SEMESTER–I (Period–I)

C10CH1

1 TOPIC

Introduction to Chemistry

Learning Objectives

Upon completion of this topic, learners will:

- Demonstrate knowledge about the origins and various stages in the development of chemistry
- Express appreciation for the scientific method
- Explain the word Chemistry and other related terminologies
- Distinguish the systems of units of measurement
- Solve simple conversion problems
- Discuss the origin of symbols of element
- Apply the symbols to write the formula and the names of compounds
- Apply the laboratory safety rules and
- Identify apparatus in the Lab.

Introduction

Science can be viewed as a continuing human effort to systematise knowledge for describing and understanding nature. In our daily life, we come across different substances present in nature and changes in them. Formation of curd from milk and rusting of iron are some of the examples of changes which we come across many times. For the sake of convenience,science is sub-divided into various disciplines such as chemistry,physics, biology, geology, etc.

Definition of Chemistry

The word "chemistry" comes from the Arabic word "*al-kimia*" meaning "the art of transformation". *Chemistry is the branch of science that studies the preparation, properties, structure and reactions of material substances.*

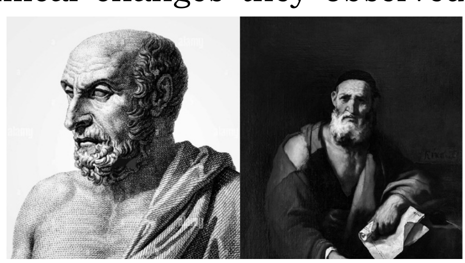
Terminologies Related to Chemistry

Some terminologies related to Chemistry are:

- **Acid:** Any substance that will release hydrogen ions when mixed with water.
- **Atom:** The smallest piece of an element that maintains the identity of that element.
- **Base:** Any substance that accepts proton and releases hydroxide ions (OH^-).
- **Compound:** A combination of more than one element.
- **Element:** A substance that cannot be broken down into simpler chemical substances by ordinary chemical means.
- **Matter:** Anything that has mass and occupies space.
- **Mixture:** A physical combination of more than one substance.
- **Molecule:** The smallest part of a substance that has the physical and chemical properties of that substance.
- **Neutron:** A subatomic particle with no charge.
- **Nucleus:** The centre of an atom that contains protons and neutrons.
- **Proton:** A subatomic particle with a positive charge.
- **Salt:** Any ionic compound that is formed from a reaction between an acid and a base.
- **Solution:** A homogeneous mixture of two or more components in which the particle size is smaller than 1 nm.
- **Solvent:** A chemical substance that dissolves another chemical substance to form a solution.
- **Valency:** The number of electrons that can be donated, accepted or shared by an atom of an element during a chemical reaction.

1.1. DEVELOPMENT OF CHEMISTRY

It was not until the era of the ancient Greeks that we have any record of how people tried to explain the chemical changes they observed and used. At that time, natural objects were thought to consist of only four basic elements: earth, air, fire, and water. In the fourth century BC, two Greek philosophers, Democritus and Leucippus, suggested that matter was not infinitely divisible into smaller



Democritus & Leucippus

particles but instead consisted of fundamental, indivisible particles called **atoms**. Unfortunately, they did not have the technology to test their hypothesis.

Over the next two millennia, alchemists achieved many advances in chemistry. Their major goal was to convert certain elements into others by a process they called **transmutation**.

In particular, alchemists wanted to find a way to transform cheaper metals into gold. Although most alchemists did not approach chemistry systematically and many appear to have been outright frauds, alchemists in China, the Arab kingdoms, and medieval Europe made major contributions, including the discovery of elements such as quicksilver (mercury) and the preparation of several strong acids.


Robert Boyle (1627-1691)

The 16th and 17th centuries saw the beginnings of the modern chemistry. During this period, great advances were made in metallurgy and the first systematic quantitative experiments were carried out. In 1661, Robert Boyle (1627–91) published "*The Sceptical Chymist*," which described the relationship between the pressure and the volume of air. In the 18th century, Joseph Priestley (1733–1804) discovered oxygen gas and found that many carbon-containing materials burn vigorously in an oxygen atmosphere, a process called **combustion**.

In the late 18th century, the French scientist Antoine Lavoisier (1743–94) showed that combustion is the reaction of a carbon-containing substance with oxygen to form carbon dioxide and water. His most important contribution was the '*law of conservation of mass.*' J. L. Proust formally stated the '*law of definite proportions*' in 1797.

In 1803, John Dalton, an English school teacher, expanded the findings if Lavoisier and Proust, and wrote his '*atomic theory of matter.*'

Joseph Gay Lussac (1778–1850) states that gases combine in specific ratio to form a compound.

Amedeo Avogadro (1776–1856) introduced the hypothesis that equal volumes of gases at the same pressure and temperature contain the same number of molecules. It is popularly known as **Avogadro's hypothesis**. It provided the first link between the macroscopic properties of a substance (volume) and the number of atoms or molecules present.

1.1.1. Scientific Method and its Steps

The scientific method is a logical and rational order of steps by which scientists come to conclusions about the world around them. **The scientific method helps to organize thoughts and procedures** so that scientists can be confident in the answers they find. The **steps of the scientific method** are:

1. **Observation:** This step could also be called "research". It is the first stage in understanding the problem you have chosen. For this stage of the scientific method, it is important to use as many sources as you can find. The more information you have; the better the design of your experiment is going to be.

2. **Hypothesis:** The next stage of the scientific method is known as the "hypothesis". This word basically means "a possible solution to a problem, based on knowledge and research". The hypothesis is a simple statement that defines what you think the outcome of your experiment will be.

3. **Prediction:** The hypothesis is your general statement of how you think the scientific phenomenon in question works. Your prediction lets you get specific—how will you demonstrate that your hypothesis is true? The experiment that you will design is done to test the prediction. An **important thing to remember** during this stage of the scientific method is that once you develop a hypothesis and a prediction, you shouldn't change it, even if the results of your experiment show that you were wrong. An incorrect prediction doesn't mean that you "failed". It just means that the experiment brought some new facts to light that may be you hadn't thought about before.

4. **Experimentation:** This is the part of the scientific method that tests your hypothesis. An experiment is a tool that you design to find out if your ideas about your topic are right or wrong. The experiment is the most important part of the scientific method.

5. **Conclusion:** The final step in the scientific method is the conclusion. This is a summary of the experiment's results, and how those results match up to your hypothesis.

