CHAPTER 1

MEASURING AND MANIPULATING DATA

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Measuring and Manipulating Data: Units, significant figures, dimensional analysis, precision and accuracy.

OBJECTIVES

- Gain knowledge about significant figures and SI units.
- Get practice of solving problems using dimensional analysis.
- Know the concept of scientific notation and relate it with decimal notation.
- Learn about factor label method.
- Understand the calculations using densities and specific gravities.

SYMBOLS AND UNITS

1.3
Units of measurement

A quantitative observation, or measurement, always consists of two parts: a number and a scale (called unit).

Part 1 – number
Part 2 – Scale (unit)

Every value of a measurement should be indicated along with its proper units, unless the value is dimensionless. The numerical values of various properties of matter can be expressed in different units. The metric system for indicating the units has now been accepted for use. This system is simple and coherent in which larger and smaller versions of each basic unit could be obtained by multiplying or dividing the unit by ten. This is a decimal system, and can be extended to all branches of science, so that all can speak the same scientific "language".

Example:

Length of a rod = 2.8 meters.

Time taken to reach a place = 230 seconds.

International System (SI) of Units

In 1799, after the French revolution, the metric system was first used in France. This system has undergone continuous evolution and improvements since then. Finally in 1960, the 11th conference on weights and measures proposed major changes in the metric system and suggested a new name, viz., the International System of Units, for the revised metric system.

The abbreviation SI, from the French words 'Le System-International', is commonly used for the revised metric system.

The seven basic units in SI system are given in Table 1. All other units are derived from these basic units.

Basic physical quantities and units.

Sl. No.	**Physical Quantity**	**SI Unit**	**Abbreviation or symbol of Unit**
1	Length	Meter	**m**
2	Mass	Kilogram	**kg**
3	Time	Second	**s**
4	Electric Current	Ampere	**A**
5	Temperature	Kelvin	**K**
6	Light Intensity	Candela	**cd**
7	Amount of substance	Mole	**mol**

Instruments are used to make measurement

Measurement	**Instrument**
Length	**Meter Stick or scale**
Volume	**Graduate Cylinder (Low precision)** **Burette, pipette, volumetric flask (Higher precision)**
Mass	**Balance**
Temperature	**Thermometer**

A measurement has a number and a unit., E.g., Mass of Apple = 2 kg

Derived physical quantities and units or derived units

There are a large number of units derived from the basic SI units.

Example:

Area = length x breadth = 4m x 3m = 12 m^2

Special Names and symbols for Certain SI derived Units.

Sl. No.	Physical Quantity	SI Units
1.	Area	'm^2'
2.	Volume	'm^3'
3.	Density	'$kg\ m^{-3}$'
4.	Velocity	'ms^{-1}'
5.	Acceleration	'ms^{-2}'
6.	Force	N (Newton) = $kg\ m\ s^{-2}$
7.	Pressure	Pa (pascal) = Nm^{-2}
8.	Energy	J (Joule) = Nm = $kg\ m^2\ s^{-2}$
9.	Work	$kg\ m^2\ s^{-2}$
10.	Force Constant	Nm^{-1}
11.	Electric conductivity	$Ohm^{-1}\ m^{-1}$
12.	Electric Charge	C (Coulomb)
13.	Dipole Moment	Cm
14.	Frequency	s^{-1} (Hz)
15.	Potential difference	V (volt)
16.	Power	W (watt)
17.	Capacitance	F (Farad)

Subsidiary Units or prefix for SI units

The basic SI units are not always convenient size for a particular measurement. Thus meter would be too big for reporting the thickness of a page but too small for reporting the distance between two places.

For this reason the SI system has recommended the use of subsidiary units which are confined to positive or negative powers of 10 times the basic units.

The prefixes used in the SI system (Those most commonly encountered are shown in blue) (or) The standard prefixes for expressing the decimal fractions

TABLE 1.2 The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)

Prefix	Symbol	Meaning	Exponential Notation*
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

Important facts about SI units

- Abbreviations of units do not have a plural ending. For example it is incorrect to write 70cms. It should be 70cm.

- By conventions the unit named after the scientists, is written by starting with small letter and not with capital letter.

- In chemistry, Angstrom, abbreviated as Å is commonly used as a unit of length because it is almost of the size of an atom. However, it does not belong to International system of units. In SI units we are supposed to use nanometers (nm) (or) picometers.

