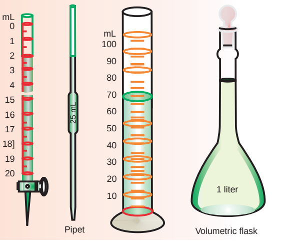
UNIT 1



Fundamental Concepts in Chemistry

Unit Outcomes

At the end of this unit, you should be able to:

- *understand the scope of chemistry;*
- *select and use appropriate SI units;*
- *understand the causes of uncertainty in measurement;*
- *express the result of any calculation involving experimental data to the appropriate number of decimal places or significant figures;*
- *use scientific methods in solving problems;*
- *demonstrate an understanding of experimental skills in chemistry;*
- *demonstrate a knowledge of laboratory procedures; and*
- *demonstrate scientific enquiry skills including observing, inferring, predicting, comparing and contrasting, communicating, analyzing, classifying, applying, theorizing, measuring, asking questions, developing hypotheses, performing and designing experiments, interpreting data, drawing conclusions, making generalizations and problem solving.*

MAIN CONTENTS

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- Prefixes used in SI Units
- Uncertainty in Measurement
- Precision and Accuracy
- Decimal Places
- Significant Figures
- Scientific Notation

1.3 Chemistry as Experimental Science
- Scientific Method
- Experimental Skills in Chemistry
- Writing a Laboratory Report

Start-up Activity

Form a group and collect as many objects as possible available to you like pen, paper, lunchbox, chalk, mobile phone, belt, towel, etc. Now, discuss the following questions and share your ideas with rest of the class.

1. Are all the objects made up of the same material? Try to classify them as plastic, leather, metal, wood, paper, etc.

2. Do these objects have the same use? Do they have multiple usages?
3. How would you decide which object is more appropriate for a particular use?
4. Can you use these objects at all temperatures?
5. What properties do you look for in a material for a particular use?
6. Are all the objects biodegradable?

   Extend your discussion to describe the extent to which chemistry is used to understand matter and our environment.

INTRODUCTION

People in the industrialized nations enjoy not only the highest standard of living, that is, the material comforts which are measured by the goods, services and luxuries available to an individual, but also quality life. Quality life depends upon business and employment, services, health and nutrition, population, leisure time in addition to standard of living. Much of this is due to chemistry.

Chemistry enables us to design all sorts of materials: drugs to fight diseases; pesticides to protect crops and our health; fertilizers to grow abundant food; fuel for transportation; fibres to provide comfort and variety in clothes; building materials for housing; plastics for diverse uses; and much more.

When we address ourselves the most fundamental question: *What is the nature of life*? Chemistry provides essential information on this subject. The theories of chemistry illuminate our understanding of the material world from tiny atoms to giant polymers.

Everything you see, smell, taste and touch is made up of matter. Even the way you perceive the world through your senses involves chemical reactions. With such an enormous range of topics, it is necessary to know about chemistry in order to understand the world around us.

Figure: 1.1 Chemistry in everyday life.

Almost everything around us involves chemistry. The world around us consists of compounds made of various elements. Human body consists mainly of carbon, oxygen and hydrogen. Our environmental issues like global warming, ozone layer depletion, acid rain, etc., also require understanding of the fundamentals of chemistry.

1.1 THE SCOPE OF CHEMISTRY

At the end of this section, you should be able to:

- define chemistry;
- distingiuish the major fields of chemistry; and
- distinguish the sub-divisions of the branches of chemistry.

1.1.1 Definition of Chemistry

Activity 1.1

Form a group and enlist the activities you perform daily such as brushing your teeth, getting ready for school, enjoying your meals, studying in the school, playing, off to bed etc.

Now, discuss the following questions and share your ideas with rest of the class.

1. What is your tooth brush, toothpaste and soap made of?
2. Why do soaps have a cleansing action?
3. What makes petrol/CNG (compressed natural gas) a better fuel than wood?
4. Is it possible to use wires made of rubber for conduction of electricity?
5. What properties of cement, iron and stone make them suitable for construction of houses, etc. but not for making an aircraft?
6. Why metals like gold and silver are preferred for making jewellery?
7. Why is it necessary to cook certain food items?

Chemistry is the science that deals with matter and the changes that it undergoes. It is a study of the composition, structure, and properties of matter and of the changes that occur in matter.

Perhaps the only permanent thing in the world is change. Iron rusts, snow melts, paints peel off and firewoods burn. We grow up, we grow old. Living plants and animals undergo ceaseless change, and even dead animals and plants continue to change as they decay. Such changes fascinated people and inspired them to look more closely at nature's way of working.

Understanding change is closely related to understanding the nature and composition of matter- the physical material of the universe. Matter is anything that occupies space and has mass.

It has long been known that matter can change or be made to change from one form to another. These changes are broadly classified into chemical and physical changes. Chemical changes, more commonly called as chemical reactions are processes whereby one substance is transformed into another as a result of combination or dissociation of atoms. We can describe the transformation both qualitatively and quantitatively with the help of chemical equations for the reaction. Some of the examples of chemical change include oxidation of matter (*rusting, burning*), fermentation, changing milk in to yogurt, and addition of water to calcium oxide.

Matter also undergoes other kinds of changes called physical changes. These changes differ from chemical reactions in that the involved substances do not change their identities. Each retains its composition. Most physical changes are accompanied by changes in physical state, such as the melting of solids and the boiling of liquids. For example, water remains H_2O whether it is in solid state (*ice*), liquid water or gaseous state (*steam*). Physical change also involves making or separating mixtures. Dissolving table salt (NaCl) in water is a physical change.

There are two kinds of physical properties, namely, extensive and intensive physical properties.

Extensive physical properties are the properties, which depend on the amount or quantity of sample and therefore, can vary from sample to sample. The extensive property of a piece of copper wire, for instance, includes its length, diameter, mass, and electrical resistance.

Intensive physical properties are properties which do not depend on the amount of a substance present. The intensive properties of a piece of copper wire include its density, colour, melting point, and hardness. Intensive properties are useful in distinguishing between different substances because they do not vary from sample to sample.

Exercise 1.1

1. How do physical and chemical changes differ?
2. Classify the following properties of a piece of copper foil into extensive and intensive physical property:
   - Thickness
   - Area
   - Conductivity
   - Specific gravity
   - Solubility
   - freezing point
   - Weight of a substance
   - Smell (odour)
3. Describe importance of chemistry with the help of examples.

1.1.2 Major Fields of Chemistry

The universe is just like a very big chemical laboratory, rearranging atoms and sub-atomic particles to produce elements and compounds. While planets are made up of rocks which are nothing but arrangement of compounds, an atmosphere is a mixture of compounds separated by distance.

Since chemistry is such an enormous area of science, for convenience it has been divided into disciplines. However, the division is never as clear-cut as it might appear to be. All sciences are related and depend on each other – they are interrelated.

All the disciplines of science share information and methods with each other. For example, biology uses the findings of both physics and chemistry to study living organisms. Chemistry utilizes the information gathered by physics about the nature of matter and energy to study the properties and interactions of substances.

There are several branches of chemistry, the major branches are, Inorganic chemistry, Organic chemistry, Physical chemistry and Analytical chemistry.

- **Inorganic chemistry** is the study of all the elements and their compounds with the exception of carbon and its compounds (which falls under the category of organic chemistry). It investigates the characteristics of substances that are not organic, such as nonliving matter and minerals found in the earth's crust. Oxides, sulphides and carbonates form the important classes of inorganic compounds.

