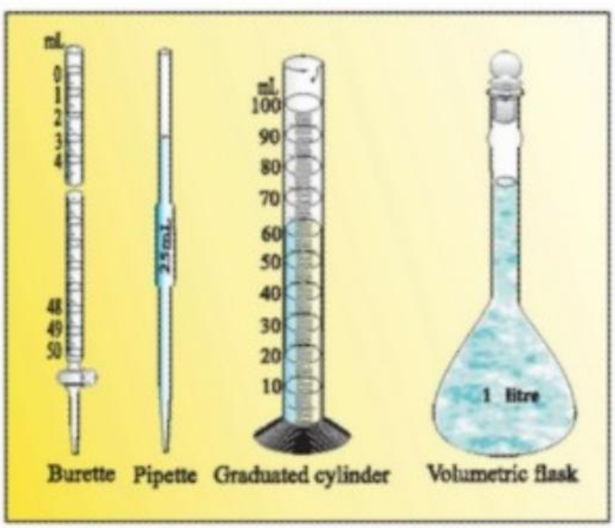

Fig 3: Some volume measuring device

(Source: Chapter 1, page no. 7, XI Textbook, NCERT)

Density: Density of a substance is its amount of mass per unit volume. SI units of density can be obtained as follows:

$$\text{SI unit of density} = \frac{\text{SI unit of mass}}{\text{SI unit of volume}}$$

$$= \frac{\text{kg}}{\text{m}^3} \text{ or kg m}^{-3}$$

This unit is quite large and a chemist often expresses density in g.cm^{-3} , where mass is expressed in gram and volume is expressed in cm^3 . All the three properties discussed above are interrelated as follows:

Density = Mass/Volume

Solids are the materials with highest density that means in solids the particles are closely packed while in liquid, molecules are less tightly packed and hence possess low density. Also in gas, molecules possess low forces of attraction and have large distances between them resulting in very low density.

1.5. Temperature: There are three common scales to measure temperature — °C (degree Celsius), °F (degree Fahrenheit) and K (Kelvin). Here, K is the SI unit. The thermometers based on these scales are shown in Fig. 4.

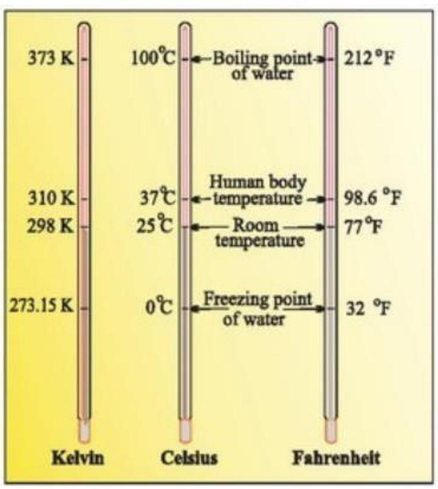

Fig. 4 Thermometers using different temperature scales

(Source: Chapter 1, page no. 7, XI Textbook, NCERT)

Generally, the thermometer with celsius scale are calibrated from 0° to 100° where these two temperatures are the freezing point and the boiling point of water respectively. The Fahrenheit scale is represented between 32° to 212°. The temperatures on two scales are related to each other by the following relationship:

The Kelvin scale is related to Celsius scale as follows:

$$K = °C + 273.15$$

It is interesting to note that temperature below 0°C (i.e. negative values) are possible in Celsius scale but in Kelvin scale, negative temperature is not possible.

2. Uncertainty in Measurement

Chemistry is the study of atoms and molecules which have extremely low masses and are present in extremely large numbers. For example a chemist has to deal with numbers as large as 602, 200,000,000,000,000,000,000 for the molecules present in 2 g of hydrogen gas or as small as 0.00000000000000000000000166g for the mass of single H atom. Similarly other constants such as Planck's constant, speed of light, charges on particles etc., involve numbers of the above magnitude. Repeatedly writing down such numbers for mathematical operations of addition, subtraction, multiplication or division would be tedious job and may result in errors. You can try any simple mathematical operation with any two numbers of the above type then you will really appreciate the difficulty in handling such numbers. A better way to represent such numbers is essential for convenient and accurate calculations. Scientific notation offers such a way.