A number of units, which are decimal fractions, or multiples of derived SI units have special names. Some of these are given in the following table:

Decimal Fractions and Multiples of SI Units having Special Names

Physical quantity	**Name of Unit**	**Symbol for Unit**	**Definition of unit**
length	Angstrom	Å	10^{-10}m
length	Micron	μ	10^{-6}m=μm
force	Dyne	dyn	10^{-5} N
pressure	Bar	bar	10^{5} N m^{-2}
energy	Erg	erg	10^{-7} J
volume	litre	L	10^{-2} m^{5} = 1 dm^{3}

Certain units are not part of the SI are approved for a limited time during the changeover to SI units. Some of these are:

Angstrom 1Å = 0.1 nm = 10^{-10} m

Standard atmosphere 1 atm = 101,325 N m^{-2}
= 1.013 x 10^{5} N m^{-2}

bar, 1 bar = 10^{5} N m^{-2}

Conventions in the use of SI units

1. The symbol for a unit is represented in Roman type.

2. No space is given between the letters of a symbol; for example, the letters k and g are written as kg (kilogram) with no space between the two.

3. Both the singular and the plural forms of a unit are represented by the same symbol without adding "s" to represent the plural. For example, both 1 gram and 10 grams should have the same symbol, g, e.g., 10g and not 10gs.

4. A unit symbol should not be followed by a period (full stop) except when it occurs at the end of a sentence (e.g., 40 kJ and not 40 kJ.).

5. Symbols for prefixes for units should also be printed in Roman type without space between the prefix and the unit, e.g., MeV, not M eV.

6. Double prefixes should be avoided, e.g., Giga Watt (GW), not kilomega Watt (kmW).

7. The combination of a prefix and a symbol should be considered as as a single symbol, which mat be raised to a power without the use of parentheses, e.g., dm^3.

8. More than one slash (/) should not be used in the same unit, e.g., cal deg^{-1} $mole^{-1}$ and not cal/deg/mole.

9. The abbreviated units should be used only when preceded with numeral; in the text without numbers, the units should be written only in the full expanded form.

   e.g., Five kilograms of glucose, not five kg of glucose.

10. Line-symbols, like square root and cube root, should not be used; instead exponents $(\ )^{1/2}$ and $(\ )^{1/3}$ are preferred.

Units in Chemistry

Symbol	**Unit**	**Unit Equivalent**
Length		
m	Meter	---
cm	Centimeter	1×10^{-2} m
mm	Millimeter	1×10^{-3} m
nm	Nanometer	1×10^{-9} m
Å	Angstrom	1×10^{-10} m
µ	Micron	1×10^{-6} m
mµ	Millimicron	1×10^{-9} m
Mass		
kg	Kilogram	1×10^{3} g
g	Gram	-
mg	Milligram	1×10^{-3} g
Volume		
l	Litre	-
ml	Milliliter	1×10^{-3} litre
‘cm^3’	Cubic centimetre	1 ml
‘m^3’	Cubicmeter	---
‘dm^3’	Cubic decimeter	----
µl	Microlitre	1×10^{-6} litre
Heat		
kcal	kilocalorie	1×10^{3} cal
cal	calorie	---
J	joule	---
kJ	kilojoule	1×10^{3} J
Pressure		
‘mm Hg’	Millimeter of mercury	$13.5951 \times 980.665 \times 10^{-2}$ Nm^{-2}
atm	atmosphere	101325 Nm^{-2}
Torr	torr	(101325 / 760) Nm^{-2}

Volume, mass and weight

Volume

One physical quantity that is very important in Chemistry is volume, which is not a fundamenta SI unit but is derived from length.

Volume = L x H x W

Its unit is derived from length m^3

Liter (L) = 1 dm^3

Other units are mL = cm^3, mL.