- **Organic chemistry** is the chemistry of carbon compounds except carbides, cyanides, carbon dioxide, carbon monoxide, carbonates and hydrogen carbonates. Perhaps the most remarkable feature of organic chemistry is that it is the chemistry of carbon and a few other elements, chiefly, hydrogen, oxygen, nitrogen, halogens and sulphur. The major nutrients in the food comprises of organic compounds such as carbohydrates, proteins, fats, vitamins, etc.

- **Physical chemistry** is the study of physical properties of materials, such as their thermal, electrical and magnetic behaviour and their interaction with electromagnetic fields. A chemical system can be studied from either a microscopic or a macroscopic point of view. The microscopic point of view makes explicit use of the concept of molecules. The macroscopic point of view studies large-scale properties of matter without explicit use of the molecule concept. Some important divisions of physical chemistry are thermodynamics, spectroscopy, quantum chemistry, chemical kinetics and electrochemistry.

- **Analytical chemistry** is a branch of chemistry which is concerned with the development of theoretical foundations and methods of chemical analyses. It involves separating, identifying and determining the relative amount of components in a sample of material. Chemical analysis may be qualitative or quantitative. Qualitative analysis reveals the chemical identity of the species in the sample while quantitative analysis establishes the relative amount of one or more of these species in numeric terms.

There is yet another important branch of chemistry, which bridges chemistry and biology, known as biochemistry. It involves the study of the science of the molecules and chemical reactions of life, and utilizes the principles and language of chemistry to explain biology at the molecular level.

Activity 1.2

Form a group and perform the following activities:

1. Investigate the ways in which the major areas of chemistry are further subdivided. You can use reference books and the internet to augment your current ideas.
2. Discuss the principles of chemistry involved in the daily-life and share your ideas with the rest of the class.

Exercise 1.2

1. Are the three states of matter inter-convertible? What type of change will it be?
2. Classical, alchemical, medical and technological traditions were chemistry's forerunners. Identify the contributions which each of these made to the development of chemistry.

1.2 MEASUREMENTS AND UNITS IN CHEMISTRY

At the end of this section, you should be able to:

- list and describe the seven SI units and their prefixes;
- write the names and symbols of derived SI units;
- use the factor-label method for solving problems and conversion to SI units;
- describe uncertainty of measurement;
- identify the digits that are certain and the ones that are uncertain given a number representing a measurement;
- identify the causes of uncertainity in measurement;
- define precision and accuracy;
- estimate the precision that is possible for any instrument, you use in the laboratory;
- explain system errors and random errors;
- analyze the given data in terms of precision and accuracy;
- define decimal places;
- determine the number of decimal places in a calculated result;
- define significant figures;
- determine the number of significant figures in a calculated result; and
- use scientific notation in writing very large or very small numbers.

1.2.1 SI Units (The International System of Unit)

In order to test a hypothesis, a scientist must gather data by measurement. Before the hypothesis is accepted, other scientists must reproduce the measured data. Data gathering and checking are much easier to accomplish if all scientists agree to use a common system of measurement. The system that has been agreed upon since 1960 is the international system of units (Système International d'Unités). The System International is a set of units and notations that are standard in science. It is a modernized version of the metric system that was established in France in 1795.

A unit of measurement is a definite magnitude of a physical quantity (mass, length, temperature, etc.) that has been chosen as the standard against which other measurements of the same physical quantity are made. For example, metre is the unit of measurement for length in the metric system.

All measurements consist of two parts: a scalar (numerical) quantity and the unit designation. For example, when an object is 2 metres long, it means that the object is two times as long as the unit standard (1 metre). In this example, the scalar quantity is 2 and the unit designation is metre.

Basic SI Units

In chemistry, two systems of units were commonly used for expressing the fundamental physical quantities such as mass, time, and length. They are:

- The cgs system – centimetre-gram-second
- The mks system – metre-kilogram-second

In the cgs system, the basic unit of length is centimetre (cm), mass is gram (g), and of time is second (s). In the mks system, the basic unit of length is metre (m), of mass kilogram (kg), and of time is second (s). In this way, each system defines individual base units for each of the fundamental physical quantities. All measured quantities can be expressed in terms of the seven base units listed in Table 1.1.

Table 1.1 The seven SI base units

Physical Quantity	SI Base unit	Symbol of unit
Mass	kilogram	kg
Length	metre	m
Time	second	s
Temperature	kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A
Luminous intensity	candela	Cd

From these seven base units, all except candela, are relevant to chemistry.

Activity 1.3

Form a group and perform the following activity.

Take a small piece of magnesium ribbon. Measure its mass and length. Now put it in 20 mL of dilute hydrochloric acid and measure the time required for magnesium ribbon to dissolve completely. Record the temperature of solution before and after putting the magnesium ribbon in it.

Discuss the following questions:

1. What instruments/equipments did you use for measuring the physical quantities?
2. What units did you use to express them?
3. What is the difference between the physical quantities, namely mass and length?
4. Which is the appropriate unit to express the time taken for the above reaction to go to completion?
5. Explain the difference between heat and temperature.
6. Which basic SI units are appropriate to express the:
    a length of a race track,
    b average room temperature, and
    c time duration for the earth to have one rotation around its axis?

i) Mass

Mass of an object is the amount of matter present in it. It is measured with an analytical balance and in contrast to weight, mass is not affected by gravity.

ii) Length

The SI base unit of length is the metre (m). To measure length much larger than the metre, we often use the kilometre (km).

In the laboratory, lengths smaller than a metre are often most convenient. For example, the centimetre (cm) and the millimetre (mm).

On the submicroscopic scale, the micrometre (μm), the nanometre (nm), etc., are used.

iii) *Time*

The SI base unit for measuring intervals of time is the second (s). Short times are expressed through the usual SI prefixes: milliseconds (ms), microseconds (μs), nanoseconds (ns), and picoseconds (ps). Long time intervals, on the other hand, are usually expressed in traditional, non – SI units: minute (min), hour (h), day (d), and year (y).

iv) *Temperature*

Temperature is a measure of the average energy of motion or kinetic energy, of a single particle in a system. The instrument for measuring temperature is called thermometer. From common experience, we know that if two objects at different temperatures are brought together, heat flows from the warmer to the colder object. For example, if you touch a hot test tube, heat will flow from the test tube to your hand. If the test tube is hot enough, your hand will get burned. The temperature of the warmer object drops and that of the colder object increases, until finally the two objects are at the same temperature (thermal equilibrium). Temperature is therefore a property that tells us in what direction heat flows.

The SI basic unit of temperature is the Kelvin (K). For most routine laboratory work, we can use a more familiar temperature scale: the Celsius scale. On this temperature scale, the freezing point of water is 0°C, and its boiling point is 100°C. Another temperature scale, probably unfamiliar to most people, is the Fahrenheit scale. The relationship between these three temperature scales is given below:

$$°C = \frac{5}{9}(°F - 32)$$

$$°F = \frac{9}{5}°C + 32$$

$$K = °C + 273.15$$

The unit for temperature in the Kelvin scale is Kelvin (K, NOT k!).