Group Activity

In the guidance of your teacher, visit few areas. Investigate and apply scientific method to find possible solution to common Liberian problems.

1.1.2. Contributors of Chemistry

There were many persons who contributed to the development of chemistry. The following table lists some major contributors of chemistry.

Table 1.1. Some Major Contributors of Chemistry

S. No.	Contributors of Chemistry	Their Contribution
1.	Democritus and Leucippus	They suggested that matter was not infinitely divisible into smaller particles but instead consisted of fundamental, indivisible particles called atoms.
2.	Robert Boyle	He formulated the Boyles's law which describes the relationship between pressure and volume of air.
3.	Antoine Lavoisier	His most important contribution was the '*law of conservation of mass.*'
4.	J. L. Proust	He stated the '*law of definite proportions.*'
5.	John Dalton	Wrote atomic theory of matter.
6.	Joseph Gay Lussac	He formulated the Gay-Lussac's Law.
7.	Amedeo Avogadro	He formulated Avogadro's hypothesis.

1.1.3. Branches of Chemistry

Chemistry plays a central role in science and is often intertwined with other branches of science. It has been further divided into different branches depending upon specialised fields of study. The various branches of chemistry are:

(i) **Inorganic Chemistry:** This branch of chemistry deals with the study of compounds of all other elements except carbon.

(ii) **Organic Chemistry:** This branch of chemistry deals with the study of carbon compounds especially hydrocarbons and their derivatives.

(iii) **Physical Chemistry:** This branch of chemistry deals with the explanation of fundamental principles governing various chemical phenomena.

(iv) **Industrial Chemistry:** This branch deals with the chemistry of processes involved in various industrial processes.

In addition to above branches there are other branches of chemistry developed in recent years. These include, *analytical chemistry,* *bio-chemistry, nuclear chemistry, pharmaceutical chemistry, geochemistry, agricultural chemistry, solid state chemistry, etc.*

1.2. UNITS OF MEASUREMENT

Measurement of any physical quantity involves comparison with a certain basic, arbitrarily chosen, internationally accepted reference standard called **unit**. The result of a measurement of a physical quantity is expressed by a number (or numerical measure) accompanied by a unit. Although the number of physical quantities appears to be very large, we need only a limited number of units for expressing all the physical quantities, since they are interrelated with one another. The units for the fundamental or base quantities are called **fundamental** or **base units**. The units of all other physical quantities can be expressed as combinations of the base units. Such units obtained for the derived quantities are called **derived units**.

1.2.1. System of Units

*A complete set of both the base and derived units, is known as the **system of units**.*

Earlier, two different systems of measurement, i.e., the **English System** and the **Metric System** were being used in different parts of the world. The metric system,which originated in France in late 18th century, was more convenient as it was based on the decimal system. Late, need of a common standard system was felt by the scientific community. Such a system was established in 1960 and is known as International System of Unit (SI).

1.2.1.1. 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 at Measures*). The CGPM is an inter governmental treaty organization created by a diplomatic treaty known as Meter Convention which was signed in Paris in 1875. The metric system was found to be more convenient as it was based on the decimal system. The fundamental units of **metric system** are, *gram* **for mass**, *the metre* **for length** and *the litre* for the **volume** of fluids.

The SI system has seven base units which are listed in Table 1.2.

Table 1.2. The Seven Basic SI Units

Physical Quantity	Abbreviation	Name of Unit	Symbol
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	kelvin	K
Luminous intensity	I_v	candela	cd
Amount of substance	n	mole	mol

Decimal fractions or multiples of units are expressed by putting certain prefixes before the names of the units. The various prefixes used for this purpose are listed in the Table 1.3.

Table 1.3. Prefixes for Expressing the Decimal Fractions in the SI System

Multiple	Prefix	Symbol
10^{12}	tera	T
10^{9}	giga	G
10^{6}	mega	M
10^{3}	kilo	k
10^{2}	hecto	h
10	deka	da
10^{-1}	deci	d
10^{-2}	centi	c
10^{-3}	milli	m
10^{-6}	micro	μ
10^{-9}	nano	n
10^{-12}	pico	p

The idea of using prefixes is illustrated by the following example:

A hundredth of a metre, corresponding to 10^{-2}, is a *centimetre* for which the symbol *cm* is used. Similarly, for 10^{3} metre we can use the term *kilometre*, symbolized as *km*.

In addition to the seven basic units, there are many derived units. These units are obtained by combination of basic units. For example, *square metre* (m^2) is the unit for area and the *cubic metre* (m^3) is the unit for volume. Some of the commonly used SI derived units are given in Table 1.4. These derived units are obtained by defining the physical quantities for which they are used.

Table 1.4. Some Common SI Derived Units

Physical Quantity	Name of Unit	Symbol for Unit	Definition of Quantity Unit	Expression in Terms of SI Basic Units
Area	—	—	Length squared	m^2
Volume	—	—	Length cubed	m^3
Density	—	—	Mass per unit volume	kg/m^3 or $kg\ m^{-3}$
Frequency	hertz	Hz	Cycles per second	s^{-1}
Electric charge	coulomb	C	Ampere times second	A s
Electric potential difference	volt	V	Energy per unit charge	$JA^{-1}\ s^{-1}$ or $kg\ m^2\ s^{-3}\ A$

1.2.2. Measurement of Mass, Length, Time, Temperature and Volume

Mass of a substance is the amount of matter present in it. The mass of a substance is constant.

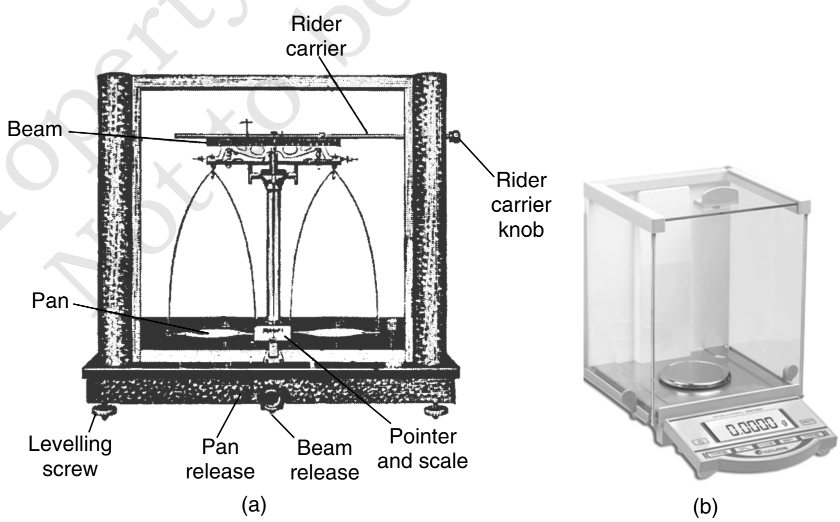


Fig. 1.1. Laboratory Balances

Standard laboratory balances can determine the mass of an object to a minimum of three significant figures.