Mass

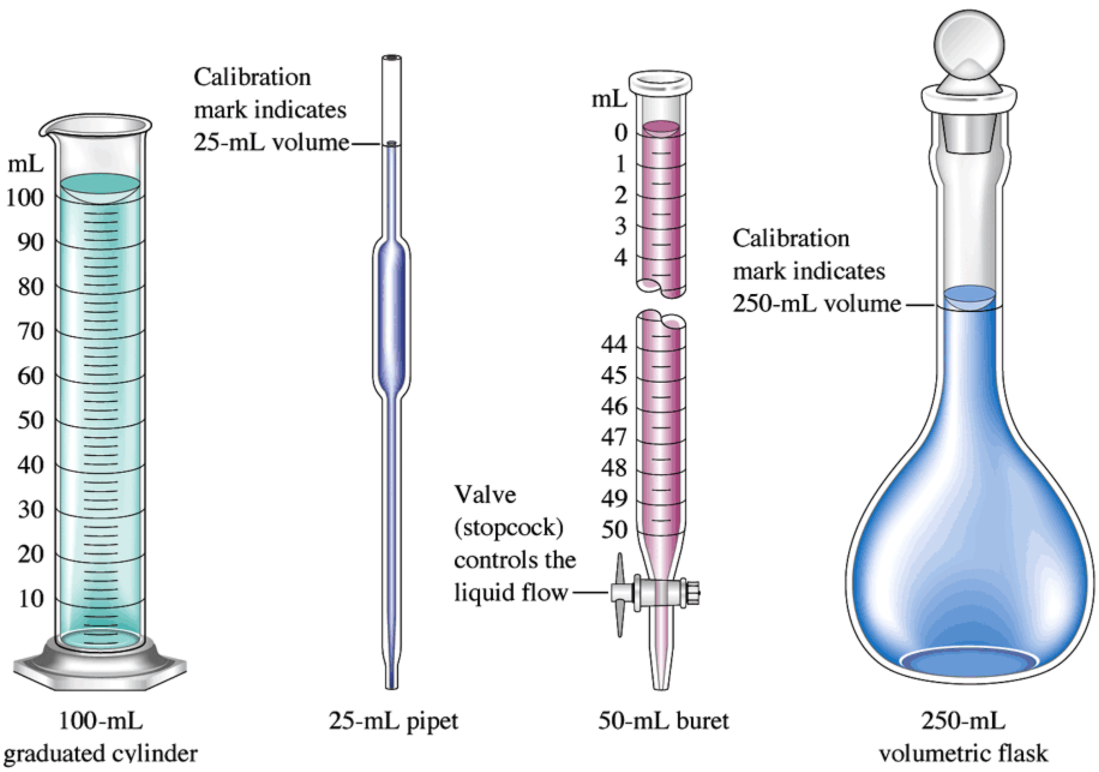
The amount of matter in an object. SI base unit is kg.

Weight

The force exerted on an object by gravity.

Glassware's used in the laboratory for measuring volume

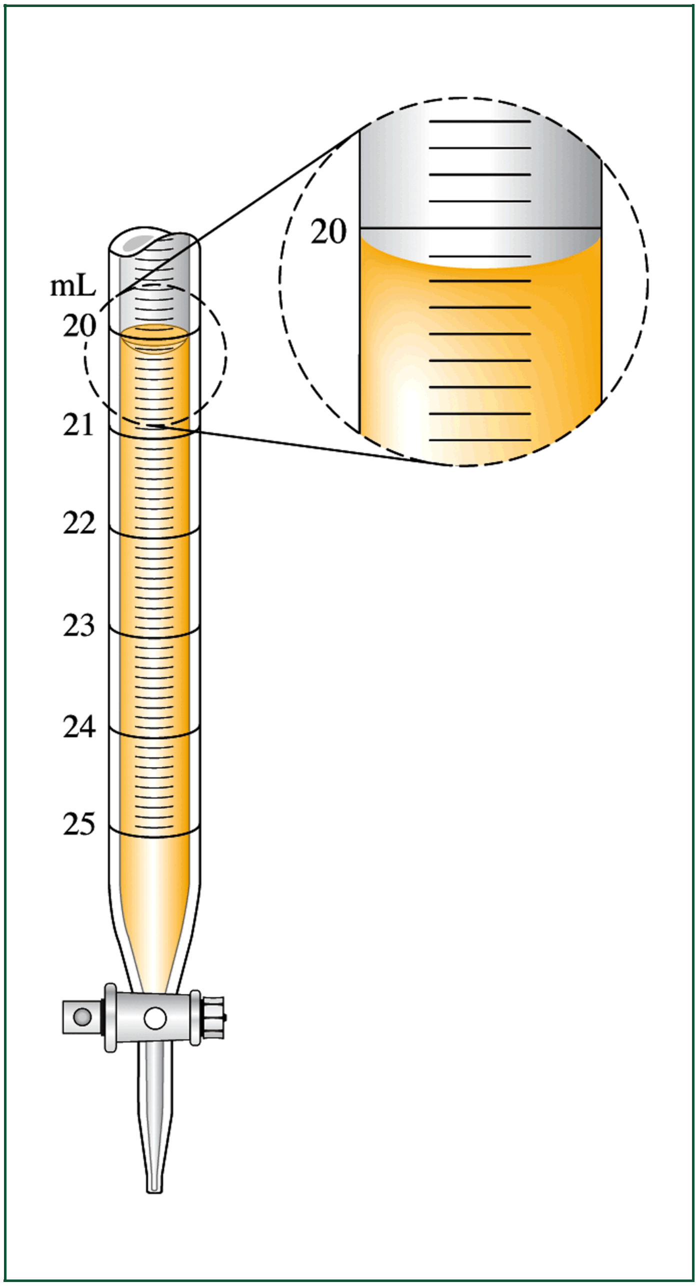
1. **Graduated cylinder**
2. **Pipette**
3. **Burette**
4. **Volumetric flask**