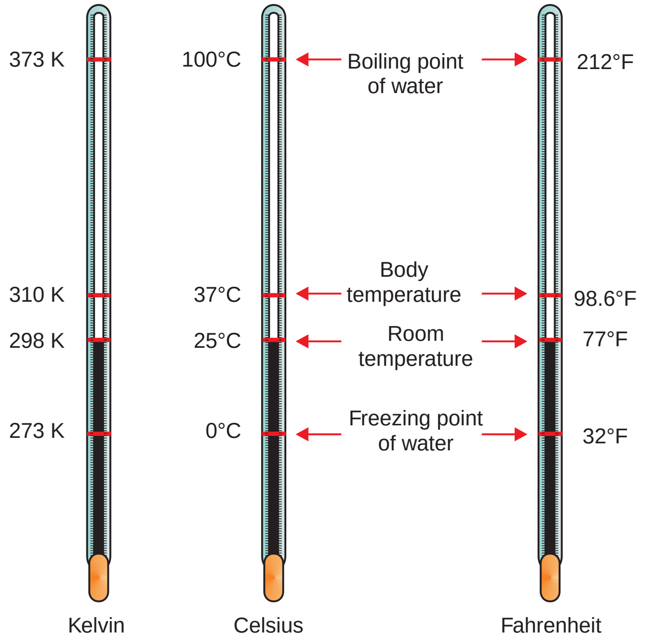


Figure1.2 The three temperature scales.

Activity 1.4

Form a group and describe the difference between mass and weight.

Share your ideas with the class.

The Kelvin scale assigns a value of zero Kelvin (0K) to the lowest possible temperature, which is called absolute zero and corresponds to –273.15°C. Note that the term absolute zero is used because this is a hypothetical temperature characterized by complete absence of thermal (kinetic) energy.

v) Mole (Amount of Substance): A mole of any substance (atoms, molecules or ions) represents 6.023×10^{23} particles of that substance. This number is also known as Avagadro's constant (N_o)

Exercise 1.3

1. The average temperature in Addis Ababa, during the summer, is about 25°C. What is the equivalent Kelvin temperature?
2. A parasite that causes trichinosis is killed when meat is cooked to 66°C. Assume you have only a Fahrenheit thermometer. Determine the minimum Fahrenheit temperature to which the meat should be heated when it is being cooked.

Derived SI Units

People often say that gold is "heavy" and aluminium is "light". Do they mean that a gold bracelet weighs more than an aluminium extension ladder?

Derived physical quantities are expressed in derived SI units. Although units used to express derived physical quantities are actually derived from basic SI units, they are often given special names for convenience. For example, force, volume, density, concentration, pressure, area, energy, etc., are derived quantities.

i *Force* is the product of mass and acceleration.

$$\text{Force} = \text{mass} \times \text{acceleration}$$
$$= \text{kg} \times \text{m/s}^2 = \text{kg m s}^{-2}$$

Therefore, kilogram–metre per second squared is the SI unit of force. This combination of units is called the Newton (N).

$$1 \text{ N} = 1 \text{ kg m s}^{-2}$$

ii *Volume* is the amount of space occupied by a solid, liquid or gas. The volume of a liquid can be measured by using graduated (measuring) cylinder, a burette, or a pipette while a volumetric flask is used to take measured volume of the liquid.

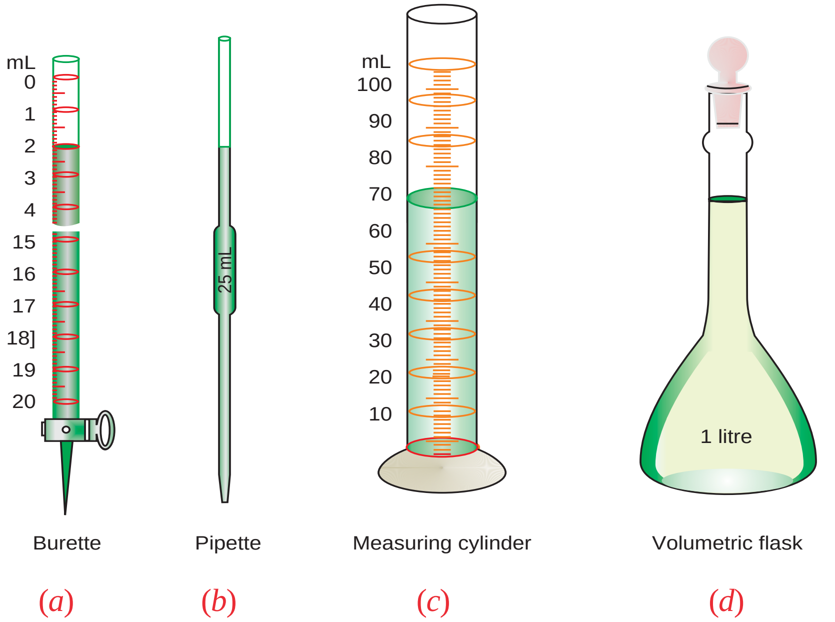


Figure 1.3 Some commonly used measuring apparatus.

Volume is a derived quantity in chemistry, so it is given a special unit, the litre (L). In SI units, one litre is defined as being equal to 1000 cubic centimetres (cm^3).

$$1L = 1000\ cm^3$$

$$1mL = 1\ cm^3$$

$$1L = 1\ dm^3$$

so,
$$1000\ cm^3 = 1dm^3$$

The volume of a solid object with a rectangular shape can be calculated as:

$$\text{Volume} = \text{length} \times \text{width} \times \text{height}$$

iii *Density* is the amount of mass in a unit volume of matter. Its symbol is ρ.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad \rho = \frac{m}{v}$$

Density can be measured in units of $g\ cm^{-3}$, $kg\ m^{-3}$ or g/mL.

For example, 1.00 g of water occupies a volume of 1.00 cm^3 or 1 mL.

$$\rho = \frac{m}{v} = \frac{1.00\,g}{1.00\,cm^3} = 1.00\ g\ cm^{-3}$$

It may be noted that the density of a substance is always measured at specific temperatures.

Example 1.1

Aluminium has a density of 2.70 g cm^{-3}. What is the mass of a piece of aluminium with a volume of 0.525 cm^3?

Solution:

Since $\rho = m/V$, it follows that $m = \rho \times V$

$$m = 2.70\ g\ \cancel{cm^{-3}} \times 0.525\ \cancel{cm^{-3}}$$

$$= 1.42\ g$$

Exercise 1.4

1. Ethanol is used in alcoholic beverages and has a density of 0.789 g/mL. What volume of ethanol (in litres) would have a mass of 500 g?
2. Calculate the density of a rectangular block of metal whose length is 8.335 cm, width is 1.02 cm, height is 0.982 cm and mass is 62.3538 g.
3. A piece of silver metal weighing 194.3 g is placed in a graduated cylinder containing 242.0 mL of water. The volume of water now reads 260.5 mL. Calculate the density of the metal.
4. Oil floats on the surface of water but mercury sinks. Explain why.

iv *Concentration*: The concentration of a solution is the amount of solute present in a given quantity of solvent or solution.

For many practical applications, the concentrations of solutions are expressed in molarity, molality and mole fraction.

Can you predict the base unit for concentration from those derived units of concentration?

For example, the molarity of a solution relates an amount of solute in moles (mol) and a solution volume in cubic decimetres (dm^3), or the amount of solute in moles (mol) and solution volume in litres (L).

$$\text{Concentration in molarity} = \frac{\text{number of moles of solute}}{\text{volume in litre of solution}}$$

Units of molarity: mol dm^{-3} and mol L^{-1}.

v Pressure is defined as force per unit area over which the force is exerted.

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

Thus, the SI unit of pressure is Newton per metre square (N m^{-2}). This unit is called Pascal (Pa) (in honour of Blaise Pascal who investigated the effect of pressure on fluids).

