

CH 110 Topic 1 and 2 Notes – INTRODUCTION TO GENERAL CHEMISTRY

1.0 INTRODUCTION

Chemistry is an active and evolving science subject that is important in society. Its study started along time ago and still goes on. We will begin its study at the macroscopic level, that is, a level where we see and measure materials which make up our world. At university level, learning of chemistry starts with a course in General Chemistry. At CBU, General Chemistry or freshmen chemistry courses are coded FO 130 and CH 110 for School of Natural Resources (SNR) students and non-SNR students, respectively. The things we will learn in the first topic of this general chemistry course after the hints of studying general chemistry are given in the course outline.

1.1 THE STUDY OF CBU GENERAL CHEMISTRY

Compared with other subjects, chemistry is commonly believed to be more difficult, at least, at the introductory level. This is mostly because of its specialized vocabulary. Once one masters the vocabulary, things should become straight forward.

1.1.1 Tips on Studying General Chemistry

- (a) Use the glossary in chemistry introductory textbooks to get concise definitions of terms arising in your course. The glossary also specifies the section of the textbook to read and use to expand your class notes on particular terms/concepts
- (b) It is to be noted that relevant sections will have examples of how to solve chemistry problems. You will need to carefully go through such problems or those given during lectures.
- (c) Additionally, for each topic covered in lectures, a tutorial consisting of questions/problems will be given. It is helpful to solve these problems individually in order to check your mastery/understanding of material you have been taught. We encourage you to be able to solve tutorial problems individually because nobody will write tests of the exam for you!!!!
- (d) To help your prepare for tests and final examination, it is helpful to get past papers, answer and check correctness of your answers to past paper questions well before your tests or exam. Inability to solve tutorial and past test/exam questions is a sign of poor preparation for tests or the exam.
- (e) University education is fast paced compared to secondary school education. In order to keep abreast, a lot of students may find the concept of “today’s portion, today itself” to be helpful. This phrase is a mnemonic that advises a science student that what one learns today should be studied today if possible. This allows one to ask for help on concepts that one has not understood before the instructor moves on to the next topic that may use previous lessons as its foundational material. This approach to studying General Chemistry also allows you to devote time to understanding a topic well before a test or an exam.

1.1.2 Purpose of the Course

The course is to make you think like a chemist, that is, to look at the macroscopic world consisting of things we see, touch and measure directly and visualize the particles and events of the microscopic world that we cannot experience without modern technology and our imagination.

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It may be confusing to students when an instructor or a textbook seems to be continually shifting back and forth between the macroscopic and the microscopic world. One should just remember that data for chemical investigations often comes from observations of large-scale phenomena (or macroscopic world), but the explanations frequently lie in the unseen and partially imagined microscopic world of atoms and molecules (or the microscopic world). In other words, chemists often see one thing (in the macroscopic world) and think another (in the microscopic world).

For example, when looking at a rusted car, roof or door frame, a chemist might think about the basic properties of individual atoms of iron and how these units interact with other atoms and molecules to produce the observed change.

1.2 THE SCIENTIFIC METHOD

This is a systematic approach to research that has 5 steps, namely;

(a) Careful definition of the problem

(b) Performing experiments which consists of

- (i) Doing the experiments
- (ii) Making careful observations and
- (iii) Recording information or data about the system (system is the part of the universe that is under investigation).

Data obtained can be both qualitative and quantitative. Qualitative data consists of general observations about the system. Quantitative data comprises numbers obtained by various measurements of the system.

Scientists use standardized symbols and equations in recording measurements and observations. This representation simplifies the process of keeping records and provides a common basis for communication with other chemists and scientists.

- (c) Interpretation of results.** Once experiments are completed and data recorded scientists attempt to explain observed phenomena. Based on the data gathered, the researcher formulates a **HYPOTHESIS**, that is, a tentative explanation for a set of observations. Further experiments are devised/done to test validity (truthfulness) of the hypothesis in as many ways as possible and the process begins a new.
- (d)** After collection of lots of data, it is often desirable to **summarize information in a concise way as a law**. In science, a **LAW, is a concise verbal or mathematical statement of a relationship between phenomena that is always the same under the same conditions**. For example, Newton's Second Law of motion says that force equals mass times acceleration ($F=ma$). This means increase in mass or acceleration of a body will proportionately increase force, conversely, a decrease in mass or acceleration will always decrease the force.
- (e)** Hypotheses that survives many experimental validation tests evolve into theories. **A THEORY is a unifying principle that explains a body of facts and/or those laws that are based on them**.

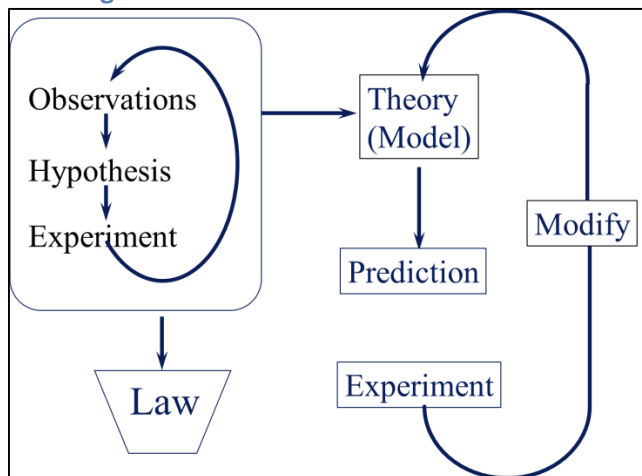
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Theories too, are constantly being tested. A disproved theory is discarded or modified to be consistent with experimental observations. Proving or disproving a theory can take many years or centuries. Sometimes this is because technology to do the relevant experiments may not be available. A good example is the ATOMIC THEORY. It took more than 2000 years to work out this fundamental principle of chemistry proposed by the ancient Greek philosopher called Democritus. A more contemporary example is the Big Bang Theory of the origin of the universe.

As shown in Figure 1.1, scientific progress is seldom straight forward. Sometimes a law precedes a theory while sometimes it is the other way round. Two scientists starting work on a project with the same objective may take different approaches.

Great discoveries are usually the result of the cumulative contributions and experience of many workers though credit for formulating a theory or a law is usually given to only one individual.

Figure 1.1: The Scientific Method



1.3 CLASSIFICATION OF MATTER

One simple definition of chemistry is that it is the study of matter and the changes it undergoes. Matter is anything that occupies space and has mass. Matter includes visible and touchable things (water, earth and trees) as well as invisible things (air). Chemists distinguish among several subcategories of matter based on composition and properties. The organization/classification of matter is shown in Figure 1.2 or Figure 1.3 below.

Figure 1.2: Classification of Matter (Ref: Zumdahl and Zumdahl)