Meniscus:

Measurement of volume using a burette.

The volume is read at the bottom of the liquid curve (called the *meniscus*).



1.4 Uncertainty in measurement

A digit that must be estimated is called uncertain.

A measurement always has some degree of uncertainty.

Example: **The volume of a 50 mL burette is read to two decimal places.**

Person	Results of Measurement of Burette reading
1	20.15 mL
2	20.14 mL
3	20.16 mL
4	20.17 mL
5	20.16 mL

Certain **Uncertain**

The last number is estimated and is therefore uncertain it is called an uncertain digit.

The number written before the uncertain digit does not change and are called certain digits.

- **These results show that the first three numbers (20.1) remains the same regardless of who makes the measurement; these are called certain digits. However, the digit to the right of the 1 must be estimated and therefore varies; it is called an uncertain digit.**

- **We customarily report a measurement by recording all the certain digits plus the first uncertain digit.**

- **It is very important to realize that a measurement always has some degree of uncertainty.**

- **Measurements are reported including all certain digits plus the first uncertain digits, these are called significant figures.**

Precision and Accuracy

Accuracy refers to how close a value is to the true value.
(or)
Accuracy represents the nearness of a measurement to its expected value.

Precision refers to the degree of agreement among several elements of the same quantity.
(or)
Precision refers to how close several measurements are to each other.

Example:

Compare several measurement of the same quantity to each other, if they are close to each other then precision is good.

The mass of an object was measured 5 times to give the following results:

Weighing	Result
1	2.4863 g
2	2.4864 g
3	2.4862 g
4	2.4861 g
5	2.4863 g

If the true mass is known to be 2.6873 g, comment on the accuracy and precision of the measurement.

Range = 2.4681 g to 2.4864 g
Narrow range therefore precise (Precision is good).

Average = (2.4863+2.4864 + 2.4862+2.4861+2.4863) g/5
= 2.4863 g

The average is very different to the true value of 2.6873 g therefore the accuracy is poor.

Error

Any difference between the measured value and the expected value is expressed as the error.

That is difference between accuracy and precision must be noted.

Errors are of two main types:

1. **Determinate Errors or Systematic error. – Error that occur in one direction each time.**

2. **Indeterminate Errors or Random Error. – Have equal probability of being high and low. E.g., the error in estimating the value of the last digit of a measurement.**

Determinate Errors or Systematic Error

These errors are determinable and are avoided if care is taken. Determinate errors have definite values and are classified into three types.

1. Instrumental errors.
2. Operative Errors.
3. Methodic Errors.

1. Instrumental Errors

Instrumental errors are introduced due to the use of defective instruments.

For example: An error in a volumetric analysis will be introduced when a 20ml pipette, which actually measures 20.1 ml, is used.

Sometimes, an instrumental error may arise from the environmental factors on the instrument.

For example: A pipette calibrated at 20°C, if used at 30°C, will introduce error in volume.

Instrumental Errors may be largely eliminated by periodically calibrating the instruments and also using appropriate blanks.

For example: In a measurement using spectrophotometer, the absorbance by the solvent must be checked and proper zero adjustment made, so that the measured absorbance due to the solute will become reliable.

Frequently, fluctuations in current voltage introduce serious instrumental errors.

The possibility of an instrumental error should always be borne in mind whenever an instrument is used to measure property.