1 Pascal (1 Pa) = 1 Newton per metre square (1 N m^{-2}).

Frequently used non-SI units for expressing pressure are millimetre of mercury (mmHg), torr, and atmosphere (atm).

$$1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ torr} = 101.3 \text{ kPa}$$

1.2.2 Common Prefixes Used in SI Units

Activity 1.5

Form a group and discuss the following:

Why would it be difficult to use kilograms to express the amounts of chemicals used in large-scale industrial quantities and the amounts of chemicals used for laboratory experiments?

After the discussion, share your ideas with the rest of the class.

We use prefixes to indicate decimal multiples or fractions of the base units. The international system uses a series of prefixes to indicate decimal fractions or multiples of various units by powers of 10. All numbers can be expressed in the form of $a \times 10^b$, where 'a' is a number between 1 and 10, and the exponent 'b' is an integer. This feature makes it easy to convert from one unit to another. Some of the most commonly encountered prefixes in chemistry are listed in Table 1.2.

Table 1.2 Some common SI prefixes

Prefix	Meaning	Symbol	Multiple/Fraction
tera	trillion	T	10^{12}
giga	billion	G	10^{9}
mega	million	M	10^{6}
kilo	thousand	k	10^{3}
deci	tenths of	d	10^{-1}
centi	hundredth of	c	10^{-2}
milli	thousandth of	m	10^{-3}
micro	millionth of	μ	10^{-6}
nano	billionth of	n	10^{-9}
pico	trillionth of	p	10^{-12}

When we solve numerical problems, we use an approach to units called dimensional analysis. Dimensional analysis was developed to ensure that our answers yield proper units. It also offers a systematic approach to solve numerical problems and check our solutions for possible errors.

In dimensional analysis, we carry units through all calculations. As we work, we multiply units together, divide them by each other, and 'cancel' them.

The key to use dimensional analysis is the correct use of conversion factors in order to change one unit into another. A conversion factor is a fraction whose numerator and denominator are the same physical quantity expressed in different units. For example, 100 cm and 1 m are the same length, 100 cm = 1 m. This relationship allows us to write two conversion factors:

$$\frac{100\text{ cm}}{1\text{ m}} \text{ and } \frac{1\text{ m}}{100\text{ cm}}$$

The first of these factors is used when we want to convert metres to centimetres and second centimetres to metres. For example, the length in centimetres of an object that is 8.50 m long is given by:

$$\text{Number of centimetres} = 8.50\ \underbrace{\text{m}}_{\text{given unit}} \times \frac{\overbrace{100\text{ cm}}^{\text{desired unit}}}{\underbrace{1\text{ m}}_{\text{given unit}}} = 850\ \overbrace{\text{cm}}^{\text{desired unit}}$$

Note that the unit of metre in the denominator of the conversion factor cancels the unit of metre in the measurement given (8.50 m). The centimetre in the numerator of the conversion factor becomes the unit of the final answer. In general, the units multiply and divide as follows:

$$\cancel{\text{given}}\ \cancel{\text{unit}} \times \frac{\text{desired unit}}{\cancel{\text{given}}\ \cancel{\text{unit}}} = \text{desired unit}$$

If the desired units are not obtained in a calculation, then we know that we made an error somewhere. Careful inspection of units often reveals the source of the error.

Exercise 1.5

1. If a man has a mass of 115 pounds, what is his mass in grams (1 lb = 453.6 g)?
2. A piece of aluminium foil is 8.0×10^{-5} cm thick. What is its thickness in micrometres?
3. Convert 75 ng to mg.
4. Convert 6.75 m^3 to μL.
5. Convert each of the following measurements to a unit that replaces the power of ten by a prefix.

   a 3.22×10^{-6} s b 9.56×10^{-3} m c 1.07×10^{3} g
6. Calcualte the mass in grams of two cubic inches (2.00 in^3) of gold. Density of gold = 19.3 g cm^{-3}.

1.2.3 Uncertainty in Measurements

In scientific work, we recognize two kinds of numbers: exact numbers (whose values are known exactly) and inexact numbers (whose values have some uncertainty).

Exact numbers are those that have defined values or are integers that result from counting number of objects. For example, by definition, there are exactly 12 eggs in a dozen and exactly 1000 g in a kilogram. The number one is any conversion factor between units, as in 1m = 100 cm, of course the number one is also an exact number.

Numbers obtained by measurement are always inexact. Uncertainties always exist in measured quantities. There are many causes of uncertainty, but the most important are usually

- the person doing the measurement,
- the measuring device,
- the environment where the measurement is being made, and
- variability in the item being measured.

Making a measurement usually involves comparing the item you are measuring with a unit or a scale of units. It is often impossible to obtain the exact value of the quantity measured, unless all the numbers are exact integers.

Activity 1.6

Form a group and perform the following activities.

1. Make a chain of paper clips or other objects of uniform length. Then use a metre stick to measure a series of lengths on the chain. For example, measure sections containing one, two, three, etc., clips. Record your results and share them with your classmates.
2. Using laboratory scale, take several mass reading for one, two, three objects of uniform size. You can use any convenient objects you find in the laboratory. Record your results and discuss them in your group. Focus especially on the similarities and differences in your measurement. Did you all find the same reading for the same object? What do you think are the cause of the uncertainties, if any?

Discuss the results with the rest of the class.

1.2.4 Precision and Accuracy in Measurements

Precision and accuracy are terms which are used to express uncertainties in measurement. Precision is a measure of how clearly individual measurements agree with one another. Accuracy refers to how closely individual measurements agree with the correct or 'true' value.

Activity 1.7

Form a group, perform the following activities and discuss each of the following questions. After the discussion, share your ideas with the rest of the class.

A laboratory instructor has given a sample of amino acid powder to four students, A, B, C and D. Each student is asked to weigh the sample and record his/her results. The true (accepted) value is 8.72 g. Their results for three trials are:

Trials	Student A	Student B	Student C	Student D
1	8.72 g	8.50 g	8.50 g	8.41 g
2	8.74 g	8.77 g	8.48 g	8.72 g
3	8.70 g	8.83 g	8.51 g	8.55 g

a Calculate the average mass from each set of data, and determine which set is the most accurate.

b Which set of data is the most precise? Is this set also the most accurate?

c Which set of data is the least accurate? Is this set also the least precise?

A possible set of results obtained from a measurement of the length of a table with a metre stick by five students is given in Table 1.3.

Table 1.3 A set of measurements of length

Student	Length (m)
1	2.157
2	2.150
3	2.153
4	2.159
5	2.156
Average	2.155

The precision of a set of measurements refers to the degree of reproducibility among the set. The precision is good (or *high*) if each of the measurements is close to the average of the series. The precision is low (or *poor*) if there is a wide deviation from the average value. The precision of the data in Table 1.3 is good. Each measurement is within 0.005 m of the average value. In contrast the accuracy of a set of measurement refers to the closeness of the average of the set to the ‘‘correct’’ or ‘‘true’’ value.

Measurements of high precision are more likely to be accurate than are those of poor precision, but even highly precise measurements are sometimes inaccurate.

Example 1.2

Four persons A, B, C and D used targets a, b, c and d respectively while practicing with a rifle. The results are illustrated in Figure 1.4. How do you explain the following results?