The mass of a substance can be determined very accurately in the laboratory by using analytical balances (Fig. 1.1).

You are already familiar with some direct methods for the measurement of length.

1.2.2.1. Measurement of Length

You are already familiar with some direct methods for the measurement of length. For example, a metre scale is used for lengths from 10^{-3} m to 10^{2} m. A Vernier callipers is used for lengths to an accuracy of 10^{-4} m. A screw gauge and a spherometer can be used to measure lengths as less as to 10^{-5} m.

1.2.2.2. Measurement of Time

To measure any time interval we need a clock. We now use an **atomic standard of time**, which is based on the periodic vibrations produced in a caesium atom. This is the basis of the **caesium clock**, sometimes called **atomic clock**. Such standards are available in many laboratories. In the caesium atomic clock, the second is taken as the time needed for 9,192,631,770 vibrations of the radiation corresponding to the transition between the two hyperfine levels of the ground state of caesium-133 atom. The vibrations of the caesium atom regulate the rate of this caesium atomic clock just as the vibrations of a balance wheel regulate an ordinary wristwatch or the vibrations of a small quartz crystal regulate a quartz wristwatch.

The caesium atomic clocks are very accurate. In principle they provide portable standard. The national standard of time interval 'second' as well as the frequency is maintained through four caesium atomic clocks.

1.2.2.3. Measurement of 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. 1.2. 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°.

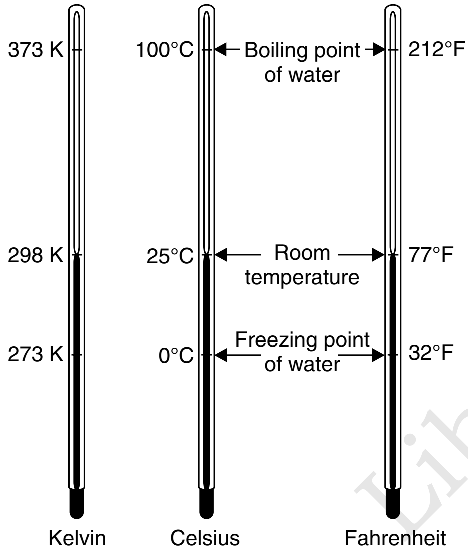


Fig. 1.2. Thermometers based on different temperature scales

The temperatures on two scales are related to each other by the following relationship:

$$°F = \frac{9}{5}(°C) + 32$$

The kelvin scale is related to celsius scale as follows :

$$K = °C + 273.15$$

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

1.2.2.4. Measurement of Volume

Volume has the units of (length)^3. So in SI system, volume has units of m^3. But since smaller volume are used in laboratories, hence the units of cm^3 or dm^3 are used. A common unit, litre (L) which is not an SI unit, is used for the measurement of volume of liquids.

$$1\ L = 1000\ mL \text{ or } 1000\ cm^3$$
$$= 1\ dm^3 = 10^{-3}\ m^3$$
$$1\ cm^3 = 10^{-3}\ dm^3 = 10^{-6}\ m^3.$$

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. 1.3.

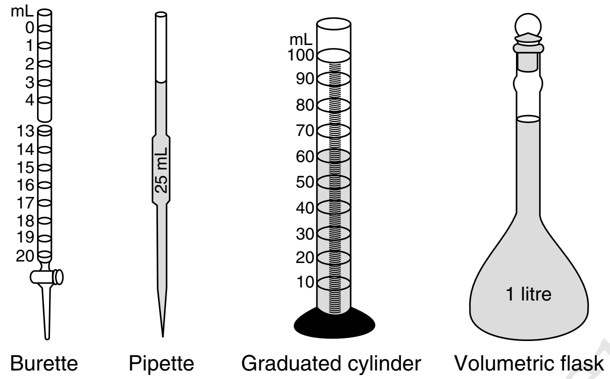


Fig. 1.3. Some volume measuring devices

1.2.3. Conversion of Units

The rule of conversion is "**multiplication is done for converting higher units to smaller units**, whereas **division is carried out for converting smaller units to higher units.**"

The units are expressed using scientific notation and converted into numerical values as per the quantities.

We can convert meter to centimeter by using the conversion factor 1 m = 100 cm. For example,

$$50 \text{ m} = 50 \times 1 \text{ m} = 50 \times 100 \text{ cm} = 5000 \text{ cm}$$

Similarly, we can convert centimeter to meter by using the conversion factor $1 \text{ cm} = \frac{1}{100}$ m. For example,

$$50 \text{ cm} = 50 \times \frac{1}{100} \text{ m} = \frac{50}{100} \text{ m} = 0.5 \text{ m}$$

1.2.3.1. Conversion of Unit of Mass

Below is the conversion table for unit of mass.

Conversion Table for Unit of Mass

1 kg = 10 hg = 100 dag = 1000 g
1 g = 10 dg = 100 cg = 1000 mg
1 dg = 10 cg = 100 mg
1 cg = 10 mg

Example 1.1: *Convert the following:*

(i) *5 g to cg* (ii) *500 g to kg*

Solution:

(i)
$$1\ g = 100\ cg$$
$$5\ g = 5 \times 100\ cg = 500\ cg$$

(ii)
$$1000\ g = 1\ kg$$
$$1\ g = \frac{1}{1000}\ kg$$
$$500\ g = \left(\frac{1}{1000}\right) \times 500\ kg = 0.500\ kg$$

1.2.3.2. Conversion of Unit of Length

Below is the conversion table for unit of length.

Conversion Table for Unit of Length

1 km = 10 hm = 100 dam = 1000 m
1 m = 10 dm = 100 cm = 1000 mm
1 dm = 10 cm = 100 mm
1 cm = 10 mm

Example 1.2: *Convert the following:*

(i) *5 m to cm* (ii) *500 m to km*

Solution:

(i)
$$1\ m = 100\ cm$$
$$5\ m = 5 \times 100\ cm = 500\ cm$$

(ii)
$$1000\ m = 1\ km$$
$$1\ m = \frac{1}{1000}\ km$$
$$500\ m = \left(\frac{1}{1000}\right) \times 500\ km = 0.500\ km$$

1.2.3.3. Conversion of Unit of Time

Below is the conversion table for unit of time.