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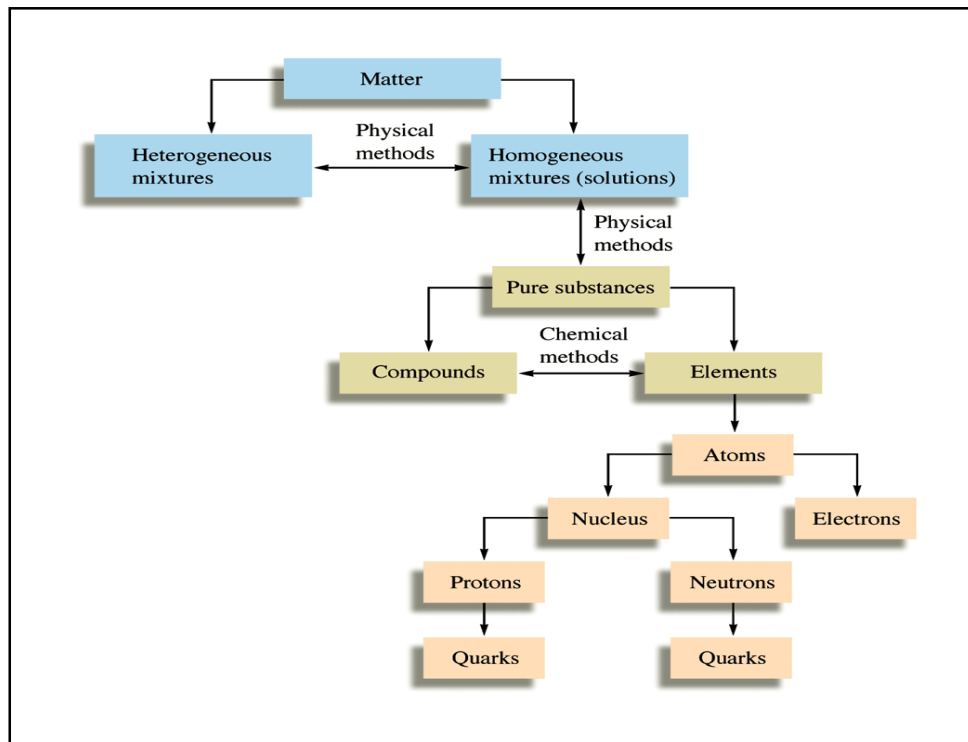
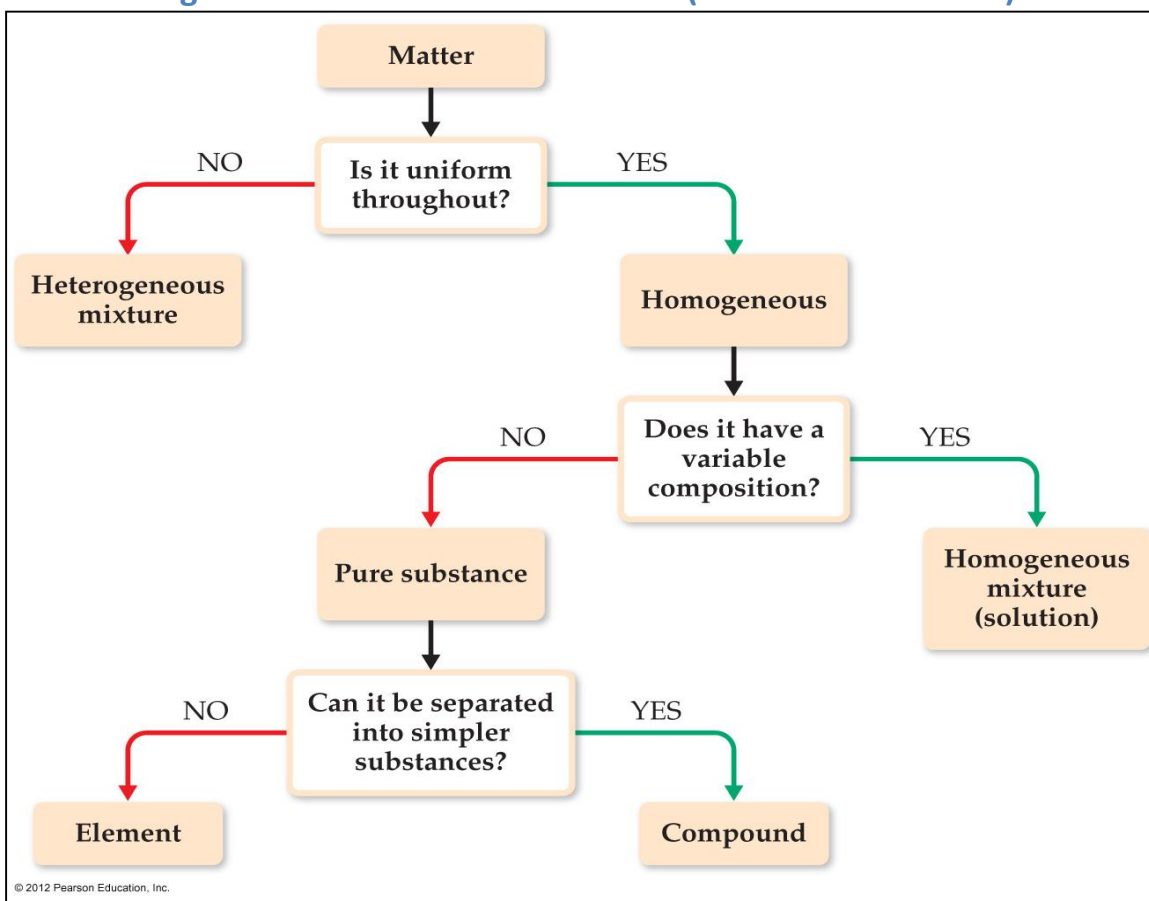


Figure 1.3: Classification of Matter (Ref: Central Science)



The classification consists of mixtures and pure substances that may be compounds or elements. Compounds have molecules or formula units made of at least two elements. Using chemical methods compounds can be separated into their constituent elements. Elements have atoms. Each atom consists of particles namely electrons and nucleons (protons and neutrons). The nucleons make up the atomic nucleus and are made up of subatomic particles called quarks.

1.3.1 Substances and Mixtures

A **substance** is a form of matter that has a definite (constant) composition and distinct properties like water, ammonia, table sugar (sucrose), gold or oxygen. A substance cannot be separated into other kinds of matter by physical processes. Substances differ in composition and are identifiable by appearance, smell taste and other properties

A **mixture** is a combination of two or more substances in which substances retain their distinct identities. For example air, soft drinks, milk and cement. Mixtures do not have constant composition. They are either homogeneous or heterogeneous.

A **homogeneous mixture** is one in which composition is the same throughout. For example, sugar dissolved in water. A homogeneous mixture is a solution.

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A **heterogeneous mixture** is one in which the composition is not uniform. For example, sand mixed with iron filings in which sand grains and iron filings remain separate.

Mixtures can be separated by physical methods. For example, sugar and water can be separated by evaporation to dryness while iron filings and sand can be separated by magnetic removal of iron filings. After separation, the components of the mixture will have the same composition and properties.

1.3.2 Substances and Compounds

Substances can either be elements or compounds. An **element** is a substance that cannot be separated into simpler substances by chemical means. To date there are 115 elements 83 of which occur naturally on earth and 32 have been created by scientists via nuclear processes.

Chemists use symbols of one or two letters to represent elements. The first letter is always capitalized but any following letters are not. For example Co is a symbol for the element cobalt whereas CO is the formula for carbon monoxide. Symbols of some elements are derived from Latin names. For example, Au is from aurum (gold), Fe is from ferrum (iron) and Na is from natrium (sodium). See also Appendix 1 of Chemistry by Raymond Chang 10th Edition.

Atoms of most elements can interact to form compounds. A **compound** is a substance composed of two or more elements chemically united in fixed proportions. Compounds unlike mixtures can be separated only by chemical means into their pure components.

1.3.3 The Three States of Matter

There are three states of matter. These are solid, liquid and gas.

1.3.4 Physical and Chemical Properties of Matter

A **physical property** can be measured and observed without changing the composition of a substance. E.g. boiling point, melting point, etc.

A **chemical property** is observed if there is a chemical change. An example is the burning of hydrogen to form water from which we cannot recover hydrogen by physical processes.

We measure properties of matter in one of the two categories, namely,

- (a) Extensive properties or
- (b) Intensive properties

An **extensive property** depends on how much matter is being considered. Mass, the quantity of matter in a given sample of substance is an extensive property. More matter means more mass. Values of extensive properties can be added together. Volume of a substance is another extensive property.

An **intensive property** does not depend on how much matter is being considered. Examples of intensive properties are density (mass divided by its volume) and temperature. Intensive properties like temperature are not additive as can be seen from the fact that the temperature of water in two beakers(T_1 and T_2) is not summed to determine the temperature of water when the contents of the two beakers are combined.

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1.4 MEASUREMENT

Chemists measure properties of substances using different instruments. These measurements are often used in calculations to obtain other related quantities. Examples of simple instruments for measuring macroscopic properties (that is, properties which can be measured directly) likely to be used in a general chemistry course are shown in the table below.

Table 1.1: Examples of simple instruments for general chemistry measurements

Instrument	Measured property in common units
Burette, pipette, graduated cylinder, volumetric flask	Volume in millilitres (mL)
Balance	Mass in grams (g)
Stopwatch	Time in seconds (s)
Thermometer	Temperature in Kelvin(K) or degrees Celsius (°C)

Microscopic properties on the atomic or molecular scale are determined by indirect methods. The measurement result in both macroscopic and microscopic a number with an appropriate unit.

1.5.1SI Units

Chemists use metric units that are related decimally, that is, by powers of 10. The metric system is used by present day chemists and has units consistent with the International System of Units (SI) and these will be used in this course. The SI base units and their accepted symbols are the metre (m) for length, the kilogram (kg) for mass, the second (s) for time, the Kelvin (K) for thermodynamic temperature, the mole (mol) for the amount of substance and the candela (cd) for luminous intensity. The Table 1.2 below shows the seven SI base units of measurements.

Table 1.2: Table of SI Base Units

Base Quantity	Name of Unit	Symbol
1. Length	metre	m
2. Mass	kilogram	kg
3. Time	second	s
4. Electrical current	ampere	A
5. Temperature	Kelvin	K
6. Amount of substance	mole	mol
7. Luminous intensity	candela	cd

All other units of measurements are derived from base units. Examples of common derived units, used in this course are given in Table 1.3. Additional derived units can be found in the IUPAC's Physical Chemistry Division green book by Mills et. al. (1993) entitled "Quantities, Units and Symbols in Physical Chemistry" from the website

http://old.iupac.org/publications/books/gbook/green_book_2ed.pdf accessed on 1 July 2015.