2. Operative errors:

These errors are also called personal errors and are introduced because of variation of personal judgments.

Example:

1. Due to colour blindness a person may arrive a wrong results in a volumetric or colorimetric analysis.

2. Using incorrect mathematical equations and committing arithmetic mistakes will also cause operative errors.

Errors of operators due to carelessness, fatigue or inadequate instruction can also lead to operative errors. Such Errors are called gross mistakes.

Arithmetic mistakes, misplacing numbers and decimals in recording data, reading a scale wrongly, reversing a sign and using a wrong scale are some among the gross mistakes.

Frequently, personal errors arise due to bias, i.e., human tendency to estimate readings in a direction that improves the precision of the results.

3. Methodic Errors

These errors are caused by adopting defective experimental methods. Proper understanding of the theoretical background of the experiment is a necessity for avoiding methodic errors. Methodic errors are generally difficult to detect and hence are the most serious of the three types of determinate errors.

Example: In volumetric analysis the use of an improper indicator leading to wrong results is an example for methodic error.

Indeterminate Error or Random errors:

These errors are also called accidental or random errors.

For example: When a student measuring the boiling point of a liquid successively reports the values as 120.0, 119.8, 120.1 and 120.2ºC there is an obvious error in his measurements. This error is called indeterminate error, which cannot be predicted or determined accurately. These errors follow a random distribution.

Indeterminate errors arise from uncertainties in a measurement that are unknown and which cannot be controlled by the experimentalist.

The cause for an indeterminate error, cannot be pinpointed.

For example: When pipetting out a liquid the speed of draining, the angle of holding the pipette, the portion at which the pipette is held etc. would introduce indeterminate error in the volume of liquid pipetted out.

Rules for improving Accuracy of data

1. Every instrument used in the analysis must be checked and calibrated.

2. Blanks should be used along with the test sample whenever possible.

3. A particular measurement should be repeatedly made under stringently identical conditions and its constancy checked.

4. All possible personal errors must be predicted and then eliminated in each measurement.

5. Errors involving arithmetic mistakes, transposition of numbers in recording the data, reading a scale backward, reversing a sign, using a wrong scale and applying a wrong equation must be avoided. The computing equipment itself must be checked for its soundness.

6. The instrument used must be re-read after the value has been recorded and then th enew value must be checked against the one that has been recorded.

7. An experiment for obtaining data should not be performed when the experimentalist is tired or has not received complete instructions.

Gross Errors

These are accidental errors. Droping the sample bottle on the floor, knocking the tape of the burette during titration, or finding a flake of rust fallen into a solution is which traces of iron have to be determined are examples for gross errors.

1.5 SIGNIFICANT FIGURES

Every measurement has some degree of uncertainty associated with it. The magnitude of uncertainty depends on the accuracy of the measuring device as well as on the skill of its operator.

Example:

The magnitude of the uncertainty associated with measurement of length of an object upon accuracy of scale used. It centimeter scale is used to measure the length of a page, we might find that it is 17.7 centimeter. In the above measurements the last digit is doubtful, hence we are using the notation $\pm$ 1 along with the doubtful digit. Hence the uncertainty in length can be expressed as 17.7 ± 0.1 cm. The other method of expressing the uncertainty in measurement is to express it in terms of significant figures.

Definition for significant figures

A measured quantity is expressed in terms of such a number which includes all digits which are certain and a last digit which is uncertain. The total number of digits in the number is called the number of significant figures.

Rules for the determination of significant figures in a measured quantity

Overview:

1. **Non zero Integers.**

2. **Zeros**

   - **Leading zeros (zeros that come before all non zero digits).**
   - **Captive zeros (zeros between non zero digits).**
   - **Trailing zeros (zeros at the right end of the numbers).**

3. **Exact numbers.**

1. Nonzero Integers

Nonzero Integers are always significant.

Example:

1. **1.2347 - has 5 significant figure.**
2. **168 cm - has three significant figures.**
3. **0.168 - has also three significant figures**

2. Zeros. There are three classes of zeros

(a) **Leading zeros (zeros that come before all nonzero digits) are never significant.**

Example:

1. **0.298 – has 3 significant figures.**
2. **0.02 – has 1 significant figures.**
3. **0.007 - has only one significant figures.**
4. **0.026 - has two significant figures.**
5. **0.0266 - has three significant figures.**

(b) **Captive zeros (zeros between nonzero digits) are always significant.**

Examples:

1. **0.203 – has 3 significant figures.**
2. **3.05 - has three significant figures.**
3. **0.205 - has three significant figures.**
4. **0.3006 - has four significant figures.**

(c) **Trailing zeros (zeros at the right end of the numbers) are significant only if the number contains a decimal point.**

The number 100 has only one significant figure, whereas the number 1.00×10^2 has three significant figures.