Conversion Table for Unit of Time

1 minute = 60 seconds
1 hour = 60 minutes= 3600 seconds
1 day = 24 hours
1 week = 7 days
1 year = 365 days

Example 1.3: *Convert the following:*

(i) *30 minutes to seconds*
(ii) *1200 seconds to hour*

Solution:

(i) $$1 \text{ minute} = 60 \text{ seconds}$$
$$30 \text{ minutes} = 60 \times 30 = 1800 \text{ seconds}$$

(ii) $$3600 \text{ seconds} = 1 \text{ hour}$$
$$1 \text{ second} = \frac{1}{3600} \text{ hour}$$
$$1800 \text{ seconds} = \left(\frac{1}{3600}\right) \times 1800 = \frac{1}{2} \text{ hour}$$

1.2.3.4. Conversion of Unit of Temperature

The three main conversions of temperature take place between:

- Celsius and Kelvin
- Fahrenheit and Kelvin
- Celsius and Fahrenheit

Here is the list of the conversion formulas for the different units of temperature.

Conversion Table for Unit of Temperature

Conversion	Formulas
Celsius to Kelvin	$K = °C + 273.15$
Kelvin to Celsius	$°C = K - 273.15$
Fahrenheit to Celsius	$°C = (°F - 32) \times \left(\frac{5}{9}\right)$
Celsius to Fahrenheit	$°F = °C \times \left(\frac{9}{5}\right) + 32$

Example 1.4: *Convert the following:*

(i) *20° Celsius to Fahrenheit*
(ii) *212 Fahrenheit to Celsius*
(iii) *30° Celsius in Kelvin*
(iv) *400 Kelvin to Celsius*

Solution:

(i) Placing the Celsius value in the formula:

$$°F = \left(°C \times \frac{9}{5}\right) + 32 = \left(20 \times \frac{9}{5}\right) + 32$$
$$= 36 + 32 = 68 \text{ °F}$$

Therefore, 20° Celsius is equal to 68° Fahrenheit.

(ii) Placing the Fahrenheit value in the formula:

$$°C = (°F - 32) \times \frac{5}{9}$$

$$= (212 - 32) \times \frac{5}{9} = 180 \times \frac{5}{9}$$

$$= 100\ °C$$

Therefore, 212° Fahrenheit is equal to 68° Celsius.

(iii) Placing the Celsius value in the formula:

$$K = 30 + 273.15 = 303.15$$

Therefore, 30° Celsius is equal to 303.15 Kelvin.

(iv) Placing the Kelvin value in the formula:

$$°C = K - 273.15 = 400 - 273.15 = 126.85\ °C$$

1.2.3.5. Conversion of Unit of Volume

We normally use liter to represent as the standard unit, and the other standard units are in the units in capacity and volume conversion chart.

Conversion Table for Unit of Volume

1 milliliter	=	0.001 liter
1 centiliter	=	0.01 liter
1 deciliter	=	0.1 liter
1 decaliter	=	10 liters
1 hectoliter	=	100 liters
1 kiloliter	=	1000 liters

Example 1.5: *Convert the following:*

(i) *65 milliliter into liter* (ii) *8 hectoliter into liter.*

Solution:

(i) 65 milliliter = 65 × 0.001 liter

= 0.065 liter

(ii) 8 hectoliter = 8 × 100 = 800 liters

1.2.4. Scientific Notations and Significant Figures

1.2.4.1. Scientific Notations

Chemistry is the study of atoms and molecules which have extremely low masses and are present in extremely large numbers. For example, 2 g of hydrogen gas has 602,200,000,000,000,000,000,000 molecules of hydrogen gas. The mass of each H atom is also found to be as small as 0.00000000000000000000000166 g. Similarly, many other constants such as Planck's constant, speed of light, charges on particles, etc. involve numbers of very small magnitude.

Looking at the above two figures and many others like those, it may look funny for a moment to write or count numbers involving so many zeroes but it offers a real challenge to do simple mathematical operations of *addition*, *subtraction*, *multiplication* or *division* with such numbers. Thus, it is always difficult to handle such small numbers.

This problem is solved by using scientific notation for such numbers. The scientific notation is also called **standard form** or **exponential notation**. In this notation, any number can be represented in the form $N \times 10^n$ where n is an exponent having positive or negative values and N is a number that can vary between 1 to 10.

While writing the scientific notation, the decimal in the given number is placed after the first non-zero digit. For doing this, the decimal has to be shifted towards right or left of the original decimal place. The number of places by which the decimal is shifted gives the value of exponent 'n'. Further, n is positive if the decimal is shifted towards left and it is negative if the decimal is shifted towards right. For example,

(i) **Number** : 136.325
Scientific notation : 1.36325×10^2

Here, the decimal has been shifted by two places towards left. Hence, $n = 2$.

(ii) **Number** : 0.00025
Scientific notation : 2.5×10^{-4}

Here, the decimal has been shifted by four places towards right. Hence, $n = -4$.

Example 1.6: *Express the following in the scientific notation.*

(i) *1,86,000* (ii) *0.00683* (iii) *7.0042.*

Solution:

(i) 1.86×10^5 (ii) 6.83×10^{-3} (iii) 7.0042.

1.2.4.2. Significant Figures

Experimental measurements have some uncertainty associated with them. However, one would always like the results to be ***precise*** and ***accurate***. **Precision** means how closely the individual measurements agree with one another. **Accuracy**, on the other hand, means the closeness to the experimental measurements and the true or accepted value. These aspects further depend on the accuracy of **measuring device** and the **skill of the operator**.

For example, the magnitude of the uncertainty associated with measurement of length of an object will depend upon accuracy of scale used. If a centimetre scale is used to measure the length of a page, we might find that it is 14.6 centimetre. In such a measurement we would be sure that the number of *tens* in the numerical expression is 1 and the number of *ones* is 4. However, we cannot be sure of the number corresponding to tenths of a centimetre. This last digit is, therefore, called doubtful digit. One of the method of expressing uncertainty is to use the notation $\pm$ 1 alongwith the doubtful digit. For example, in the above measurement of length, the doubtful digit is 6, which corresponds to the tenths of a centimetre. Hence, the uncertainty in length can be expressed by recording the above measurement as **14.6 $\pm$ 0.1 cm**.