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Table 1.3: Table of Examples of Common Derived Units

Quantity	Name	Symbol	SI base units	Derived Units
area			m^2	
volume			m^3	
velocity			m s^{-1}	
density			kg m^3	
force	Newton	N	m kg s^{-2}	
pressure or force/area	Pascal	Pa	$\text{kg m}^{-1} \text{s}^{-2}$	N m^{-2}
energy	joule	J	$\text{m}^2 \text{kg s}^{-2}$	N m
power	watt	W	$\text{m}^2 \text{kg s}^{-3}$	J s^{-1}
Specific heat			$\text{m}^2 \text{s}^{-2} \text{K}^{-1}$	$\text{J kg}^{-1} \text{K}^{-1}$

To make the numbers used with either base units or derived convenient for the metric or SI system, **prefixes** are attached indicating decimal multiples of the units. Accepted prefixes, symbols, and multiples are shown in Table 1.4. The use of prefix steps smaller than 10^3 is discouraged.

Table 1.4: Accepted SI Prefixes and Symbols for Multiples and Submultiples of Units

Prefix	Symbol	Meaning	Example
yotta	Y	$1\,000\,000\,000\,000\,000\,000\,000\,000 = 10^{24}$	1 yottametre (Ym) = 1×10^{24} m
zetta	Z	$1\,000\,000\,000\,000\,000\,000\,000\,000 = 10^{21}$	1 zettametre (Zm) = 1×10^{21} m
exa	E	$1\,000\,000\,000\,000\,000\,000\,000\,000 = 10^{18}$	1 exametre (Em) = 1×10^{18} m
peta	P	$1\,000\,000\,000\,000\,000\,000\,000\,000 = 10^{15}$	1 petametre (Pm) = 1×10^{15} m
tera	T	$1\,000\,000\,000\,000\,000\,000\,000\,000 = 10^{12}$	1 terametre (Tm) = 1×10^{12} m
giga	G	$1\,000\,000\,000\,000\,000\,000\,000\,000 = 10^9$	1 gigametre (Gm) = 1×10^9 m
mega	M	$1\,000\,000\,000\,000\,000\,000\,000\,000 = 10^6$	1 megametre (Mm) = 1×10^6 m
kilo	k	$1\,000 = 10^3$	1 kilometre (km) = 1×10^3 m
hecto	h	$100 = 10^2$	1 hectometre (hm) = 1×10^2 m
deka	da	$10 = 10^1$	1 dekametre (dam) = 1×10^1 m
-	-	$1 = 10^0$	1 metre (m) = 10^0 m
deci	d	$0.1 = 10^{-1}$	1 decimetre (dm) = 1×10^{-1} m
centi	c	$0.01 = 10^{-2}$	1 centimetre (cm) = 1×10^{-2} m
milli	m	$0.001 = 10^{-3}$	1 millimetre (mm) = 1×10^{-3} m
micro	μ	$0.000\,001 = 10^{-6}$	1 micrometre (μm) = 1×10^{-6} m
nano	n	$0.000\,000\,001 = 10^{-9}$	1 nanometre (nm) = 1×10^{-9} m
pico	p	$0.000\,000\,000\,001 = 10^{-12}$	1 picometre (pm) = 1×10^{-12} m
femto	f	$0.000\,000\,000\,000\,001 = 10^{-15}$	1 femtometre (fm) = 1×10^{-15} m
atto	a	$0.000\,000\,000\,000\,000\,001 = 10^{-18}$	1 attometre (am) = 1×10^{-18} m
zepto	z	$0.000\,000\,000\,000\,000\,000\,001 = 10^{-21}$	1 zeptometre (zm) = 1×10^{-21} m
yocto	y	$0.000\,000\,000\,000\,000\,000\,000\,001 = 10^{-24}$	1 yoctometre (ym) = 1×10^{-24} m

1.5.2 Mass and Weight

These two terms, though often used interchangeably, are different quantities.

Mass is a measure of the amount of matter in an object which is constant and does not depend on location.

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Weight is the force that gravity exerts on an object. It depends on the location from and on the celestial body attracting an object. For example, objects on the moon weigh a sixth of their weight on earth. Also, the further a body is from the centre of the attracting body, the lighter it is.

Chemists are primarily interested in mass which can be determined with a balance. The process of measuring mass, oddly, is called weighing.

1.5.3 Volume

Since the SI unit of length is metre (m), the SI derived unit for volume is cubic metre (m³).

Generally, chemists work with much smaller volumes such as cubic centimetre (cm³) and the cubic decimetre (dm³).

$$1 \text{ cm}^3 = (1 \times 10^{-2} \text{ m})^3 = 1 \times 10^{-6} \text{ m}^3 \text{ and}$$

$$1 \text{ dm}^3 = (1 \times 10^{-1} \text{ m})^3 = 1 \times 10^{-3} \text{ m}^3$$

Another common unit of volume is the litre (L). A litre is the volume occupied by one cubic decimetre.

1 litre of volume is equal to 1,000 millilitres (mL) or 1,000 cm³, that is,

$$1 \text{ L} = 1,000 \text{ mL} = 1,000 \text{ cm}^3 = 1 \text{ dm}^3$$

and one millilitre is equal to one cubic centimetre, that is,

$$1 \text{ mL} = 1 \text{ cm}^3$$

1.5.4 Density

The equation for density is

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \quad \text{or} \quad d = \frac{m}{v} \quad (1.1)$$

The SI derived unit of density of kilogram per cubic metre (kg/m³) is too large for chemical applications. Usually chemists use density units of grams per cubic centimetre (g/cm³) or its equivalent, grams per millilitre (g/mL) for solids and liquids, respectively.

Gas density can be expressed in units of grams per litre (g/L):

$$1 \text{ g/cm}^3 = 1 \text{ g/mL} = 1000 \text{ kg/m}^3$$

$$1 \text{ g/L} = 0.001 \text{ g/mL}$$

Example 1.1: A piece of platinum metal with density of 21.5 g/cm³ has a volume of 4.49 cm³. What is its mass?

Solution: From Equation 1.1, we have

$$d = \frac{m}{v} \quad \therefore m = d \times v = 21.5 \frac{\text{g}}{\text{cm}^3} \times 4.49 \text{ cm}^3 = 96.5 \text{ g}$$

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Example 1.2: The density of ethanol, a colourless liquid that is commonly known as grain alcohol is 0.798 g/mL. Calculate the mass of 17.4 mL of the liquid.

Solution: Like in Example 1.1, we use Equation 1.1 where we have

$$d = \frac{m}{v} \quad \therefore m = d \times v = 0.798 \frac{\text{g}}{\text{mL}} \times 17.4 \text{ mL} = 13.9 \text{ g}$$

Example 1.3: The density of sulphuric acid in a certain car battery is 1.41 g/mL. Calculate the mass of 242 mL of the liquid.

Solution: Like in Examples 1.1 and 1.2, we use Equation 1.1 where we have

$$d = \frac{m}{v} \quad \therefore m = d \times v = 1.41 \frac{\text{g}}{\text{mL}} \times 242 \text{ mL} = 341 \text{ g}$$

1.5.5 Temperature Scales

Three common temperature scales are

- (i) degrees Fahrenheit (°F)
- (ii) degrees Celsius (°C) and
- (iii) Kelvin (K)

The scale that is commonly used outside the laboratory in Zambia is the Celsius or centigrade scale. In the USA, the scale commonly used outside the laboratory is the Fahrenheit. The Fahrenheit scale defines the freezing and boiling points of water. The table below compares the three temperature scales.

Table 1.5: Comparison of three common temperature scales

Scale	Water freezing point	Room temperature	Normal body temperature	Water boiling point
Fahrenheit	32°F	77°F	98.6°F	212°F
Celsius	0°C	25°C	37°C	100°C
Kelvin	273K	298K	310K	373K

As shown in Table 1.4, the Kelvin is the SI base unit of temperature. It is the absolute scale. This means the zero on the Kelvin scale, denoted as 0K, is the lowest temperature that can be attained theoretically. On the other hand 0°F and 0°C are based on the behaviour of an arbitrarily chosen substance, water. Equation 1.2 is the conversion equation for degrees Fahrenheit to degrees Celsius is

$$t_{\text{°C}} = (t_{\text{°F}} - 32^{\text{°F}}) \times \frac{5^{\text{°C}}}{9^{\text{°F}}} \quad (1.2)$$

Equation 1.3 is the conversion equation for degrees Celsius to degrees Fahrenheit is

$$t_{\text{°F}} = (t_{\text{°C}} \times \frac{9^{\text{°F}}}{5^{\text{°C}}}) + 32^{\text{°F}} \quad (1.3)$$

Both Celsius and Kelvin Scales have units of equal magnitude, that is, 1°C = 1K. Absolute zero of 0K = -273.15°C. Therefore the conversion equation for degrees Celsius to Kelvin is

$$t_{\text{K}} = (t_{\text{°C}} + 273.15^{\text{°C}}) \times \frac{1 \text{ K}}{1^{\text{°C}}} \quad (1.4)$$

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Example 1.4:

- (a) Solder is an alloy of tin and lead whose melting point is 224°C. What is its melting point in degrees Fahrenheit?
- (b) Helium has the lowest boiling point of all the elements at -452°F. Convert this temperature to degrees Celsius.
- (c) Mercury, the only liquid metal at room temperature, melts at -38.9°C. Convert its melting point to Kelvin.