The number one hundred written as 100. Also has three significant figures.

Examples:

1.
7.0 has two significant figures.
2. **0.070 has two significant figures.**
3. **0.0700 has three significant figures.**
4. **0.4670 has 4 significant figures.**
5. **4680. – has 4 significant figures.**

3. Exact numbers

Many times calculations involve numbers that were not obtained using measuring devices but were determined by counting.

Example: 10 experiments, 3 apples, 8 molecules.

Such numbers are called exact numbers. They can be assumed to have an infinite number of significant figures.

[Exact numbers (numbers in counting and definitions) are considered to have infinite numbers of significant figures and therefore they never limit the number of significant figures in calculations.]

20 molecules, 1 atm = 760 mmHg, 2pr

4. If a number ends in zero that are not to the right of a decimal, the zeros may (or) may not be significant.

Example:
1900 g may have two, three (or) four significant figures.
The uncertainty is the last point can be removed by expressing in scientific notation.

In scientific notation the number is written in the standard exponential form as $N \times 10^n$

N = a number with a single non-zero digit to the left of the decimal point.

'n' = Some Integer.

Hence the mass of 1900 g can be expressed in scientific notation as

1.9×10^3g (Two significant Figures).
1.90×10^3g (Three significant Figures)
1.900×10^3g (Four significant Figures)

5. Rules for Significant Figures in Mathematical Operations
(For multiplication or division)

Multiplication and Division: The number of significant figures in the result is the same as the number in the least precise measurement used in the calculation.

4.56 x 1.4 = 6.38 --Corrected--> 6.4

Limiting term has two significant figures (1.4) **Two significant figures** (6.4)

The Product should have only two significant figures, since 1.4 has two significant figures.

6. **Rules for Significant Figures in Mathematical Operations (Addition or subtraction)**

Addition and Subtraction: the result has the same number of decimal places as the least precise measurement used in the calculation.

For example:

```
12.11
18.0    ←     Limiting term has one decimal place

 1.013
---------     Corrected to
31.123        -------------------> 31.1
---------
```

One decimal place

The correct result is 31.1, since 18.0 has only one decimal places.

Rules for Rounding

1. In a series of calculations, carry one extra digit through to the final result, and then round off.

2. If the digit to be removed:

(a) is less than 5, the preceding digit stays the same.

For example: 1.33 rounds to 1.3.

(b) is equal to or greater than 5, the preceding digit is increased by one.

For example, 1.36 rounds to 1.4.

Exponential Notation

Some numbers are either so very large or so very small that using our ordinary decimal system proves to be inconvenient and it is cumbersome.

For instance it is waste of zero's to write 0.00000000000000000000 472, but it is a vague feeling about how small number. It is simply to write 4.72 x 10^{-20}. A number written this way is said to be written in exponential, or scientific notation. Few facts about numbers written this way.

1. **Each number is written as a product of two numbers. The few numbers are**

   (i) **a coefficient**
   (ii) **10 raised to some power. The power (exponent) should be a positive or negative integer or zero.**

2. **When two numbers written exponentially are to be multiplied (or) divided, first multiply (or) divide the two co-efficient and then add (or) subtract the exponents.**

N x 10^{n}

Where N is a number between 1 and 10

'n' is a positive or negative whole number.

If n is +ve the dp in N is moved n places to the right.

If n is –ve the dp in N is moved n places to the left.

766000000 = 7.66 x 10^{8} (+Ve n)

0.00006215 = 6.215 x 10^{-5} (-ve n)

The total number of significant figures should not change when the number is written in scientific notation.

1.6 DIMENTIONAL ANALYSIS (OR) FACTOR LABEL METHOD

Dimensional Analysis is a technique employed for solving problems. In this technique unit used to express different physical quantities involved in the calculation are analyzed to check whether these quantities are put in the calculation properly or not.