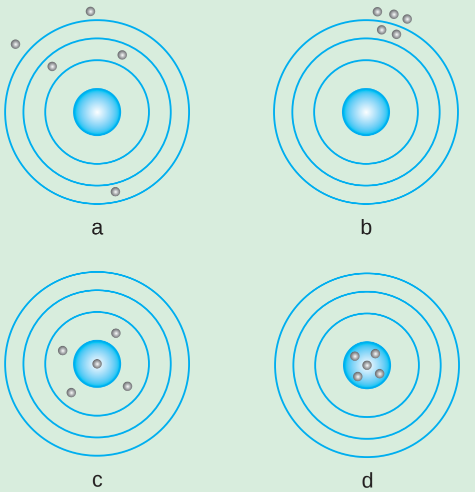
a Target a represents low accuracy and low precision.

b Target b represents low accuracy and high precision.

c Target c represents high accuracy and low precision.

d Target d represents high accuracy and high precision.

Read the figure’s legend and tally with your answers.



(The central blue region is the centre of the target)

a measurements of low accuracy and low precision are scattered and off-centre;
b those with low accuracy and high precision form a tight off-centre cluster;
c those with high accuracy and low precision are evenly distributed but are distant from the centre; and
d those with high accuracy and high precision are bunched in the centre of the target.

Figure 1.4 Comparing precision and accuracy.

Exercise 1.6

Four students measured the mass of a piece of metal whose accurate mass is 34.75 g. Their results are 34.2 g, 33.75 g, 35.0 g, and 34.69 g.

a What is the best estimate for the mass of the piece of metal? Explain why?

b Explain whether the results are precise or accurate?

The precision of a result is a measure of the certainty of the value. Usually the result is quoted as a plus-or-minus (±) value. For example, the accepted value of a universal gas constant (R) is 8.314 J mol^{-1} K^{-1}. One student might quote a result of (8.34 ± 0.03) J mol^{-1} K^{-1}, while another student gives a value of (8.513 ± 0.0006) J mol^{-1} K^{-1}. The result of the former student is more accurate (*i.e., closer to the true* (*accepted*) *value*), while that of the latter student is the more precise (*i.e., has the smallest uncertainty*).

Precision and accuracy are linked with two common types of errors called Random and Systematic errors.

Random error makes a measurement less precise but not in any particular direction. In other words, the actual value may be either greater or smaller than the value one records. Random errors arise mostly from inadequacies or limitations in the instrument. On the other hand this may be a result of how precisely someone can read a metre or a scale.

For example, consider the measuring cylinder shown in Figure 1.3c; we would probably take the reading as 70.0, but in doing so, we say that it is nearer to 70.0 than it is to 69.9 or 70.1. In other words, it is greater than 69.95 (had it been less we would have recorded it as 69.9) and smaller than 70.05 (had it been greater we would have recorded it as 70.1); hence, we should record this value as 70.0 ± 0.05.

In some cases, such as many thermometers, it is only possible to read a scale to the nearest 0.2 (that is, one would record 28.0, 28.2, 28.4, 28.6, etc., but never an odd final digit such as 28.3, 28.5, 28.7 etc.). In this case, the uncertainty would be ± 0.1, because a reading of 28.2 means it is greater than 28.0, but less than 28.4.

Systematic errors produce values that are either entirely higher or smaller than the actual value. It always affects a result in a particular direction, and skews the accuracy of the experiment in that direction. Systematic errors arise from flaws or defects in the instrument or from errors in the manner that the measurement was taken.

For example, when you are taking the initial reading of a burette that is placed well above head height, you might decide to read the top rather than the bottom of the meniscus. Systematic errors can lead to inconsistent results.

Exercise 1.7

Indicate which of the following allow you to give exact numbers when you measure them:

a The mass of a paper clip;

b The surface area of a coin;

c The number of inches in a mile;

d The number of ounce in a pound;

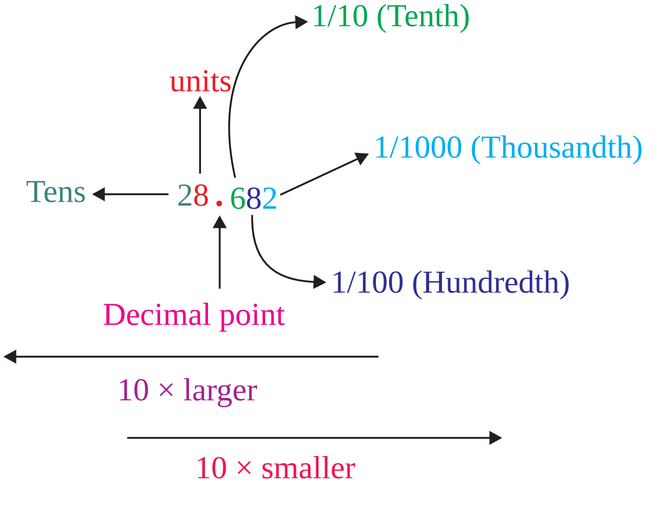
e The number of microseconds in a week;

f The number of pages in this text book.

1.2.5 Decimal Places

A decimal place refers to the number of digits to the right of the decimal point. Each successive position to the right of a decimal point has denominator increased by a power of ten. For example, 0.087 is a number given to three decimal places, and in 0.087, 0 is the first decimal place, 8 is the second, 7 is the third.

A point or dot (•) used to separate the whole number part from the fractional part of a number is called a decimal point.



To express a value to the n^{th} decimal place, look at the values of the $(n + 1)^{th}$ digit, as stated in the following rules.

1 If the $(n + 1)^{th}$ digit is 4 or less, leave the n^{th} digit unchanged. For instance, 369.648 rounds to 369.6 if we want one decimal place.

2 If $(n + 1)^{th}$ digit is greater than 5 round up the n^{th} digit. For instance, 936.758 can be rounded to 936.76 if we want to express the result upto two decimal places.

3 If the digit to be removed is 5, the preceding number increases by one unit if it is odd and remains unchanged if it is even. For example, 17.75 rounds to 17.8, but 17.65 rounds to 17.6. Note that if 5 is followed only by zeros, the left-most digit is unchanged. But if the 5 is followed by non-zeros, the final digit is increased by 1. For example, 17.6500 rounds to 17.6, but 17.6513 rounds to 17.7.

Example 1.3

1. Round 7.1284 to 2 decimal places.
2. Round 0.1284 to 1 decimal place.
3. Round 26.895 to 2 decimal places.

Solution:

1. The 3rd decimal number, 8, is bigger than 5, so we add 1 to the 2nd decimal number, 2, and drop the rest of the decimal numbers. Our answer is 7.13.
2. The 2nd decimal number, 2, is less than 5, so we do nothing to the 1 and we drop the rest of the decimal numbers. Our answer is 0.1.
3. The second decimal number, 9, is odd, so we add 1 to 9 to get 10, and drop the rest of the decimal numbers. But, we have to carry the 1 to the 8 to get 9. So our answer is 26.90 or just 26.9.

1.2.6 Significant Figures

Significant figures are those digits that correctly indicate the precision of a measurement. So significant figures show both the limits of accuracy and where the uncertainty begins. For this reason, it is important to indicate the margin of error in measurement by clearly indicating the number of significant figures, which are the meaningful digits in a measured or calculated quantity.

It may be noted that measurements that were taken by the five students in Table 1.3 agree on the first three digits (2.15); they differ only in the fourth digit. The last digit in a scientific measurement is usually regarded as uncertain. All digits known with certainty, plus one of uncertain value, are called significant figures.

The measurements in Table 1.3 have four significant figures. In other words, we are quite sure that the length of the table is between 2.15 m and 2.16 m. Our best estimate, including the uncertain digit, is the average value, i.e., 2.155 m.

The last digit in a significant figure is uncertain because it reflects the limit of accuracy.

The following guidelines apply to determining the number of significant figures in a measured quantity. It has to be decided whether zeros are significant in three different situations.