A convenient method of expressing the uncertainty in measurement is to express it in terms of *significant figures* instead of using the notation $\pm$ 1. In this method, it is assumed that all the digits are known with certainty except the last digit which is uncertain to the extent of $\pm$ 1 in that decimal place. Thus, 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**.

The number of significant figures in a measurement is the number of figures that are known with certainty plus one that is uncertain, beginning with the first non-zero digit.

In order to determine the significant figures in a measured quantity the following rules should be applied.

1. ***All non-zero digits are significant.***

   For example, 165 cm has **three** significant figures; 0.165 has also **three** significant figures. Similarly, 2006 has four significant figures, 9.05 has three significant figures, etc.

2. ***Zeros to the left of the first non-zero digit in the number are not significant.***

   For example, 0.005 g has only **one** significant figure, 0.026 g has **two** significant figures.

3. ***Zeros between non-zero digits are significant.***

   For example, 2.05 g has **three** significant figures.

4. ***Zeros to the right of the decimal point are significant.***

   For example, 5.00 g, 0.050 g, 0.5000 g have **three**, **two** and **four** significant figures respectively.

5. ***If a number ends in zeros that are not to the right of a decimal, the zeros may or may not be significant.***

   For example, 1500 g may have two, three or four significant figures.

   The ambiguity in the last point can be removed by expressing the number 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.

For example, a mass of 1500 g can be expressed in scientific notation in the following forms depending upon whether it has two, three or four significant figures.

1.5×10^3 g (**Two** significant figures)

1.50×10^3 g (**Three** significant figures)

1.500×10^3 g (**Four** significant figures)

In these expressions all the zeros to the right of the decimal point are significant. The exponential notation is an excellent way of expressing significant figures in very large or very small measurements. For example, Avogadro's constant is expressed as 6.022×10^{23} mol^{-1} and Planck's constant as 6.62×10^{-34} Js.

Calculations with Significant Figures

During quantitative studies, the scientists have to do calculations with numbers used for various measured physical quantities. These numbers generally, have different number of significant digits depending upon the accuracy with which a particular measurement is made. While carrying out calculations with these numbers, the rule used is that *the accuracy of the final result is limited to the least accurate measurement.* In other words, *final result cannot be more accurate than the least accurate number involved in the calculation.*

Rounding Off

While carrying out calculations with experimentally measured quantities, the final result often contains figures that are not significant. When this happens, the final result is rounded off. In rounding off, the extra digits are dropped with or without minor changes in the figures retained. The rules employed for rounding off a number to the required number of significant digits are as follows:

1. *If the digit following the last digit to be retained is less than five, the last digit is left unchanged.* For example, suppose in the final result obtained as 46.32 only two figures are to be retained as significant figures. The last figure to be retained is 6 and the figure following it is 3. Since 3 is less than five, the 6 will be retained as such without change and the final result would be expressed as 46.
2. *If the digit following the last digit to be retained is more than five, the last digit retained is increased by one.* For example, suppose the result 52.87 is to be rounded off to three significant digits. The last digit to be retained is 8 and the digit following it is 7. Since 7 is more than five, therefore 8 would be increased by one to 9. The result in terms of significant figures would be expressed as 52.9.
3. *If the digit following the last digit to be retained is equal to five, the last digit is left unchanged if it is even and is increased by one if it is odd.*

Example 1.7: *Express the following numbers to four significant figures:*

(i) *5.607892* (ii) *32.392800*

Solution:

(i) As the fifth digit 8 is greater than 5, therefore the result will be expressed as 5.608.

(ii) It will be expressed as 32.39. The digit 2 is dropped and since it is less than 5, the figure is not rounded off to next number.

1.2.5. Accuracy and Precision

1.2.5.1. Accuracy

Accuracy is how close a measured value is to the **actual (true) value**. Accuracy describes how close an approximation is to a correct answer. Thus, accuracy "*is the measure of the difference between the experimental value or mean value of a set of measurements and the true value.*"

Accuracy = *Mean value – True value*

Smaller is the difference between the mean value and *the true value more is the accuracy.*

For *example*, suppose your math textbook tells you that the value of pi (π) is 3.14. You do a careful measurement by drawing a circle and measuring the circumference and diameter, and then you *divide the circumference by the diameter* to get a value for pi (π) of 3.16.

The accuracy of your answer is how much it differs from the accepted value.

In this case, the accuracy is 3.16 – 3.14 = **0.02.**

1.2.5.2. Precision

Precision is how close the measured values are **to each other**. Precision describes how many digits we use to approximate a particular value. In simple words, it is the difference between the measured value and the arithmetic mean value for a series of measurements, i.e.,

Precision = *Individual value – Arithmetic mean value*

Look at the following example:

The value of π (pi) is 22/7. Suppose the value is written as 3.1417 and 3.1392838. Looking at the value, the second number has higher precision, but it would appear that the first is more accurate. (Actual value is 3.142857143.)

1.3. CHEMICAL SYMBOLS, FORMULAE AND NAMING COMPOUNDS

1.3.1. Origin of Symbols

In order to represent the elements, instead of using full lengthy names, scientists use abbreviated names. These abbreviated names of the elements are known as symbols. Thus, ***symbol*** *of an element may be defined as the abbreviation used for the name of the element.*

The abbreviation used for the name of an element is more precisely known as chemical symbol of that element.

Dalton suggested symbols for the atoms of different elements as shown in Fig. 1.4. He was the first scientist to use the symbols for elements in a quantitative sense. When he used symbol for an element he meant a definite quantity of that element, that is, one atom of the element.

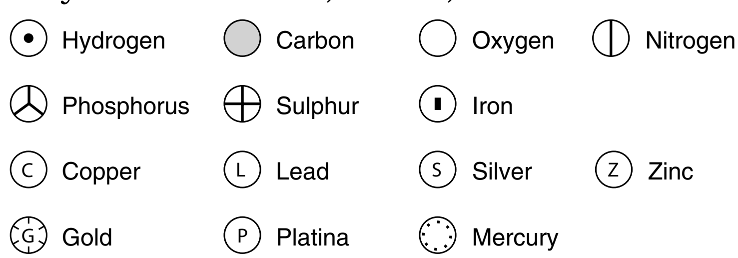


Fig. 1.4. Symbols for some elements as proposed by Dalton

Nowadays, IUPAC (International Union of Pure and Applied Chemistry) approves names of elements. The symbols of elements are generally either the first letter or the first two letters or the first and the third letters of the name of the element. Some symbols derived from the first letter of the names of the elements. For example, Hydrogen (H), Nitrogen (N), Carbon (C), etc.