Converting from one unit to Another

- **To convert one unit to another, use the equivalence statement that relates the two units.**
- **Derive the appropriate unit factor by looking at the direction of the required change (to cancel the unwanted units).**
- **Multiply the quantity to be converted by the unit factor to give the quantity with the desired units.**

Proper use of "unit factors" leads to proper units in your answer.

Example 1

Consider the conversion of mass from gram to Kilogram. Suppose a packet of sugar weighs 1200 grams. Let us find out its weight in kilograms.

1000 g = 1kg

Therefore

$$1200 \text{ g contains} = \frac{1200g}{1000g} \times 1kg$$

$$= 1.2 \text{ kg.}$$

The equation given above is called unit conversion factors.

Example 2

Suppose we want to calculate the amount of a substance in moles from the given concentration in moles per litre and volume in litres. This can be achieved as

Amount of substance (mol) = concentration x volume

$$= \frac{mol}{L} \times L = mol$$

Example 3

For the conversion of pounds to kilograms we choose the conversion factor that cancels pounds.

1 kg = 2.2205 pounds.

$$\text{Mass (Kg)} = 1lb \times \frac{1kg}{2.205lb}$$

Therefore

1 lb = 0.45 kg.

1.7 Temperature

SI Unit Kelvin (K)

Other units are degree ^{o}C and Degree Fahrenheit ^{o}F.

Celsius scale $= {}^{o}C$
Kelvin scale $=$ K

Zero Kelvin is the lowest temperature that theoretically can be reached.

Freezing point of water = $0^{o}C$ = 273.15K

$$T(K) = T(^{o}C) + 273.15$$

$$T(^{o}C) = T(K) - 273.15$$

1.8 Density

Density

The density of an object is its mass per unit volume. It is expressed as

$$Density = \frac{mass(m)}{Volume(v)}$$

(or)

$$d = \frac{m}{v}$$

where

d → density (g/mL)
m → mass (g)
v → volume (mL)

Suppose an object has a mass of 15.0 g and a volume of 10.0 cm^3. Substituting you find that

$$d = \frac{15.0g}{10.0cm^3} = 1.50g/cm^3$$

Density is an important characteristic property of a material. The density of some substances are given in the following table.

Sl. No.	Substance	Density (g/mL) at 20°C
1.	Water	0.998
2.	Lead	11.3
3.	Oxygen	1.33×10^{-3}
4.	Ethylene glycol	1.114
5.	Toluene	0.866

In addition to characterizing a substance, the density provides a useful relationship between mass and volume. For example, suppose an experiment calls for a certain mass of liquid, rather than weight the liquid on a balance, you might instead measure out the corresponding volume.

The following example illustrates this idea.

Example

An experiment required 43.7g of isopropyl alcohol. Instead of measuring out sample on a balance, a chemist dispenses the liquid into a graduated cylinder. The density of isopropyl alcohol is 0.785 g/mL. What volume of isopropyl alcohol should be used?

Solution

Rearrange the formula defining the density to obtain the volume

$$v = \frac{m}{d}$$

$$v = \frac{43.7g}{0.785g/mL} = 55.7mL$$

Specific gravity

It is defined as the ratio of density of the material to the density of equal volume of water. It is dimensionless quantity. It is numerically equal to the density of the material.

Relation between various units
(Which will be useful for calculations)

Units of length

1 ft	**=**	**12 inches**		
1 yard	**=**	**3 ft (feet)**	**=**	**0.9144 m**
1 m	**=**	**39.37 in**		
1 m	**=**	**10 dm**		
1 dm	**=**	**10 cm**		
1 cm	**=**	**10 mm**		
1 Å	**=**	**10^{-10}m**		
1 nm	**=**	**10^{-9} m**		
1 pm	**=**	**10^{-12} m**		
1 mm	**=**	**10^{-3} m**		
1 cm	**=**	**10^{-2} m**		
1mm	**=**	**10^{-6}m**		
1 inch	**=**	**2.54cm**		
1km	**=**	**1000 m**		
1 m	**=**	**1.094 yd**		
1 mile	**=**	**1760 yd**		
1 mile	**=**	**5280 ft**	**=**	**1.609 km**