1. If the zeros precede the first non-zero digit, they are not significant. Such zeros merely locate the decimal point. *i.e.*, they define the magnitude of measurement. For example, 0.004 m has one significant figure, and 0.00016 m has two significant figures.
2. If the zeros are between non-zero digits, they are significant. For example, 204408 kg has six significant figures while 0.05504 has four significant figures.
3. If the zeros follow non-zero digits, there is ambiguity if no decimal point is given. For example, if a volume is given as 200 cm^3, there is no way of expressing if the final two zeros are significant. But if the volume is given as 200 cm^3, zeros after a non-zero digit preceded by a decimal point make all figures significant. Thus, 200.cm^3 has three significant figures. If it is given as 200.0 cm^3, it has four significant figures.
4. Non-zero digits are always significant.

Example 1.4

1. How many significant figures are there in:

   a 0.0004802 b 6, 834 c 5, 2100

2. A calculator display shows the result of a calculation to be 67340.468. How many significant figures are there?

Solution:

1. a four significant figures

   b four significant figures

   c ambiguous

2. Eight significant figures.

1.2.7 Scientific Notation

The ambiguity in significant figures can be avoided by expressing the measurements in scientific notation. Scientific notation is a way of expressing large or small numbers as factors of the powers of 10. The exponents of 10 can be used to make the expression of scientific measurements more compact, easier to understand, and simpler to manipulate. For example, the mass of an electron is 0.000 000 000 000 000 000 000 000 000 91 g, and the value of Avogadro's number is 602, 000, 000, 000, 000, 000, 000, 000 mol^{-1}. In scientific notation, these values can be expressed as 9.1×10^{-28} g and 6.02×10^{23} mol^{-1} respectively.

To express numbers in scientific notation, you use the form $a \times 10^{b}$, where a is a decimal number between 1 and 10 (*but not equal to 10*), and is known as the digit term, and b is a positive or negative integer or zero and is called the exponent.

To express a number in scientific notation, count the number of places you must move the decimal point in order to get '*a*' between 1 and 10. Moving the decimal point to the right (*if the number is less than 1*) indicates a negative exponent, and moving the decimal point to the left (*if the number is greater than 1*) indicates a positive exponent.

Example 1.5

Express the following numbers in scientific notation (each with three significant figures)

a 7500000 b 0.000777

Solution:

a 7.50×10^{6} b 7.77×10^{-4}

Exercise 1.8

1. Express 0.0000000013 in scientific notation.
2. Express each of the following with the number of significant figures indicated:
   a 5,000.083 (to three significant figures)
   b 3,986.0 (to four significant figures)

Follow the following rules for proper answers of the results of substraction, addition, multiplication, and division of numerical values.

1. For addition and subtraction, the answer should contain no more digits to the right of the decimal point than any individual quantity. i.e., use the least number of decimal places.
2. For multiplication and division, a result can only be as accurate as the factor with the least number of significant figures that goes into its calculation. i.e., use the least number of significant figures.

Example 1.6

1. What is the area, in square metres, of a room that is 12.42 m long and 4.81 m wide?
2. Perform the following calculation and round off the answer to the correct number of digits.

$$49.146 + 72.13 - 9.1434 = ?$$

Solution:

1. The length of the room is expressed to four significant figures and the width to three. By whatever method we use to carry out the multiplication, we are limited to three significant figures in our answer.

$$12.42 \text{ m} \times 4.81 \text{ m} = 59.7 \text{ m}^2$$

2. In this calculation, we must add two numbers and, from their sum, subtract the third. We express the answer to two decimal places, the number of decimal places in '72.13'. We do this in two ways below:

a

$$\begin{array}{r} 49.146 \\ +\,72.13 \\ \hline 121.276 \end{array} = \begin{array}{r} 121.28 \\ -9.1434 \\ \hline 112.1366 \end{array} = 112.14$$

b

$$\begin{array}{r} 49.146 \\ +72.13 \\ \hline 121.276 \\ -\,9.1434 \\ \hline 112.1326 \end{array} = 112.13$$

The preferred method is b, where we do not round off the intermediate result: 121.276.

Exercise 1.9

1. Perform the indicated operations and give answers with the proper number of digits.
   - a   451 g – 15.46 g – 20.3 g
   - b   15.436 L + 5.3 L – 6.24 L – 8.177 L
   - c   48.2 m + 3.82 m + 48.4394 m
   - d   148 g + 2.39 g + 0.0124 g
   - e   37 m × 2.340 m × 0.52 m
   - f   62.89 m ÷ 4.7 m
2. The distance between carbon atoms in a diamond is 154 pm. Convert this distance to millimetres.
3. In a certain part of the country, there is an average of 710 people per square mile and 0.72 internet services per person. What is the average number of internet services in an area of 5.0 km^2?

1.3 CHEMISTRY AS EXPERIMENTAL SCIENCE

At the end of this section, you should be able to:

- define scientific method;
- describe the major steps of the scientific method;
- use scientific methods in solving problems;
- demonstrate some experimental skills in chemistry; and
- describe the procedures of writing laboratory reports.

1.3.1 The Scientific Method

Activity 1.8

Form a group and perform the following activity.:

- a   Collect a pastic bag filled with different items provided by your teacher.
- b   Decide on the question you would like to answer about your bag. Write it down. (Do not open the bag)
- c   Guess what the answer to your question might be. Write down. (Do not open the bag)
- d   Open your bag and answer the questions.

e Be sure to count the total number of items.

Now, discuss which part of the activity (a, b, c, d, or e) introduces the scientific terminology: hypothesis, data collection, experimentation, etc.

Discuss the results with the rest of the class.

Science is an organized body of knowledge that is based on a method of looking at the world. The scientific methods are unique, and require any explanation of what is seen to be based on the results of experiments and observations. These experiments and observations must be verifiable by anyone who has the time and means needed to reproduce them.

Although two different scientists rarely approach the same problem in exactly the same way, there are guidelines for the practice of science that have come to be known as the scientific method. These guidelines are outlined in Figure 1.5.

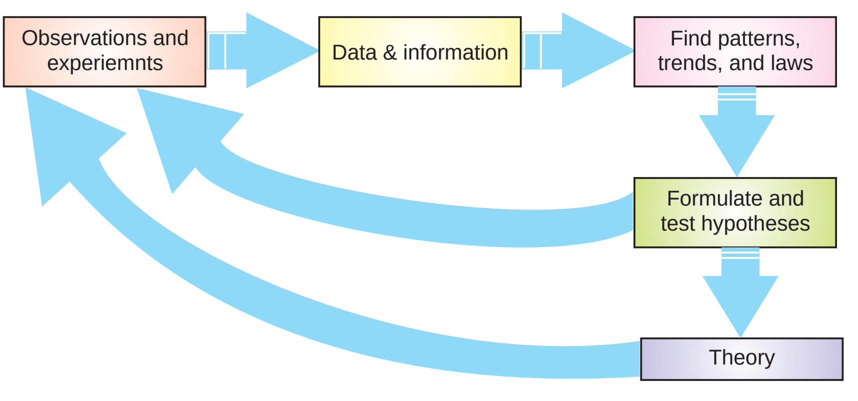


Figure1.5 Steps in scientific methods.

The scientific method is a general approach to problems. It involves making observations, collecting information/data, seeking patterns in the observations, formulating hypothesis to explain the observations, and testing these hypothesis by further experiments. If a hypothesis successfully posses many tests, it becomes a theory, which is a tested explanation of a hypothesis.