Solutions:

- (a) Conversion is carried out using Equation 1.3, that is,

$$t_{\text{°F}} = 224^{\circ}\text{C} \times \frac{9^{\circ}\text{F}}{5^{\circ}\text{C}} + 32^{\circ}\text{F} = 435^{\circ}\text{F}.$$

- (b) Here, conversion is obtained by Eqn 1.2, that is,

$$t_{\text{°C}} = (-452^{\circ}\text{F} - 32^{\circ}\text{F}) \times \frac{5^{\circ}\text{C}}{9^{\circ}\text{F}} = -269^{\circ}\text{C}$$

- (c) Getting the melting point of mercury in Kelvin uses Equation 1.4, that is,

$$t_{\text{K}} = (-38.9^{\circ}\text{C} + 273.15^{\circ}\text{C}) \times \frac{1\text{ K}}{1^{\circ}\text{C}} = 234.3\text{ K}$$

Practice Exercise 1.1: Convert

- (a) 327.5°C (the melting point of lead) to degrees Fahrenheit.
- (b) 172.9°F (the boiling point of ethanol) to degrees Celsius; and
- (c) 77 K (the boiling point of liquid nitrogen) to degrees Celsius.
- (d) Helium has the lowest boiling point of all the elements at -452°F. Convert this temperature to degrees Celsius.

1.6 HANDLING OF NUMBERS OF SCIENTIFIC MEASUREMENTS

There are two important techniques of handling numbers associated with chemical measurements, namely,

- (a) scientific notation and
- (b) significant figures

1.6.1 Scientific Notation

Scientific notation is used when chemists handle extremely large or extremely small numbers.

For example, 1 g of the element hydrogen has roughly 602,200,000,000,000,000,000 or 6.022×10^{23} atoms. Each hydrogen atom has a mass of 0.000 000 000 000 000 000 00166 or 1.66×10^{-24} g.

Handling such numbers is cumbersome and mistakes of an extra or a missed zero are easily made. Therefore, when working with very large and/or very small numbers, a system called scientific notation is used. In this system, regardless of their magnitude, all numbers are expressed in the form $N \times 10^n$ where N , the coefficient, is a number between 1 and 10 and n , the exponent, is a positive or negative integer.

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Given a certain number, being asked to express it in scientific notation means finding n . We count the number of places that the decimal point must be moved to give the number N (which is between 1 and 10).

If the decimal point has to be moved to the left, then n is a positive integer; if it has to be moved to the right, n is a negative integer.

Example 1.5:

(a) Express 568.762 in scientific notation

$568.762 = 5.68762 \times 10^2$. The decimal point moved two places to the left $\therefore n = +2$.

(b) Express 0.000 00772 in scientific notation

$0.000\ 00772 = 7.72 \times 10^{-6}$. The decimal point moved six places to the right $\therefore n = -6$.

1.6.1.1 Arithmetic operations of addition and subtraction using scientific notation

Addition and subtraction using scientific notation requires writing each quantity say N_1 and N_2 , with the same exponent. Then we combine N_1 and N_2 while ensuring that the exponents remain the same. For example

- (i) $7.4 \times 10^3 + 2.1 \times 10^3 = (7.4 + 2.1) \times 10^3 = 9.5 \times 10^3$
- (ii) $4.31 \times 10^4 + 3.9 \times 10^3 = (4.31 \times 10^4) + (0.39 \times 10^4) = (4.31 + 0.39) \times 10^4 = 4.70 \times 10^4$
- (iii) $(2.22 \times 10^{-2}) - (4.10 \times 10^{-3}) = (2.22 \times 10^{-2}) - (0.41 \times 10^{-2}) = (2.22 - 0.41) \times 10^{-2} = 1.81 \times 10^{-2}$

1.6.1.2 Arithmetic operations of multiplication and division using scientific notation

To multiply numbers expressed in scientific notation ($A = N_1 \times 10^{n_1}$ and $B = N_2 \times 10^{n_2}$), the parts N_1 and N_2 of A and B , respectively, are multiplied in the usual way but we add the exponents of A and B together, that is, we add n_1 and n_2 as shown in Equation 1.5 below.

$$A \times B = (N_1 \times 10^{n_1}) \times (N_2 \times 10^{n_2}) = N_1 \times N_2 \times 10^{n_1+n_2} \quad (1.5)$$

To divide the numbers A and B using numbers expressed in scientific notation ($A = N_1 \times 10^{n_1}$ and $B = N_2 \times 10^{n_2}$), we divide the parts N_1 and N_2 of A and B , respectively, as usual and subtract the exponents n_1 and n_2 as shown in Equation 1.6 below.

$$A \div B = (N_1 \times 10^{n_1}) \div (N_2 \times 10^{n_2}) = \left(\frac{N_1}{N_2}\right) \times 10^{n_1-n_2} \quad (1.6)$$

The examples below clearly shows the above statements

- (i) $(8.0 \times 10^4) \times (5.0 \times 10^2) = (8.0 \times 5.0) \times 10^{4+2} = 40 \times 10^6 = 4.0 \times 10^7$
- (ii) $(4.0 \times 10^{-5}) \times (7.0 \times 10^3) = (4.0 \times 7.0) \times 10^{-5+3} = 28 \times 10^{-2} = 2.8 \times 10^{-1}$
- (iii) $(6.9 \times 10^7) \div (3.0 \times 10^{-5}) = \left(\frac{6.9}{3.0}\right) \times 10^{7-(-5)} = 2.3 \times 10^{12}$
- (iv) $(8.5 \times 10^4) \div (5.0 \times 10^9) = \left(\frac{8.5}{5.0}\right) \times 10^{4-9} = 1.7 \times 10^{-5}$

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1.6.2 Significant Figures

Only integers are exact. Measurement quantities, which are not integers have a margin of error in the measurement. This is clearly shown by having a number of significant figures, which are the meaningful digits in a measured or calculated quantity. When significant figures are used, the last digit is understood to be uncertain.

1.6.2.1 Guidelines for determining significant figures

There are five guidelines for determining significant figures of a given number. These are

- (i) Any non-zero digit is significant. Thus 845 cm has three significant figures, 1.234 kg has four significant figures, etc.
- (ii) Zeros between non-zero digits are significant. Thus 606 m contains three significant figures, 40501 kg contains five significant figures and so on.
- (iii) Zeros to the left of the first non-zero digit are not significant. Their purpose is to indicate the placement of the decimal point. For example, 0.08 L contains one significant figure, 0.000034 g contains two significant figures and so on.
- (iv) If a number is > 1 then all the zeros written to the right of the decimal point count as significant figures. Thus 2.0 mg has two significant figures; 40.062 has five significant figures; and 3.040 dm has four significant figures.

If a number is < 1 , then only the zeros that are at the end of the number and the zeros that are between non-zero digits are significant. This means 0.090 kg has two significant figures, 0.3005 has four significant figures, 0.00420 min has three significant figures and so on.

- (v) In numbers that do not contain decimal points the trailing zeros (that is, zeros after the last non-zero digit) may or may not be significant. Thus, 400 cm may have one significant figure (the digit 4), two significant figures (40) or three significant figures (400). We cannot know the correct number of significant figures without more information. By using scientific notation, however, we avoid this ambiguity. In this particular case we can express the number 400 as

4×10^2 for one significant figure

4.0×10^2 for two significant figures

4.00×10^2 for three significant figures

Example 1.6:

Determine the number of significant figures in the following measurements

(a) 478 cm (b) 6.01 g (c) 0.825 m (d) 0.043 kg (e) 1.310×10^{22} atoms (f) 7000 mL

Answers

(a) Three

(b) Three

(c) Three

(d) Two

(e) Four

(f) Ambiguous case. The number of significant figures may be 4 (7.000×10^3), 3 (7.00×10^3), 2 (7.0×10^3) or 1 (7×10^3).