Units of Volume

1 m^3	**=**	**10^3 L**
1 dm^3	**=**	**1 L**
1 cm^3	**=**	**10^{-3} L or 1 mL**

1ft^3	**=**	**28.32 L**
1 L	**=**	**1.06 qt**
1 gal	**=**	**4 qt**

1mL	**=**	**10^{-3} L**
μL	**=**	**10^{-6} L**

Units of Mass

1 kg	**=**	**103g**
1mg	**=**	**10-3g**
1 metric ton	**=**	**103kg**
1lb	**=**	**453.6g**

Common Derived SI Units

Symbol and Meaning		SI Unit
A	**Area**	**m^2**
V	**Volume**	**m^3**
ρ	**Density**	**$kg\ m^{-3}$**
v, u	**Velocity**	**$m\ s^{-1}$**
'a'	**Acceleration**	**$m\ s^{-2}$**
F	**Force**	**Newton (N) = $kg\ m\ s^{-2}$**
P	Pressure	Pascal (Pa) = $N\ m^{-2} = kg\ m^{-1}\ s^{-2}$
E	Energy	J =N m = $Kg\ m^2\ s^{-2}$ =AVs
'k'	Force constant	$N\ m^{-1}$
G	Electric conductivity	$\Omega^{-1}\ m^{-1}$
'q'	Electric charge	C (coulomb) = AV
μ	Dipole moment	C m
γ	Frequency	s^{-1} (Hz)
V	Potential difference	V (volt) = $J\ C^{-1} = kg\ m^2\ s^{-3}\ A^{-1}$
W	Power	W (watt)
C	Capacitance	F (farad)

Some Useful Constants and Their Values in SI and c.g.s units.

Constant	Symbol	Value in SI units	Value in c.g.s. units
Avogadro's number	N_A	6.022 x 10^{23} mol^{-1}	6.022 x 10^{-23} mol^{-1}
Planck's constant	'h'	6.626 x 10^{-34} J s	6.626 x 10^{-27} erg s
Boltzmann constant	'k'	1.3806 x 10^{-23} J K^{-1}	1.3806 x 10-16 erg K^{-1}
Gas constant	$R = kN_A$	8.3144 J K^{-1} mol^{-1}	1.987 cal K^{-1} mol^{-1}
Speed of light	'c'	2.9979 x 108 m s^{-1}	2.9979 x 10^{10} cm s^{-1}
Charge of proton	'e'	1.602 x 10^{-19} C	9.8802 x 10^{-10} e.s.u
Charge of electron	$-e$	-do-	-do-
Electron mass	'm_e'	9.1095 x 10^{-31}kg	9.1095 x 10^{-28} g
Proton mass	'm_p'	1.6726 x 10^{-27} kg	1.6726 x 10^{-24} g
Neutron mass	'm_n'	1.6749 x 10^{-27} kg	1.6749 x 10^{-24} g
Faraday	$F = eN_A$	9.6485 x 104 C mol^{-1}	5.9499 x 10^{14} e.s.u.mol^{-1}
Bohr radius	'a_0'	5.29 x 10-11 m	0.529 Å
Vacuum permittivity	ε_0	8.854 x 10^{-12} J^{-1} C^2 m^{-1}	----
Permittivity factor	$4\pi\varepsilon_0$	1.11265x$10^{-10}$$N^{-1}$$C^2$ m^{-2}	----
Gravitational constant	G	6.672 x 10^{-11} N m^2 kg^{-2}	----
Atmospheric pressure	P	= 1.01325 x 10^5 N m^{-2}	1 atm = 76 cm Hg = 760 mm Hg

IMPORTANT CONVERSION FACTORS

1 cal	=	4.184 J = 4.184×10^{-7} erg
	=	41.293 atm cm^3
Newton	=	10^5 dynes
Joule	=	10^7 ergs = 0.239 cal
Electron volt. eV	=	1.6021×10^{-12} erg $molecule^{-1}$
	=	1.6021×10^{-19} J $molecule^{-1}$
	=	96.48 kJ mol^{-1} = 23.06 kcal mol^{-1}
	=	8065.5 cm^{-1}
1 atm	=	760 mm Hg = 760Torr
	=	101325 N m^{-2} = 101.325 kPa
1 mm Hg	=	1 torr = 133.322 N m^{-2}
1 a.m.u	=	1.66×10^{-24} g = 1.66×10^{-27} kg
	=	931.5 MeV
1 cm^{-1}	=	1.986×10^{-23}J
	=	1.240×10^{-4} eV
ID	=	3.336×10^{-30} C m
1Å	=	10^{-8} cm = 10^{-10} m