Activity 1.9

Form a group and imagine/enact that you are a group of scientists who have discovered a new chemical compound of great use in curing cancer.

1. How would you present your work to other scientists of the world?
2. Is there a particular format to write scientific reports?

   Discuss the essential features of a well-designed experiment.
3. Consider the following:
   - a Strike a match stick to light it.
   - b Record all your observations in writing.
   - c Examine your written observations and consider their objectivity.

   Which of these do you think are just descriptions of your observations?

   In contrast, which are the ideas that you formed based on your observations?

   Finally, which of your data could be subjective or partly subjective? Why?

Present your observations to your classmates. As you do so, describe the objectivity or subjectivity of each statement that you make. In particular, discuss those that you inferred based on observation.

Scientists seek general relations that unify their observations. A concise verbal statement or a mathematical equation that summarizes a broad variety of observations and experience is known as scientific law. A familiar example is the law of gravity. It summarizes the experience that what goes up must come down.

Scientists also seek to understand laws. A tentative explanation of a law is called a hypothesis. A hypothesis is useful if it can be used to make predictions that can be tested by further experiments and can thereby be verified.

A hypothesis that continually withstands such tests is called a theory. A theory is an explanation of the general principles of certain phenomena that has considerable evidence or facts to support it. It may serve to unify a broad area and may provide a basis for explaining many laws.

There is no fool-proof, step-by-step scientific method that people use. Their approaches depend on their temperaments, circumstances and training. Rarely will two people approach the same problem in the same way. Scientific progress is not smooth, certain and predictable. The path of any scientific study is likely to be irregular and uncertain. Progress is often slow, and many promising leads turn out to be dead ends. Serendipity (*fortunate accidental discovery*) as well as perseverance has played an important role in the development of science.

1.3.2 Some Experimental Skills in Chemistry

Activity 1.10

Form a group and perform the following activities. Discuss the results with the rest of the class.

Consider the reaction between a copper wire and concentrated nitric acid. Observe this reaction and suggest possible solutions for the following questions:

Has the copper wire disappeared?

Have the copper atoms disappeared? If not disappeared, where are they?

What are the expected products?

Is this change physical or chemical?

Caution: The gaseous product formed is pungent, irritating and poisonous; do not inhale it and do not allow it to escape in the air.

Experimental Skills

Chemistry has two main roles in the curriculum. Chemistry is pre-requisite for many other courses in higher education, such as medicine, biological and environmental sciences. It is an experimental science that combines academic study with the acquisition of practical and investigatory skills in order to:

- plan experimental activity, i.e., planning;
- carry out experimental work, i.e., implementing;
- analyse and draw conclusions from the results of experimental work, i.e., analysing evidence and drawing conclusions; and
- evaluate the work, i.e., evaluating evidence and procedures.

To acquire these experimental skills and investigations you should be able to:

- follow a sequence of instructions;
- use techniques, apparatus and materials;
- make and record observations, measurements and estimates;
- interpret, evaluate and report upon observations and experimental results;
- design/plan an investigation, select techniques, apparatus and materials; and
- evaluate methods and suggest possible improvements.

It is not possible to prepare an exhaustive list of skills, but the major skills that are ideally developed in a laboratory environment include:

1. Skills in the safe handling of chemical materials, taking into account their physical and chemical properties, including any specific hazards associated with their use.
2. Skills required for conducting the standard laboratory procedures involved in synthetic and analytical work, in relation to both inorganic and organic systems.
3. Skills in monitoring, by observation and measurement, of chemical properties, events or changes, and the systematic and reliable recording and documentation thereof.
4. Competence in planning, design and execution of practical investigations, from the problem-recognition stage to the evaluation and appraisal of results and findings; this includes the ability to select appropriate techniques and procedures.
5. Skills in the operation of standard chemical instrumentation such as that used for structural investigations and separation.
6. Ability to interpret data derived from laboratory observations and measurements in terms of their significance and the theory underlying them.
7. Ability to conduct risk assessments concerning the use of chemical substances and laboratory procedures.

Chemistry Laboratory Apparatus

Laboratory equipment comprises different sets of apparatus, which are designed to perform various tasks in the laboratory. On the basis of their use, these apparatus can be broadly classified into three categories:

1. Reaction vessels, *e.g.*, Beakers, flasks, boiling tubes and test tubes.
2. Measuring equipments, *e.g.*, Pippetes, burettes, balances and thermometers.
3. Support and heating devices, *e.g*. Stand and clamp, tripod and gauze, spirit burner and Bunsen burner.

The practical activities are intended to support conceptual development. Proficiency in handling of apparatus is the result of continual practice.

Note: The information about care and safety associated with the use of some of these apparatus/devices have been discussed at various places in this unit. However, adequate information about these is obtained when these are actually used to perform experiments in the laboratory or elsewhere.

Chemistry Laboratory Safety Rules

The chemistry laboratory may be considered as a place of discovery and learning. However, by the very nature of laboratory work, it can be a place of danger if proper common-sense precautions are not taken. It is your duty in law to take reasonable care for your own health and safety and that of others working in the laboratory. Therefore, it is essential that the students are taught what can go wrong, how to prevent such events from occurring, and what to do in case of an emergency.

Protect your eyes

- Appropriate eye protection must be worn at all times! Inform your teacher if you wear contact lenses.

Wear appropriate protective clothing

- Your clothing should cover your legs to the knees; shorts are not appropriate for the laboratory. Loose clothing should not be worn because it may dip into chemicals or fall into a flame and catch fire. Further, laboratory aprons can be used to protect your clothing.

Wear shoes that cover your feet

- Due to the dangers of broken glass and corrosive liquid spills in the laboratory, open sandals or bare feet are not permitted in the laboratory. Remember! leather shoes protect your feet from chemical spills – canvas shoes do not.

Tie back loose hair

- Dangling hair can fall into the Bunsen burner and catch fire or can fall into a chemical solution

Eating and drinking in the laboratory

- Do not taste any chemical! Even food, drink and chewing gum are prohibited in the chemistry laboratory. These activities are ways by which you can accidentally ingest harmful chemicals

Smelling chemicals

- Do not smell any chemicals directly!
- Smell chemicals only if your teacher specifically tells you to do so, then use your hand to fan the vapour towards your nose.

Pipetteing out solutions

- Do not suck the solutions in the pipette by mouth!
- Use a rubber suction bulb (pipette bulb) or other device to fill a pipette.

General precautions

- Wash your hands with soap and water before leaving the laboratory even if you have been wearing gloves.
- Know the hazards of the materials being used.
- When lighting the Bunsen burner, first light the match stick then turn on the gas.
- Know how to interpret data from a MSDS (Material Safety Data Sheets).
- Read the labels on the reagent bottles carefully to make sure that you are using the right chemical.

- Never add water to concentrated acid solutions. The heat generated may cause spattering. Instead, as you stir, add the acid slowly to water.
- Hold your hand over the label while pouring.
- For minor skin burns, immediately plunge the burned portion into cold water and inform the teacher.
- If you get any chemical in your eye, immediately wash the eye with the eye-wash fountain and notify the teacher.
- Work with volatile chemicals under a fume hood.
- Never look directly into a test tube. View the contents from the side.
- Get acquainted with the location and proper usage of the safety equipments like eye wash fountain, safety shower, fire extinguisher, emergency exits.
- Carry out only the experiments assigned by your teacher.
- Use equipment only as directed.
- Never place chemicals directly on the pan balances.
- Use glycerin when inserting glass tubing into rubber stoppers.
- Be cautious of glassware that has been heated. Handle hot glassware with gloves or beaker tongs.
- Add boiling chips to liquid to be boiled.
- Point test tubes that are being heated away from you and others.
- Check glassware for stars or cracks.
- Never use laboratory glassware for eating or drinking purposes.
- Never remove chemicals from the laboratory.
- Never work alone in the laboratory. In case of a problem, you may need another person to prevent injury or even save your life!