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1.6.2.2 Rules for handling significant figures during addition and subtraction calculations

In addition and subtraction, the answer cannot have more digits to the right of the decimal point than either of the original numbers. Consider these examples.

- (i)
$$\begin{array}{r} 89.332 \\ + 1.1 \\ \hline = 90.432 \end{array}$$
 ← one digit after the decimal point
← round off to 90.4
- (ii)
$$\begin{array}{r} 2.097 \\ - 0.12 \\ \hline = 1.977 \end{array}$$
 ← two digits after the decimal point
← round off to 1.98

The rounding procedure is as follows.

- (a) To round off a number at a certain point we simply drop off the digit that follow if the first of them is less 5. Thus 8.724 rounds too 8.72 if we want only two digits after the decimal point.
- (b) If the first digit following the point of rounding off is equal to or greater than 5 we add 1 to the preceding digit. Thus, 8.727 rounds off to 8.73 and 0.425 rounds off to 0.43.

1.6.2.3 Rules for handling significant figures during multiplication and division calculations

In multiplication and division, the number of significant figures in the final product or quotient is determined by the original number that has the smallest number of significant figures. The following examples illustrate this rule.

- (i)
$$\begin{array}{r} 4.5039 \\ \times 2.8 \\ \hline = 12.60192 \end{array}$$
 ← one digit after the decimal point
← round off to two significant figures 13
- (ii)
$$\frac{6.85}{112.04} = 0.0061138878$$
 ← round off to three significant figures = 0.00611
Three significant figures

It should be kept in mind that the exact numbers obtained from definitions or by counting numbers of objects can be considered to have an infinite number of significant figures. If an object has a mass of 0.2786 g, then the mass of eight such objects is $0.2786 \times 8 = 2.229$ g. We do not round off this product to one significant figure because the number 8 is 8.0000... by definition.

Similarly, to take the average of two measured lengths 6.64 cm and 6.68 cm, we write

$$\frac{6.64 \text{ cm} + 6.68 \text{ cm}}{2} = 6.66 \text{ cm.}$$

Example 1.7:

Carry out the following operations to the correct number of significant figures.

- (i) $11,254.1 \text{ g} + 0.1983 \text{ g}$

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- (ii) $66.59 \text{ L} - 3.113 \text{ L}$
- (iii) $8.16 \text{ m} \times 5.1355$
- (iv) $0.0154 \text{ kg} \div 88.3 \text{ mL}$
- (v) $2.64 \times 10^3 \text{ cm} \times 3.27 \times 10^2 \text{ cm}$

Answers:

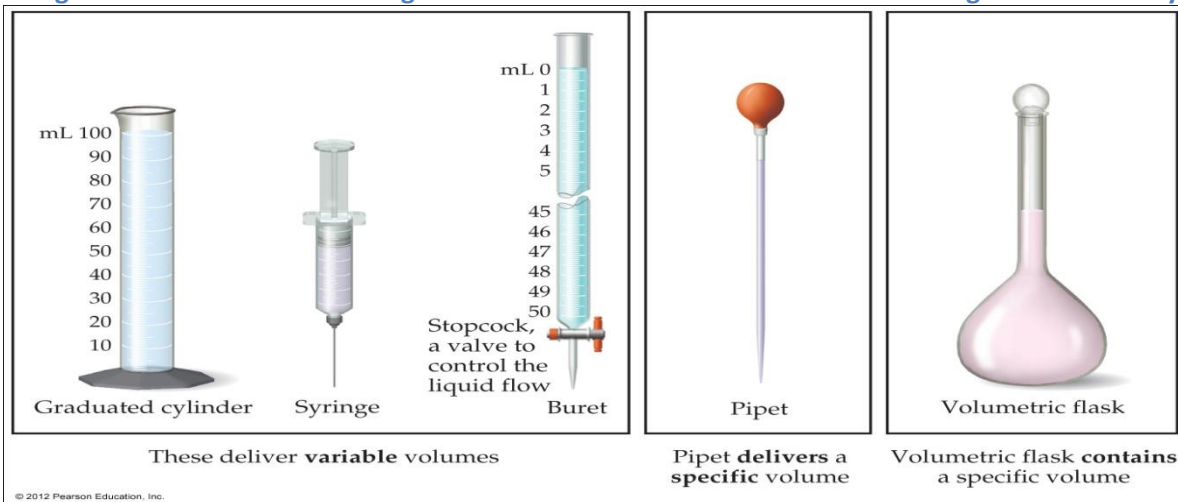
- (i) $1,254.1 \text{ g} \leftarrow \text{one decimal}$
 $+0.1983 \text{ g}$
 $= 1,254.2983 \text{ g} \leftarrow \text{round off to one decimal} = \underline{1,254.3 \text{ g}}$
- (ii) $66.59 \text{ L} \leftarrow \text{two decimals}$
 $- 3.113 \text{ L}$
 $= 63.477 \text{ L} \leftarrow \text{round off to two decimals} = \underline{63.48 \text{ L}}$
- (iii) $8.16 \text{ m} \leftarrow \text{three significant figures}$
 $\times 5.1355$
 $= 41.90568 \text{ m} \leftarrow \text{round off to three significant figures} = \underline{41.9 \text{ m}}$
- (iv) $\frac{0.0154 \text{ kg}}{88.3 \text{ mL}} = 0.00017440546 \text{ kg/mL} \leftarrow \text{round off to 3 sig figs} = 0.000174 \text{ kg/mL}$
 $\text{Three significant figures} \quad \text{or } 1.74 \times 10^{-4} \text{ kg/mL}$
- (v) $2.64 \times 10^3 \text{ cm} \times 3.27 \times 10^2 \text{ cm} = (2.64 \times 3.27) \times 10^{3+2} \text{ cm}^2 = 8.6328 \times 10^5 \text{ cm}^2$
 $\text{Three sig figs for both numbers} \quad \text{round off to 3 sig figs} = 8.63 \times 10^5 \text{ cm}^2$

NB: when using electronic machines you round off the answer of a chain of calculations at the end of your calculations.

1.7 MEASUREMENT ERROR OR UNCERTAINTY

The numbers associated with a measurement is obtained by using a measurement device. Table 1.1 and Figure 1.4 show examples of such devices used by chemistry students.

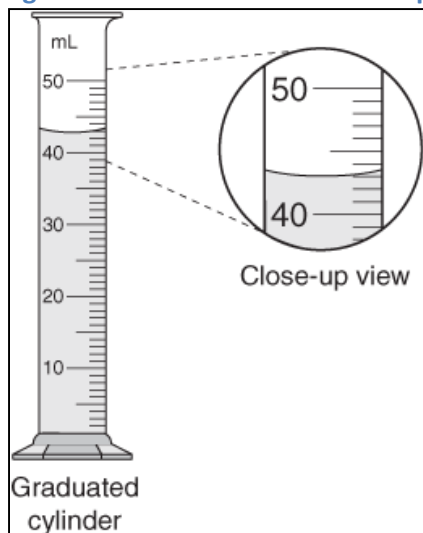
Figure 1.4: Different Measuring Devices with Different Uses and Different Degrees of Accuracy



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A measurement that results from use of device always has some degree of uncertainty as shown in Table 1.6 below. As shown Figure 1.5, uncertainty in the measurement is indicated by recording certain digits and the first uncertain digit. These digits are called significant figures of a measurement.

Figure 1.5: Basis for Number of Reported 3 Significant Figures



**Reading is estimated to be 43.2 mL.
The last digit is an estimated number.**

Table 1.6: Examples of Uncertainty of Volumetric Glassware¹

Glassware	Volume in mL	± Uncertainty in mL	% of Volume
Volumetric pipettes	1.00	0.01	1.00
	5.00	0.01	0.20
	10.00	0.02	0.20
	25.00	0.03	0.12
Volumetric flasks	50.00	0.05	0.10
	100.00	0.08	0.08
	250.0	0.10	0.04
Burette	50.00	0.05	0.10
	100.00	0.10	0.10
Erlenmeyer flasks	100	5	5
	250	10	4
Beaker	50	5	10
	100	5	5
Graduated cylinder	10.0	0.1	1
	100.0	0.5	0.5

¹J. S. Fritz and G. H. Schenk, *Quantitative Analytical Chemistry*, 3rd ed., Allyn & Bacon, Boston, 1974, p. 560

http://academics.wellesley.edu/Chemistry/Chem105manual/Appendices/uncertainty_volumetric.html

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1.7.1 Precision and Accuracy

Accuracy tells how close a measurement is to the true value of the quantity that was measured while **precision** refers to how closely two or more measurements of the same quantity are with one another.