Some symbols derived from the first two letters of the names of the elements. For example, Aluminium (Al), Barium (Ba), Lithium (Li), etc.

Some symbols derived from the first and the third letters of the names of the elements. For example, Arsenic (As), Magnesium (Mg), Chlorine (Cl), etc.

However, there are certain symbols which seem to have no relationship to their names. The symbols of these elements are, in fact, derived from their latin names. For example, Iron (*Ferrum*) –Fe, Gold (*Aurum*)– Au, etc.

It is important to note that the first letter of every chemical symbol is capital letter but, if the symbol consists of two letters, the second letter is not capital letter. Thus:

Symbol for aluminium is Al and not AL

Symbol for lead is Pb and not PB

1.3.2. Writing Chemical Formula

As already discussed, the elements are represented in the abbreviated form by their symbols. Similarly, a compound is represented in the abbreviated form by its chemical formula.

The expression of the composition of a substance by chemical symbols and numerical subscripts is called the ***chemical formula*** *of the substance.*

The formula of a simple substance is obtained by writing the symbol of the element and indicating the number of atoms in a molecule of the substance by a subscript. For example, a molecule of hydrogen contains two atoms. Hence, hydrogen is represented by the formula H_2. A molecule or ozone contains three atoms of oxygen. Hence, ozone is represented by the formula O_3.

In order to represent the chemical formula of a compound, one must know what are the elements present in the compound and what is the number of atoms of one element that combine with a definite number of atoms of the other element. *For example*, a molecule of water contains two hydrogen atoms and one oxygen atom. It is represented by the formula H_2O.

1.3.2.1. Writing the Formula of a Binary Molecular Compound

In binary molecular compounds, the atoms of the two elements are held by covalent bonds. A covalent bond, as you know, is formed by sharing of electrons between the two atoms. The number of electrons that an atom of the elements contribute for sharing is known as valency of the element. Knowing the valencies of the two elements involved in the formation of a binary molecular compound, the formula of the compound can be derived.

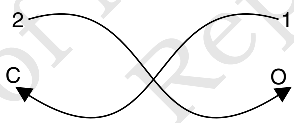
Step 1. *Write the constituent elements and their valencies as shown below:*

Valency	**4**	**2**
Element	**C**	**O**

Step 2. *Reduce the valency numerals to simplest whole numbers by dividing by some common factor, if any*

2	**1**
C	**O**

Step 3. *Criss-cross the reduced valency numerals and write them as subscripts at bottom right hand side of the symbols.*



The subscript 1 is not written. Thus, the formula of the compound is CO_2.

1.3.2.2. Writing the Formula of an Ionic Compound

The formula of an ionic compound represents the simplest whole number ratio of ions in it. The total positive charge on cations is equal to the total negative charge on anions. Knowing the formulae of the ions present in the compound, the formula of the compound can be written by the following steps:

Step 1. *Write the formulae of the ions or radicals of the compound side by side with cation on the left hand side and anion on the right hand side.*

Step 2. *Enclose the compound ion (if any) in a bracket.*

Step 3. *Reduce the valency numerals to a simple ratio by dividing with a common factor, if any.*

Step 4. *Criss-cross the valencies, i.e., shift the valency numerals crosswise to the lower right-hand corner of the ions. This is done to achieve electrical neutrality.*

Let us apply the above steps to write formula of **calcium phosphate**.

Step 1. Writing the formulae of the ions.

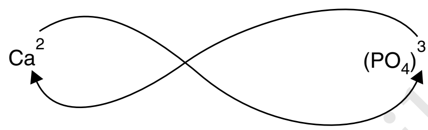
Ca^{2+} PO_4^{3-}

Step 2. Enclose the compound ion phosphate in a bracket.

Ca^{2+} $(PO_4)^{3-}$

Step 3. Not applicable, because ratio is already simple.

Step 4. *Criss-cross the valencies.*



Thus, the formula of calcium phosphate is: $\mathbf{Ca_3(PO_4)_2}$

On the basis of above steps the formulae of some ionic compounds are:

(i) Potassium chlorate ($KClO_3$)

(ii) Lithium nitride (Li_3N)

(iii) Sodium oxalate [$Na_2C_2O_4$]

(iv) Calcium cyanide [$Ca(CN)_2$]

(v) Magnesium nitride (Mg_3N_2)

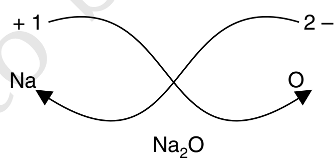
(vi) Potassium manganate (K_2MnO_4)

Example 1.8: *Write down the formulae of*

(i) *Sodium oxide* (ii) *Magnesium hydroxide*

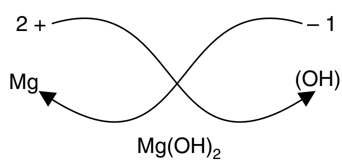
Solution:

(i) Sodium oxide contains Na^+ and O^{2-} ions.



The formula of sodium oxide is $\mathbf{Na_2O}$.

(ii) Magnesium hydroxide



The formula of magnesium hydroxide is $Mg(OH)_2$.

Example 1.9: *Write names of the compounds represented by the following formulae.*

(i) $Al_2(SO_4)_3$ (ii) $CaCl_2$

Solution:

(i) Aluminium sulphate (ii) Calcium chloride

Group Activity: Playing Game for Writing Formula

Make placards with symbols and valencies of the elements separately. Each student should hold two placards, one with the symbol in the right hand and the other with the valency in the left hand. Keeping the symbols in place, students should criss-cross their valencies to form the formula of a compound.

1.3.3. Types of Formula

There are different types of formulas. These include *empirical*, *molecular structural* formulas. In this section, we will study empirical and molecular formulas.

1.3.3.1. Empirical Formula

The empirical chemical formula represents the relative number of atoms of each element in the compound. Some compounds, like water, have the same empirical and molecular formula, because they are small and have the same ratio of atoms in molecules and number of atoms in a molecule. The empirical and molecular formula for water looks like this:

$$H_2O$$

The empirical formula is determined by the weight of each atom within the molecule. Therefore, for a slightly bigger molecule like hydrogen peroxide, the empirical formula shows only the ratio of atoms. In this case:

$$HO:$$

However, this empirical chemical formula only shows the basic foundation of the molecule. In reality, two HO: molecules come together to form a hydrogen peroxide molecule.