2.1. Scientific Notation

In scientific notation numbers are represented in exponential notation, i.e., in the form N × 10 n where N is a number (called digit term) which varies between 1.000... and 9.999... and the number n which is called exponent, is a number having positive or negative values. To transform a number larger than 9.999...to scientific notation the decimal point is moved to left until there is one non zero digit before the decimal point. If the decimal point is moved to the left by 'x' places then there is increase in the exponent by x. Example is given below.

Example: For writing number 232.508 in scientific notation we shift the decimal to two places to the left and then exponent n =2 .Thus, this number can be written as 2.32508 $\times 10^2$ in scientific notation.

To transform a number smaller than 1.000...to scientific notation the decimal point is moved to the right until there is one non zero digit before the decimal point. If the decimal point is moved to the right by 'x' places then exponent n= - x. Example is given below.

Example: 0.00016 can be written as 1.6×10^{-4} . Here in the scientific notation the decimal has been shifted to four places to the right and exponent of 10 is (– 4). Similarly the number of molecules in 2g hydrogen given above can be expressed as 6.022×10^{-23} and mass of single hydrogen atom given above can be expressed as 1.66×10^{-24} g in scientific notation. Mathematical operations on numbers expressed in scientific notations. For performing mathematical operations on numbers expressed in scientific notations, the following points are to be kept in mind.

Multiplication and Division

These two operations follow the same rules which are applied for exponential numbers. In multiplication, the digit terms (numbers N) are multiplied and the exponents (n) are added up while in carrying out division, operation of division is performed on digit terms (numbers N) and the exponents (n) are subtracted.

Example-1:

$$(5.6\times10^5)\times(6.9\times10^8)=(5.6\times6.9)\times(10^{5+8})$$
$$= (5.6 \times 6.9) \times (10^{13})$$
$$= 38.64 \times 10^{13}$$
$$= 3.864 \times 10^{14}$$

Example- 2 :

$$(9.8 \times 10^{-2})\times(2.5 \times 10^{-6}) = (9.8 \times 2.5) \times [10^{(-2)+(-6)}]$$
$$= (9.8 \times 2.5) \times (10^{-8})$$
$$= 24.50 \times 10^{-8}$$
$$= 2.450 \times 10^{-8}$$

Example -3:

$$(2.7 \times 10^{-3}) \div (5.5 \times 10^4) = (2.7 \div 5.5) \times [10^{(-3)-(4)}]$$
$$= 4.909 \times 10^{-7}$$

Example- 4:

$$(5.7 \times 10^{-6}) \div (4.2 \times 10^{-3}) = (5.7 \div 4.2) \times [10^{(-6)-(-3)}]$$
$$= (5.7 \div 4.2) \times (10^{-3})$$
$$= 1.357 \times 10^{-7}$$

Addition and Subtraction

For these two operations, first the numbers are written in such a way that they have same exponent. After that, the coefficients (digit terms) are added or subtracted as the case may be.

Example 1: Addition of 6.65×10^4 and 8.95×10^3

$$(6.65 \times 10^4) + (8.95 \times 10^3) = (6.65 \times 10^4) + (0.895 \times 10^4)$$
$$= (6.65 + 0.895) \times 10^4$$
$$= 7.545 \times 10^4$$

Example 2: Addition of 4.56×10^3 and 2.62×10^2

$$(4.56 \times 10^3) + (2.62 \times 10^2) = (45.6 \times 10^2) + (2.62 \times 10^2)$$
$$= (45.6 + 2.62) \times 10^2$$
$$= 58.22 \times 10^2$$

Example 3: Subtraction of 4.5×10^{-3} and 2.6×10^{-4}

$$(4.5 \times 10^{-3}) - (0.26 \times 10^{-3}) = (4.5 - 0.26) \times 10^{-3}$$
$$= 4.24 \times 10^{-3}$$