When doing target practice at the range or in a game of darts, hitting the bull's eye consistently as shown in Figure 1.6a is considered high accuracy. Normally, highly accurate scores/measurements like the ones shown in this figure are usually precise too. However, highly precise measurements (or scores) do not necessarily guarantee accurate results as shown in Figure 1.6b. Figure 1.6c shows measurements (or scores) that are inaccurate and imprecise.

1.7.2 Random and Systematic Errors

In a measurement process, the difference between the true value and the measurement is called the measurement error. There are two types of measurement error, namely,

- (i) Random or reproducibility errors and
- (ii) Systematic errors.

A **random error** (also called indeterminate error) means that a measurement has an equal chance of being high or low while a **systematic error** (or determinate error) occurs in the same direction each time; it is either always high or always low.

Figure 1.2a shows small random errors and essentially no or small systematic errors. Figure 1.2b shows small random errors but large systematic errors. Figure 1.2c shows large random errors (poor technique) and most likely large systematic errors.

In the laboratory large random errors or poor technique usually comes from lack of measurement skill – plainly speaking, failing to master experimental activities and equipment. Large systematic errors are often associated with faulty experimental laboratory operating procedures or poorly calibrated equipment. Ultimately large systematic errors also reflect poor laboratory skill of the laboratory managers who are probably uncertified analysts operating laboratories that do not participate in proficiency analytical testing (PAT). Of necessity, such laboratories are not internationally accredited.

In quantitative measurements, some people often use precision to indicate accuracy. In this situation, one assumes that the average of a series of precise measurements is accurate or close to the true value. As shown in Figures 1.2a and 1.2b, this assumption is only valid if systematic errors are absent!

Figure 1.6: Accuracy versus Precision

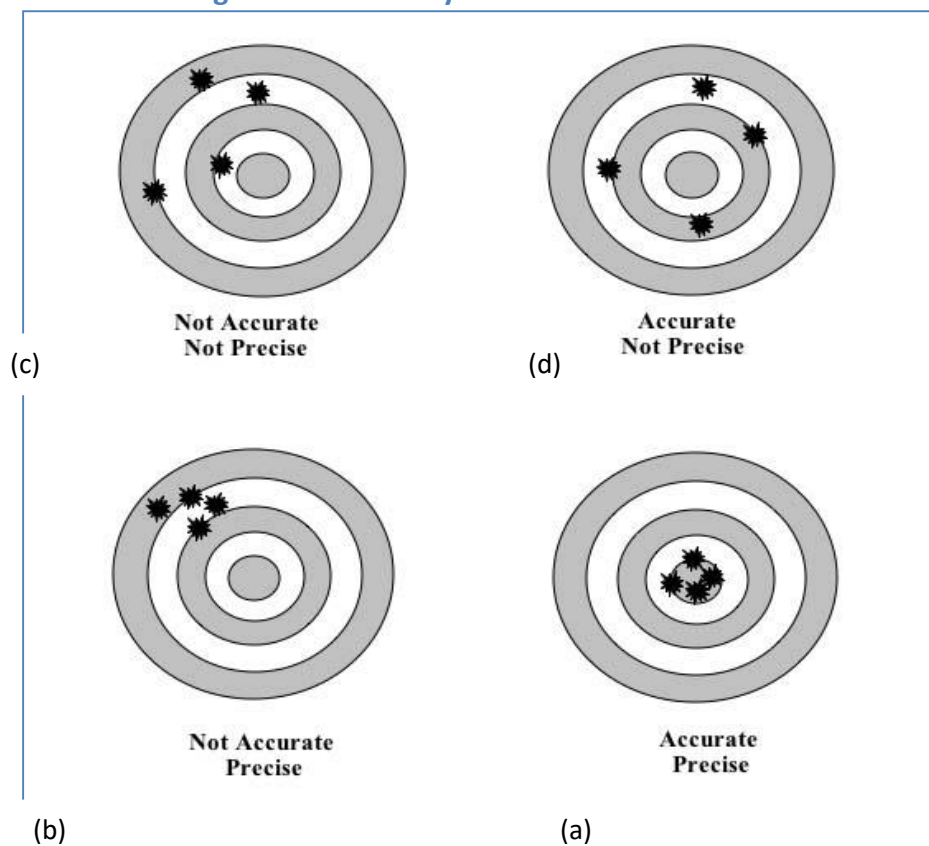


Figure 1.7: Reproducibility Error versus Systematic Error

	Accurate	Inaccurate (systematic error)
Precise		
Imprecise (reproducibility error)		

http://academics.wellesley.edu/Chemistry/Chem105manual/Appendices/uncertainty_analysis.html

(accessed on July 3, 2015)

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Table 1.7: Summary on Precision and Accuracy Attributes

Precision	Accuracy
Reproducibility	correctness
check by repeating measurements	check by using a different method
poor precision results from poor technique	poor accuracy results from procedural or equipment flaws
poor precision is associated with 'random errors' - error has random sign and varying magnitude. Small errors more likely than large errors.	poor accuracy is associated with 'systematic errors' - error has a reproducible sign and magnitude.

<http://antoine.frostburg.edu/chem/senese/101/measurement/>

1.7.3 Notes on Uncertainty Reporting as Significant Figures and Uncertainty Analysis

http://academics.wellesley.edu/Chemistry/Chem105manual/Appendices/uncertainty_sigfigs.html

http://academics.wellesley.edu/Chemistry/Chem105manual/Appendices/uncertainty_analysis.html and <http://www.physics.rutgers.edu/ugrad/389/errors.pdf> (accessed 3 July 2015)

1.8 DIMENSIONAL ANALYSIS OR FACTOR LABEL METHOD

Dimensional analysis amounts to using the units of a parameter to solve problems. This problem solving approach uses conversion factors to change the units of a parameter. **Solving a dimensional analysis problem starts with an equivalence statement that gives two conversion factors.**

In general terms, an equivalence statement is a mathematical expression of two values linked by an equals sign such that the first value is put on the left hand side (LHS) of the equals sign while the second value is put on the right hand side (RHS) as shown in equation 1.7 below

$$\text{LHS} = \text{RHS} \quad (1.7)$$

This statement gives two conversion factors. The first one is obtained by dividing both sides of Equation 1.7 with LHS to give Equation 1.8a

$$\frac{\text{LHS}}{\text{LHS}} = \frac{\text{RHS}}{\text{LHS}} \quad \text{or} \quad 1 = \frac{\text{RHS}}{\text{LHS}} \quad (1.8a)$$

The second conversion factor is obtained by dividing both sides of Equation 1.7 with RHS to give Equation 1.8b as shown below.

$$\frac{\text{LHS}}{\text{RHS}} = \frac{\text{RHS}}{\text{RHS}} \quad \text{or} \quad \frac{\text{LHS}}{\text{RHS}} = 1 \quad (1.8b)$$

When these conversion factors of Equations 1.8a and 1.8b are contracted into a single statement, they give Equation 1.8

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$$\frac{RHS}{LHS} = 1 = \frac{LHS}{RHS} \quad (1.8)$$

When solving a problem using dimensional analysis, one must multiply the value to be converted by the conversion factor that will give the correct units in the answer.

It is to be noted that the problem being solved can use any type of units (that is SI base units, SI derived units, imperial units, some other units). A few examples will clarify this.

Example 1.8: Write the conversion factors of the equivalence statement 1 foot = 12 inches.

Solution: Noting that the abbreviation of foot or feet is ft and that of inch or inches is in, use of Equation 1.8a gives the first conversion factor, that is,

$$\frac{LHS}{RHS} = 1 \text{ or } \frac{1 \text{ ft}}{12 \text{ in}} = 1.$$

Use of Equation 1.8b gives the second conversion factor which is

$$\frac{RHS}{LHS} = 1 \text{ or } \frac{12 \text{ in}}{1 \text{ ft}} = 1$$

The contracted statement of the two conversion factors formulated from Equation 1.8 is

$$\frac{1 \text{ ft}}{12 \text{ in}} = 1 = \frac{12 \text{ in}}{1 \text{ ft}}$$

Example 1.9: Given the equivalence statements 11 yds = 2 rods, 40 rods = 1 furlong and 8 furlongs = 1 mile, convert 1.25 miles to (i) rods, (ii) 6 furlongs, (iii) metres, and (iv) kilometres.

Solution:

(ii) It is easiest to solve (ii) using the last equivalence statement which, according to Equation 1.8 gives conversion factors (a) and (b).

$$\frac{8 \text{ furlongs}}{1 \text{ mile}} = 1 = \frac{1 \text{ mile}}{8 \text{ furlongs}}$$

Since we are converting miles to furlongs, multiplying conversion factor (a) gives

$$1.25 \text{ miles} = 1.25 \text{ miles} \times \frac{8 \text{ furlongs}}{1 \text{ mile}} = 10 \text{ furlongs} = 10.0 \text{ furlongs}$$

The answer has the desired units, thus, the appropriate conversion factor has been used.