1.3.3.2. Molecular Formula

The molecular formula comes in to show the actual number of atoms within each molecule. Thus, for hydrogen peroxide the molecular formula is thus:

$$H_2O_2$$

As you can see, this somewhat confuses the actual structure of hydrogen peroxide. While the empirical chemical formula gives clues that the molecule has two oxygen atoms bonded together in the middle, the molecular formula does not make that clear at all. However, the molecular formula is often used to describe molecules, simply because it is convenient and most molecules can be looked up after their formula is identified.

1.3.4. Naming Compounds

The simplest compounds are binary compounds. *A binary compound is a compound that contains only two elements*. For example, NaCl is a binary compound of sodium and chlorine. CaF_2 is binary compound of calcium and fluorine. It may be noted that a binary compound may contain more than two atoms.

While writing the formula of a binary compound the symbol of the more electronegative element is written on right hand side while that of less electronegative element is written on left hand side. The number of atoms of each element are indicated by subscripts on the right hand side bottom of the symbol. For example, for a compound formed by combination of a metal with a non-metal, the symbol of the metal element is written first (left hand side) and the symbol of the non-metal element is written on right hand side.

While naming the binary compounds, the first element (less electronegative element) is named as such while the name of the second element (more electronegative element) is written with an-ide ending. For example,

KI is named as potassium iodide.

$CaCl_2$ is named as calcium chloride.

The names of some non-metallic elements with-*ide* endings are given below:

Hydrogen	—	Hydride
Fluorine	—	Fluoride
Chlorine	—	Chloride
Oxygen	—	Oxide
Sulphur	—	Sulphide
Carbon	—	Carbide

While naming binary compounds of metals and non-metals, the subscript numerals are ignored. For example, $BaCl_2$ is named barium chloride and not barium dichloride.

On the other hand, *while naming the binary compounds of two non-metals, the subscript numerals have to be taken into consideration* and are indicated as a part of the name. The reason for using prefixes is that the same two non-metallic elements may combine to form many compounds. For example, phosphorus and chlorine combine to form two compounds PCl_3 and PCl_5. A subscript 2 is indicated by the prefix *di*; subscript 3 by *tri*; subscript 4 by tetra, and so on. There should be no gap between the prefix and the name of the element. Some examples are given in Table 1.5.

Table 1.5. Names of some Binary Compounds of two Non-metals

Formula	Name	Formula	Name
CO	Carbon monoxide	CCl_4	Carbon tetrachloride
CO_2	Carbon dioxide	PCl_5	Phosphorus pentachloride
SO_2	Sulphur dioxide	N_2O_4	Dinitrogen tetroxide
PCl_3	Phosphorus trichloride	N_2O	Dinitrogen oxide

1.4. APPARATUS AND SAFETY RULES

1.4.1. Laboratory Apparatus

Experiments cannot be carried out in the absence of scientific apparatus and equipment. Let us discuss about some common laboratory apparatuses.

The common laboratory apparatuses are:

(a) **Beaker:** A beaker is a simple container commonly used in laboratories for mixing and heating liquids. They are generally cylindrical in shape, with a flat bottom.

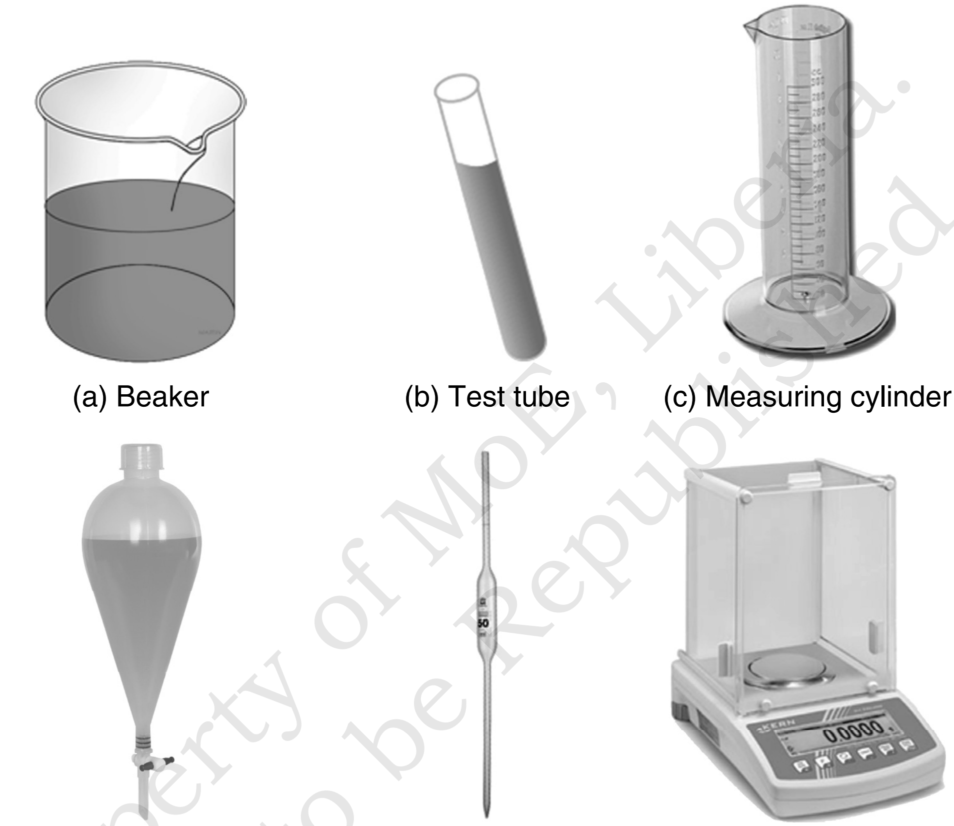
(b) **Test tube:** A thin cylindrical glass tube closed at one end, used to hold small amounts of material for laboratory testing or experiments.

(c) **Measuring cylinder:** A measuring cylinder is used to measure the volume of a liquid in laboratory. It has a narrow cylindrical shape. Each marked line on the measuring cylinder represents the amount of liquid that has been measured.

(d) **Separating funnel:** A separating funnel is a laboratory glassware used in liquid-liquid extractions to separate (partition) the components of a mixture into two immiscible solvent phases of different densities.

(e) **Pipette:** A pipette, pipet, or chemical dropper is a laboratory tool commonly used in chemistry, biology and medicine to transport a measured volume of liquid, often as a media dispenser.