4. Significant Figures

Every experimental measurement other than counting has some amount of uncertainty associated with it. Suppose we measure the mass of an object on platform balance and obtain the value as 12.3 g. Again we measure the mass of the same object on analytical balance the value of mass obtained is now 12.3028 g. This means that mass obtained by weighing on analytical balance is slightly higher than that obtained by weighing on platform balance. Therefore, in this example there is uncertainty about the number 3 placed after decimal if measurement is done on platform balance. Similarly in the measurement using analytical balance there is uncertainty about last digit which is 8 in the value of mass given above. However, one would always like the results to be precise and accurate. Precision and accuracy are often referred to while we talk about the measurement.

Precision refers to the closeness of various measurements for the same quantity. However, accuracy is the agreement of a particular value to the true value of the result. For example: A true value for a result is 2.00 g and a student 'A' takes two measurements and reports the results as 1.95 g and 1.93 g. These values are precise as they are close to each other but are not accurate. Another student repeats the experiment and obtains 1.94 g and 2.05 g as the results for two measurements. These observations are neither precise nor accurate. When a third student repeats these measurements and reports 2.01g and 1.99g as the result, these values are both precise and accurate. This can be more clearly understood from the data given in Table 4.

Measurements/ g						
	1	2	Average (g)	True Value (g)	Accuracy	Precision
Student A	1.95	1.93	1.940	2.000	low	high
Student B	1.94	2.05	1.995	2.000	low	low
Student C	2.01	1.99	2.000	2.000	high	high

Another example to understand the concept of precision and accuracy is shown in Fig.5.

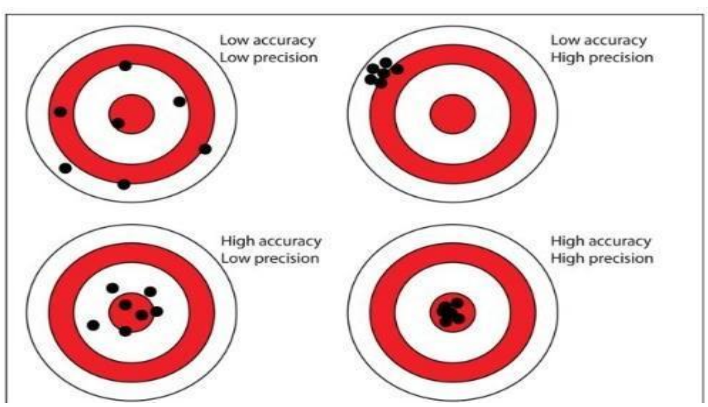

Fig. 5 Precision versus Accuracy

(Source:http://cdn.antarcticglaciers.org/wpcontent/uploads/2013/11/precision_accuracy.png)

The uncertainty in the experimental or the calculated values is indicated by mentioning the number of significant figures. Significant figures are meaningful digits which are known with certainty plus one more digit which is uncertain. It is assumed that when a number is written down, all digits preceding the last one are known with certainty and there is an uncertainty of about one unit in the last digit.

For example if we write a result as 11.2 mL, we say the 11 is certain and 2 is uncertain and the uncertainty would be ± 1 in this last digit. Unless otherwise stated, an uncertainty of ± 1 in the last digit is always understood.

There are certain rules for determining the number of significant figures. These are stated below:

1) All non-zero digits are significant.

   For example 285 cm, 0.25 mL, 5004 and 2.05 have three, two, four and three significant figures respectively
2) Zeros proceeding to first non-zero digit are not significant. Such zero indicates the position of decimal point.
   For example 0.03 and 0.0052 have one and two significant figures respectively
3) Zeros between two non-zero digits are significant. For example 2.005 has four significant figures.
4) Zeros at the end or right of a number are significant provided they are on the right side of the decimal point. For example 0.200 g has three significant figures. The terminal zeros are not significant if there is no decimal point, .e.g., 100 has only one significant figure, but 100. has three significant figures and 100.0 has four significant figures.
5) Such numbers are better represented in scientific notation. We can express the number 100 as $1 \times 10^2$for one significant figure, 1.0×10^2 for two significant figures and 1.00×10^2 for three significant figures. For numbers written in scientific notation, all digits are significant e.g., 4.01×10^2 has three significant figures, and 8.256×10^{-3} has four significant figures.