We can check this by using conversion factor (b) which gives

$$1.25 \text{ miles} = 1.25 \text{ miles} \times \frac{1 \text{ mile}}{8 \text{ furlongs}} = \frac{1.25}{8} \frac{\text{mile}^2}{\text{furlongs}}$$

This result is wrong because dimensional analysis yields the wrong units!

(i) Using the answer of (ii) above and the appropriate conversion factor from using the second equivalence statement, according to Equation 1.8 gives conversion factors (a) and (b).

$$\frac{40 \text{ rods}}{1 \text{ furlong}} = 1 = \frac{1 \text{ furlong}}{40 \text{ rods}}$$

(a)
(b)

Again use of conversion factor (a) gives

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$$1.25 \text{ miles} = 10 \text{ furlongs} \times \frac{40 \text{ rods}}{1 \text{ furlong}} = 400 \text{ rods} = 4.00 \times 10^2 \text{ rods}$$

Since the answer has the correct units, the correct conversion factor has been used.

- (iii) Using the answer of (i) above and the appropriate conversion factor from using the first equivalence statement, according to Equation 1.8 gives conversion factors (a) and (b).

$$\begin{array}{ccc} \frac{11 \text{ yds}}{2 \text{ rods}} & = 1 = & \frac{2 \text{ rods}}{11 \text{ yds}} \\ \text{(a)} & & \text{(b)} \end{array}$$

Use of conversion factor (a) converts miles to yards when the answer to (i) is employed as shown below

$$1.25 \text{ miles} = 400 \text{ rods} \times \frac{11 \text{ yds}}{2 \text{ rods}} = 2200 \text{ yds} = 2.20 \times 10^3 \text{ yds}$$

To do the final conversion of yards to metres, we need an equivalence statement for yards and metres that was not provided in the question. This statement can be obtained from tables or can be arrived at using other equivalence statements. The latter approach will be using to conclude solving this problem.

The equivalent statements required are

1. 1 yd = 3 ft whose conversion factors are

$$\begin{array}{ccc} \frac{1 \text{ yd}}{3 \text{ ft}} & = 1 = & \frac{3 \text{ ft}}{1 \text{ yd}} \\ \text{(a)} & & \text{(b)} \end{array}$$

2. 1 ft = 12 in whose conversion factors are

$$\begin{array}{ccc} \frac{1 \text{ ft}}{12 \text{ in}} & = 1 = & \frac{12 \text{ in}}{1 \text{ ft}} \\ \text{(a)} & & \text{(b)} \end{array}$$

3. 1 in = 2.53 cm whose conversion factors are

$$\begin{array}{ccc} \frac{1 \text{ in}}{2.53 \text{ cm}} & = 1 = & \frac{2.53 \text{ cm}}{1 \text{ in}} \\ \text{(a)} & & \text{(b)} \end{array}$$

4. 1 m = 100 cm whose conversion factors are

$$\begin{array}{ccc} \frac{1 \text{ m}}{100 \text{ cm}} & = 1 = & \frac{100 \text{ cm}}{1 \text{ m}} \\ \text{(a)} & & \text{(b)} \end{array}$$

If we denote a conversion factor of the above equivalent statements as CV(nx) where n is number of the equivalent statement and x indicates the appropriate conversion factor, a or b, then the equivalence statement for yards and metres is obtained from the multiplication:

$$1 \text{ yd} = 1 \text{ yd} \times CV(1b) \times CV(2b) \times CV(3b) \times CV(4a)$$

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Putting in the actual conversion factors yields

$$1 \text{ yd} = 1 \text{ yd} \times \frac{3 \text{ ft}}{1 \text{ yd}} \times \frac{12 \text{ in}}{1 \text{ ft}} \times \frac{2.53 \text{ cm}}{1 \text{ in}} \times \frac{1 \text{ m}}{100 \text{ cm}}$$

This gives the 5th equivalent statement

$$5. \quad 1 \text{ yd} = \frac{3 \times 12 \times 2.53 \text{ m}}{100} = 0.9108 \text{ m}$$

This equivalence statement gives the conversion factors given below according to Equation 1.8

$$\frac{1 \text{ yd}}{0.9108 \text{ m}} = 1 = \frac{0.9108 \text{ m}}{1 \text{ yd}}$$

(a) (b)

Thus converting 1.25 miles to metres simply requires multiplication of 2200 yds with conversion factor 5b, that is

$$1.25 \text{ miles} = 2200 \text{ yds} \times \frac{0.9108 \text{ m}}{1 \text{ yd}} = 2004 \text{ m} = 2.00 \times 10^3 \text{ m}$$

- (iv) Using the answer of (iii) above and the appropriate conversion factor from the first equivalence statement, 1 km = 1000 m, that according to Equation 1.8 are conversion factors (a) and (b) below .

$$\frac{1 \text{ km}}{1000 \text{ m}} = 1 = \frac{1000 \text{ m}}{1 \text{ km}}$$

(a) (b)

This gives

$$1.25 \text{ miles} = 2004 \text{ m} \frac{1 \text{ km}}{1000 \text{ m}} = 2.004 \text{ km} = 2.00 \text{ km}$$

Example 1.10: Science fiction often uses nautical analogies to describe space travel. If the starship U.S.S. Enterprise is travelling at warp factor 1.71, what is its speed in knots? Use the following equivalent statements to solve your problem:

1. Warp 1.71 = 5.00 times the speed of light
2. Speed of light = $3.00 \times 10^8 \text{ m/s}$
3. 1 knot = 2000 yd/h exactly

Solution:

Conversion factors from the above equivalent statements are

1. 1.71 Warp factor = 5.00 times the speed of light has conversion factors

$$\frac{1.71 \text{ Warp factor}}{5.00 \text{ light speed}} = 1 = \frac{5.00 \text{ light speeds}}{1.71 \text{ Warp factor}}$$

(a) (b)

2. Speed of light or 1 light speed = $3.00 \times 10^8 \text{ m/s}$ has conversion factors

$$\frac{1 \text{ light speed}}{3.00 \times 10^8 \text{ m/s}} = 1 = \frac{3.00 \times 10^8 \text{ m/s}}{1 \text{ light speed}}$$

(a) (b)

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3. 1 knot = 2000 yd/h has conversion factors

$$\frac{1 \text{ knot}}{2000 \text{ yd/h}} = 1 = \frac{2000 \text{ yd/h}}{1 \text{ knot}}$$

(a) (b)

Missing equivalence statements with their conversion factors are

4. 1 yd = 0.9108 m whose conversion factors are

$$\frac{1 \text{ yd}}{0.9108 \text{ m}} = 1 = \frac{0.9108 \text{ m}}{1 \text{ yd}}$$

(a) (b)

5. 1 h = 3600 s whose conversion factors are

$$\frac{1 \text{ h}}{3600 \text{ s}} = 1 = \frac{3600 \text{ s}}{1 \text{ h}}$$

(a) (b)

Using information from the latter two conversion factors can be used to determine the equivalent statement relating yd/h to m/s as

$$1 \frac{\text{yd}}{\text{h}} = 1 \frac{\text{yd}}{\text{h}} \times CV(4b) \times CV(5a)$$

On substituting the relevant values gives the 6th equivalence statement for this problem

$$1 \frac{\text{yd}}{\text{h}} = 1 \frac{\text{yd}}{\text{h}} \times \frac{0.9108 \text{ m}}{1 \text{ yd}} \times \frac{1 \text{ h}}{3600 \text{ s}}$$

The simplified equivalence statement is

6. 1 yd/h = 2.53×10^{-4} m/s whose conversion factors are

$$\frac{1 \text{ yd/h}}{2.53 \times 10^{-4} \text{ m/s}} = 1 = \frac{2.53 \times 10^{-4} \text{ m/s}}{1 \text{ yd/h}}$$

(a) (b)

Conversion factor 6b can be used to get the 7th equivalence statement knots and m/s as follows

$$1 \text{ knot} = 2000 \text{ yd/h} = 2000 \text{ yd/h} \times \frac{2.53 \times 10^{-4} \text{ m/s}}{1 \text{ yd/h}} = 2 \times 0.253 \text{ m/s} = 0.503 \text{ m/s}$$

7. 1 knot = 0.503 m/s whose conversion factors are

$$\frac{1 \text{ knot}}{0.503 \text{ m/s}} = 1 = \frac{0.503 \text{ m/s}}{1 \text{ knot}}$$