Demonstrate safe behaviour

- Obey all safety instructions given by your teacher or found in you experimental procedure.

- Clean up spills immediately if you know. If you are uncertain how to clean up a spill or if a large spill occurs, notify your teacher immediately.
- Before leaving the lab be sure to replace the lids to all containers, return equipment and chemicals to their proper places and clean up your work area.
- Know how to dispose off waste. Dispose off all waste materials according to your instructional procedure or your teacher's instructions.
- Remember that the lab is a place for serious work! Careless behavior may endanger yourself and others and will not be tolerated!
- Know how to respond to an emergency.
- Report any accidents or unsafe conditions immediately!
- For some experiments, it may be helpful to anticipate data. You should always read the experiment in advance.

Note: Additional safety precautions will be announced in class prior to experiments where a potential danger exists.

1.3.3 Writing a Laboratory Report

The purpose of writing an introductory laboratory experiment is to give practice in writing laboratory reports that answer the general questions:

- What did you do?
- Why did you do it?
- How did you do it?
- What happened?

A laboratory report is a written composition of the results of an experiment. It should be written precisely and clearly, using good grammar and punctuation. Each report must include: title, objective, materials and (equipment) used, procedure, observation, result, discussion, and conclusion.

1. *Title*: Create a title in less than ten words that reflects the factual content of your report

2. *Objective*: This section states the purpose of your experiment. Be specific about the outcomes that you plan to achieve when you designed your experiment.
3. *Materials used*: Describe the substances, equipment and instrumentation that is to be used in your work. Copy the format for this section from your laboratory manual or from the standard procedure supplied by the teacher.
4. *Procedure*: Describe how you performed the experiment, and mention each step in chronological order.
5. *Data/Observations*: This section demonstrates that you carried out an experiment carefully and knowledgeably. The person reading your report should find it clear and convincing enough to take your experimental results seriously.
6. *Result and Discussion*: In this section of the report, present your results and discuss them. Also report possible errors in the procedure and results, including possible inaccuracies.

   Include any problems that you encountered during your work. Present them objectively. If possible suggest ways in which such problems could be reduced at least if not overcome.
7. *Conclusion*: This section should be brief, as it refers back to the objectives and considers how and to what degree they have been met. Review the purpose of the experiment, and summarize the implications of the results.

Unit Summary

- The science of chemistry deals with the composition, physical properties, and chemical properties of matter.
- Matter is made up of atoms and molecules.
- All matter fits into two categories: substances and mixtures.
- The SI system has seven base units, six of which are used in chemistry.
- Some measurements are expressed directly in terms of base units as well as multiples or submultiples of a base unit. For example, you might express a length in metres as well as in kilometres or millimetres.

- A measured quantity must be expressed with the proper number of significant figures to indicate its precision.
- In reporting calculated quantities special attention must be paid to the concept of significant figures.
- Calculations can be done by the unit-conversion method.
- Techniques of estimating answers are also useful in problem solving.
- The scientific method involves making observations, doing experiments forming hypothesis, gathering data, and formulating laws and theories.

Check List

Key terms of the unit

- *Accuracy*
- *Analytical chemistry*
- *Basic unit*
- *Derived unit*
- *Extensive*
- *Hypothesis*
- *Inorganic chemistry*
- *Intensive*
- *Law*
- *Organic chemistry*
- *Physical chemistry*
- *Precision*
- *Scientific method*
- *Scientific notation*
- *Significant figure*
- *Theory*
- *Uncertainty*

Review Exercise

Part I: Multiple Choice Type Questions

1. Which of the following numbers has five significant figures?

   a 61,530 b 0.6153 c 0.006154 d 615.40

2. What is the mass of 30.0 mL of a liquid that has a density of 1.60 g/mL?

   a 18.8 g b 48.0 g c 31.6 g d 53.3 g

3. A blood sample analysis showed a total of 0.00361 g lead. Express it in micrograms?

   a 0.361 µg b 3.61×10^3 µg c 3.61×10^6 µg d 3.61×10^4 µg

4. Which of the following statements is correct?

   a One metre has about the same length as a yard.

   b A scientific notation for 0.00035 is 3.5×10^{-3}.

   c A substance with a density of 0.80 g/cm^3 sinks in water.

   d The prefix mega- means multiplied by 1,000,000.

5. A conversion factor, like the definition 100 cm = 1 m, is:

   a approximate c cannot be predicted

   b exact d subject to individual opinion

6. Which of the following is an extensive physical property?

   a colour c mass

   b density d melting point

7. The product of 12.76 cm and 0.0030 cm has __________ significant figures.

   a 2 b 3 c 4 d 5

8. A measurement gave a mass of 0.45 kg. This is the same as:

   a 4.5 g b 45 g c 450 g d 4,500 g

9. Which of the following statements is not correct?

   a Density has no units.

   b Every measurement has a unit tied to it.

   c Physical quantities are properties that can be measured.

   d The kelvin degree is larger than the celsius degree.

10. The number 52.415 ± 0.0001 has ___________ significant figures.

   a 4 b 5 c 6 d 7

Part II: Answer the following questions.

1. Convert – 40°C to °F.
2. Round off 45.68 metres to one decimal place.
3. How many seconds are there in 8 minutes?
4. Change each of the following measurements to one in which the unit has an appropriate SI prefix.

a 3.76×10^3 m

b 6.34×10^{-6} g

c 1.09×10^{-9} g

5. How many significant figures are there in each of the following measured quantities?

a 4.200×10^5 s   c 6.02×10^{23} e 0.00075 m

b 0.1050 °C   d 0.049300 g   f 8008 m

6. Express each of the following measured quantities in exponential notation. Assume all the zeros in parts c and d, are significant.

a 0.00090 cm   b 20.00 s   c 9,000 s   d 2,800 m

7. Perform the indicated operations and give answers with the proper number of significant figures.

a 48.2 m + 3.82 m + 48.4394 m

b 451 g – 15.46 g

c 15.44 mL – 9.1 mL + 105 mL

d $73.0 \times 1.340 \times 0.41$

e $\dfrac{22.61 \times 0.0587}{135 \times 28}$

8. A 25.0 mL sample of liquid bromine weighs 78.0 g. Calculate the density of the bromine.

9. Some metal chips with a total volume of 3.29 cm^3 are placed on a piece of paper and weighed. The combined mass is found to be 18.43 g, and the paper itself weighs 1.2140 g. Calculate the density of the metal to the proper number of significant figures.

10. A rectangular block of lead is 1.20 cm × 2.41 cm × 1.80 cm, and it has a mass of 59.01 g. Calculate the density of lead.

11. A block of lead, with dimensions 2.0 dm × 8.0 cm × 35 mm, has a mass of 6.356 kg. Calculate the density of lead in g cm^{-3}.

12. Demonstrate that kg L^{-1} and g cm^{-3} are equivalent units of density.

13. Steam is sometimes used to melt ice. Is the resulting change physical or chemical?

14. Which of the following lengths is the shortest and which is the longest? 1583 cm, 0.0128 km, 17931 mm, and 14 m

15. Which SI unit can be used for expressing the height of your classroom ceiling?