(1) Counting numbers of objects, for example, 2 balls or 20 eggs have infinite significant figures as these are exact numbers and can be represented by writing infinite number of zeros after placing a decimal i.e., 2 = 2.000000 or 20 =20.000000.

(2) Addition and Subtraction of Significant Figures: When adding or subtracting the decimal places the result cannot have more digits to the right of the decimal point than the number with the least decimal places.

For example:12.11 (2 decimal places)
18.0 (one decimal place)
1.012(three decimal place)
31.122(three decimal place)

Here, 18.0 has only one digit after the decimal point and the result should be reported only up to one digit after the decimal point which is 31.1.

Multiplication and Division of Significant Figures

In multiplication and division operations, the result must be reported up to the same number of significant figures as there are in the number with least significant figures.

Example:
$2.5 \times 1.25 = 3.125$
Since 2.5 has two significant figures which is less than the number of significant figures in number 1.25, therefore result should not have more than two significant figures. Thus, result should be reported as 3.1
While limiting the result to the required number of significant figures as done in the above mathematical operation, one has to keep in mind the following points for rounding off the numbers.

1. If the rightmost digit to be removed is more than 5, the preceding number is increased by one.

For example: 1.386, if we have to remove 6, we have to round it to 1.39

2. If the right most digit to be removed is less than 5, the preceding number is not changed.

For example: 4.334, if 4 is to be removed, then the result is rounded upto 4.33. If the right most digit to be removed is 5, then if preceding number even number then it is not changed. If it is an odd number, it is increased by one.

For example: if 6.35 is to be rounded by removing 5 result should be reported as 6.4. However, if 6.25 is to be rounded off result should be 6.2.

5. Dimensional Analysis: Often while calculating, there is a need to convert units from one system to other. The method used to accomplish this is called factor label method or unit factor method or dimensional analysis.

Example 1: A piece of metal wire is 3 inch (represented by in) long. What is its length in

cm? Solution:

We know, 1 in = 2.54 cm

From this equivalence we can write $\frac{1in}{2.54\,cm} = 1 = \frac{2.54\,cm}{1in}$ thus, $\frac{1in}{2.54\,cm}$ equals 1 and $\frac{2.54\,cm}{1in}$ also equals 1. Both of these are called *unit factors*. If some number is multiplied by these unit factors, i.e. 1, it will not be affected in value.

Say, the 3 in given above is multiplied by the unit factor. So,

$$3\text{ in} = 3\text{ in}\ \frac{2.54\,\text{cm}}{1\text{ in}} = 3\ \ 2.54\text{ cm} = 7.62\text{ cm}$$

Now the unit factor by which multiplication is to be done is that unit factor which gives the desired units i.e., the numerator should have that part which is required in the desired result.

$$\frac{2.54\,\text{cm}}{1\,\text{in}}$$

It should also be noted in the above example that units can be handled just like other numerical part. It can be cancelled, divided, multiplied, squared etc.

6. **Summary:** In this module, we emphasized on measurements of different properties of matter. When the properties of a substance are studied, measurement is inherent. The quantification of properties requires a system of measurement and units in which the quantities are to be expressed. Many systems of measurement exist out of which the English and the Metric Systems are widely used. The scientific community, however, has agreed to have a uniform and common system throughout the world which is abbreviated as SI units (International System of Units). Since measurements involve recording of data which are always associated with a certain amount of uncertainty, the proper handling of data obtained by measuring the quantities is very important. The measurements of quantities in chemistry are spread over a wide range. Hence, a convenient system of expressing the numbers in scientific notation is used. The uncertainty is taken care of by specifying the number of significant figures in which the observations are reported. The dimensional analysis helps to express the measured quantities in different systems of units. Hence, it is possible to interconvert the results from one system of units to another.