(f) **Balance:** Weighing scales are used to measure weight or mass of the chemicals.



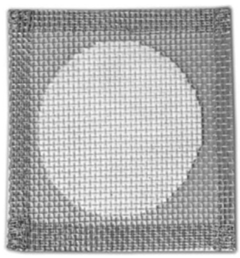
(a) Beaker (b) Test tube (c) Measuring cylinder

(d) Separating funnel (e) Pipette (f) Balance

Fig. 1.5. Some common Laboratory Apparatuses

Besides, the other apparatuses used in a laboratory are test tube stand, wash bottle, China dish, Petri dish, tripod, wire gauze, water condenser and watch glass. We will study different experiments using these laboratory apparatuses in some of the subsequent units.

Test tube stand



Wire gauze

Wash bottle

China dish

Petri dish

Tripod

Fig. 1.6. Some other Laboratory Apparatuses

ACTIVITY 1.1

Making a Poster of Laboratory Apparatuses

Take a drawing sheet and draw various laboratory apparatuses on it. Label these apparatuses properly. Display this poster to the whole class.

1.4.2. Basic Safety Rules in the Laboratory

In a laboratory, improper handling of chemicals and equipment can cause injury or accident. Thus, we must follow safety rules inside the laboratory to avoid accidents and injuries. These rules are:

- Always wear lab-coats and closed shoes.
- Never taste any chemical or transfer chemicals by mouth pipette.
- Read labels carefully before using chemicals.
- Do not use any chemicals in the absence of teacher.
- Do not perform unauthorized experiments.
- Keep your lab space clean and organized.
- Do not leave an on-going experiment unattended.
- Check your glassware for cracks and chips each time you use it.
- Safety glasses must be worn whenever chemicals are being heated, mixed or poured.
- Wash hands before leaving the lab and before eating.
- If any accident (burns) or injury (cut) occurs, immediately call your teacher for help.
- Before leaving a lab unattended, turn off all ignition sources and lock the doors.

Group Activity: Discussing Laboratory Safety Rules

In group, discuss some laboratory safety rules. Students should note these on their notebooks and apply them in the laboratory.

1.4.2.1. Some Hazard Signs

Following are some hazard signs:

Fig. 1.7. Some Safety Hazard Signs

GLOSSARY

- **Accuracy:** The quality or state of being correct.
- **Chemistry:** The branch of science that studies the preparation, properties, structure and reactions of material substances.
- **Combustion:** The process where a substance burns in the presence of Oxygen, giving off heat and light in the process.
- **Hypothesis:** A possible solution to a problem, based on knowledge and research.
- **Metric System:** The decimal measuring system based on the metre, litre, and gram as units of length, capacity, and weight or mass.

- **Precision:** The quality, condition, or fact of being exact and accurate.
- **Transmutation:** Conversion of one chemical element into another.

SUMMARY

- Chemistry is the branch of science that studies the preparation, properties, structure and reactions of material substances.
- In the fourth century BC, two Greek philosophers, Democritus and Leucippus, suggested that matter was not infinitely divisible into smaller particles but instead consisted of fundamental, indivisible particles called atoms.
- The 16th and 17th centuries saw the beginnings of the modern chemistry.
- In 1661, Robert Boyle (1627–91) published “The Sceptical Chymist,” which described the relationship between the pressure and the volume of air.
- Amedeo Avogadro (1776-1856) introduced the hypothesis that equal volumes of gases at the same pressure and temperature contain the same number of molecules.
- Measurement of any physical quantity involves comparison with a certain basic, arbitrarily chosen, internationally accepted reference standard called unit.

EVALUATION

I. Multiple Choice Questions

1. This branch deals with the study of compounds of other elements except carbon.

(a) Physical Chemistry (b) Industrial Chemistry

(c) Inorganic Chemistry (d) Bio-Chemistry

2. How many base units are there?

(a) Three (b) Five

(c) Seven (d) None of these

3. Which of the following is NOT an SI unit?

(a) kelvin (b) metre

(c) ampere (d) yard.

4. The expression of the composition of a substance by chemical symbols and numerical subscripts is called chemical _________.
   (a) symbol (b) formula
   (c) Both (a) & (b) (d) None of these.
5. In a laboratory, improper handling of chemicals and glassware can cause
   (a) injury (b) accident
   (c) both (a) and (b) (d) neither (a) nor (b).

II. State True or False

1. Antoine Lavoisier published "*The Sceptical Chymist,*" which described the relationship between the pressure and the volume of air.
2. The scientific method helps to organize thoughts and procedures.
3. Joseph Gay Lussac stated the law of definite proportions.
4. Stapler and calculator are examples of laboratory apparatus.
5. We should not return chemicals or reagents to bottles.

III. Answer the Following Questions

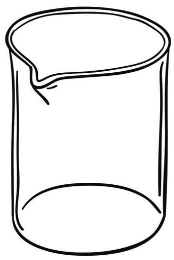
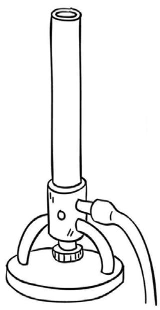
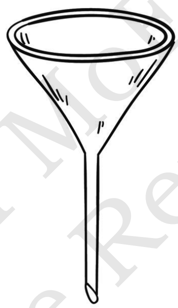
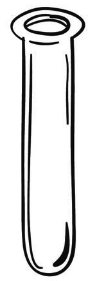
1. Define chemistry.
2. Convert each of the following:
   (i) 5 gram to centigram
   (ii) 25000 millimeter to meter
   (iii) 1200 seconds to minutes
   (iv) 40° Celsius to Fahrenheit
3. Express the following in the scientific notation.
   (i) 3,86,000 (ii) 9007
   (iii) 0.02683 (iv) 9000.0
4. Express the result of the following calculation to the appropriate number of significant digits.
   $816 \times 0.02456 + 215.67$.
5. Write at least five symbols of atoms which were proposed by Dalton.
6. Write down the formulae of:
   (i) aluminium chloride (ii) magnesium hydroxide.
7. Write names of the compounds represented by the following formulae.
   (i) K_2SO_4 (ii) KNO_3

8. Give the names of the elements present in the following compounds:

(i) Baking powder (ii) Aluminium chloride

9. Write five safety rules of laboratory.

10. Identify the following objects and write their names in your notebook.

		
(a)	(b)	(c) ....
		
(d)	(e)	(f) ....