(a) (b)

Conversion of Warp factors to knots requires multiplication of 1.71 Warp factors by CV(1b), CV(2b) and CV(7a) as shown below

$$1.71 \text{ Warp} = 1.71 \text{ Warp} \times CV(1b) \times CV(2b) \times CV(7a)$$

On substituting appropriate conversion factors we have

$$1.71 \text{ Warp} = 1.71 \text{ Warp} \times \frac{5.00 \text{ light speeds}}{1.71 \text{ Warp factor}} \times \frac{3.00 \times 10^8 \text{ m/s}}{1 \text{ light speed}} \times \frac{1 \text{ knot}}{0.503 \text{ m/s}}$$

This simplifies to

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$$\underline{1.71 \text{ Warp factor} = 3.00 \times 10^9 \text{ knots}}$$

2014 FO 130 TEST 1 – Question 1(c)

A star is estimated to have a mass of $2 \times 10^{36} \text{ kg}$. Assuming it to be a sphere of average radius $7.0 \times 10^5 \text{ km}$, calculate the average density of the star in units of grams per cubic centimetre

Solution:

According to Equation 1.1, density of a substance is given as

$$d = \frac{m}{v} \quad (1.1)$$

while volume of a sphere is given by Equation 1.9 below.

$$v = \frac{4\pi r^3}{3} \quad (1.9)$$

where r is the radius of the sphere.

Substituting equation 1.9 into Equation 1.1 gives Equation 1.10 which is the density of a sphere.

$$d = \frac{3m}{4\pi r^3} \quad 1.10$$

The mass and radius of the star are given in kg and km, respectively. For its density to be given in grams per cubic centimetre (g/cm^3), then mass has to be expressed in grams (g) while radius has to be in centimetres (cm). This requires two equivalence statements.

1. $1 \text{ kg} = 1000 \text{ g}$ whose conversion factors are

$$\frac{1 \text{ kg}}{1000 \text{ g}} = 1 = \frac{1000 \text{ g}}{1 \text{ kg}}$$

(a) (b)

2. $1 \text{ km} = 100\,000 \text{ cm}$ whose conversion factors are

$$\frac{1 \text{ km}}{100\,000 \text{ cm}} = 1 = \frac{100\,000 \text{ cm}}{1 \text{ km}}$$

(a) (b)

Using conversion factor 1b, the mass of the star in g is

$$2 \times 10^{36} \text{ kg} = 2 \times 10^{36} \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 2 \times 10^{39} \text{ g}$$

Similarly, using conversion factor 2b, the radius of the star in cm is

$$7 \times 10^5 \text{ km} = 7 \times 10^5 \text{ km} \times \frac{100\,000 \text{ cm}}{1 \text{ km}} = 7 \times 10^{10} \text{ cm}$$

Substituting these values in equation 1.10 gives the density of the star as

$$d = \frac{3m}{4\pi r^3} = \frac{3 \times 2 \times 10^{39} \text{ g}}{4\pi (7 \times 10^{10} \text{ cm})^3} = \frac{6 \times 10^{39} \text{ g}}{1.372 \times 3.142 \times 10^{33} \text{ cm}^3}$$

The density of the star works out to be

$$d = \frac{6 \times 10^{39} \text{ g}}{4.311 \times 10^{33} \text{ cm}^3} = 1.39 \times 10^6 \text{ g/cm}^3$$

Since the density of the star has to be given to be 2 significant figures, the answer is

$$\underline{d = 1.4 \times 10^6 \text{ g/cm}^3}$$

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2014 FO 130 TEST 1 – Question 1(i)

The ideal gas equation is given by $PV = nRT$, where P is pressure in Newton per square area, V is volume, n is the number of moles, T is temperature and R is the gas constant. Show by units analysis that the derived unit of $R = \text{Joule} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$.

Solution:

Solving this problem requires making R the subject of the formula of the given equation, that is,

$$R = \frac{P \times V}{n \times T}$$

In the above equation, P is pressure, V is volume, n is the number of moles and T is temperature. Table 1.3 shows the unit of pressure. The table below therefore shows the units of the parameters and parameter products of the above equation. The last entry shows that the derived unit of $R = \text{Joule} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$.

Variable or Product of Variables	Units
P	$\text{Pascal} = \text{Pa} = \text{kg m}^{-1} \text{s}^{-2} = \text{N m}^{-2}$
V	m^3
n	mol
T	K
$PV = P \times V$	$(\text{N m}^{-2}) \times (\text{m}^3) = \text{N m} = \text{Joule}$
$nT = n \times T$	$(\text{mol}) \times (\text{K}) = \text{mol K}$
$R = \frac{PV}{nT}$	$\frac{\text{Joule}}{\text{mol K}} = \text{Joule mol}^{-1} \text{K}^{-1}$

1.9 DALTON'S ATOMIC THEORY OF MATTER

This theory which states that matter is composed of discrete units called atoms began as a philosophical concept in ancient Greece by Democritus (ca 460 – ca 370 BC). Lavoisier (1743 -1794), the “father of modern chemistry”, laid the experimental foundation of modern chemistry by putting together the first extensive list of elements that included Oxygen, Hydrogen, Silicon and Sulphur.

1.9.1 Postulates of Dalton's Atomic Theory

The Atomic Theory entered the scientific mainstream in the early 19th century when John Dalton (1766-1844) provided the basic theory by observing that all matter whether in the form of an element, a compound or a mixture – is composed of small particles called atoms. The postulates or basic statements of Dalton's atomic theory are an explanation of the structure of matter in terms of different combinations of very small particles. These postulates or **hypotheses** are;

- All matter is composed of indivisible atoms. An **atom** is an extremely small particle of matter that retains its identity during chemical reactions.
- An **element** is a type of matter composed of only one kind of atom, each atom of a given kind having the same properties. Mass is one such property. Thus, the atoms of a given element have a characteristic mass.

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- (iii) A **compound** is a type of matter composed of atoms of two or more elements chemically combined in fixed proportions. The relative numbers of any two kinds of atoms in a compound occur in simple ratios. For example, water a compound of hydrogen and oxygen has an atomic ratio of 2 to 1, respectively.
- (iv) A chemical reaction consists of the rearrangement of atoms present in the reacting substances to give new chemical combinations present in the substances produced by the reaction. Atoms are not created, destroyed, or broken into smaller particles by any chemical reactions.

Today we know that atoms are not truly indivisible; they are themselves made up of particles. Nevertheless, Dalton's postulates are essentially correct.

1.9.2 Atomic Symbols and Models

It is convenient to use symbols for the atoms of different elements. An **atomic symbol** is a one- or two-letter notation used to represent an atom corresponding to a particular element. Typically, the atomic symbol consists of the first letter, capitalized, from the name of the element, sometimes with an additional letter from the name in lowercase. For example, chlorine has the symbol Cl. Other symbols are derived from a name in another language (usually Latin). Sodium is given the symbol Na from its Latin name, natrium.

1.9.3 Deductions from Dalton's Atomic Theory

The atomic theory explains the difference between an element and a compound. It also explains two laws that existed at the time of Dalton's formulation of his postulates; namely **the law of conservation of mass** which states that the total mass during a chemical reaction remains constant and the **law of definite proportions (or law of constant composition)** which states that a pure compound, whatever its source, always contains definite or constant proportions of the elements by mass. The atomic theory was also used to predict the **law of multiple proportions** which states that for two (or more) compounds composed of the same elements, for a given mass one of the elements, the ratio of masses of any other element in the compounds is a small whole-number ratio.

For example, hydrogen peroxide contains 0.06300 g of hydrogen for each gram of oxygen present. Water contains 0.1260 g of hydrogen for each gram of oxygen present. For the fixed mass of oxygen (1 g in each case), the ratio of masses of hydrogen is $\frac{0.1260 \text{ g}}{0.06300 \text{ g}} = \frac{2}{1}$.

Sometimes, the ratio does not appear to be integral at first, but it can be converted to an integral ratio while keeping the value the same by multiplying both numerator and denominator by the same small integer. For example, a compound of chromium and oxygen contains 3.2499 g of chromium per gram of oxygen, while a second compound of these elements contains 2.1666 g of chromium per gram of oxygen. Per gram of oxygen, the ratio of masses of chromium is

$$\frac{3.2499 \text{ g}}{2.1666 \text{ g}} = \frac{1.5000}{1}$$

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We can convert that ratio to an integral ratio by multiplying both numerator and denominator by 2:

$$\frac{3.2499 \text{ g}}{2.1666 \text{ g}} = \frac{1.5000 \times 2}{1 \times 2} = \frac{3}{2}$$

1.10 EARLY EXPERIMENTS TO CHARACTERISE THE ATOM

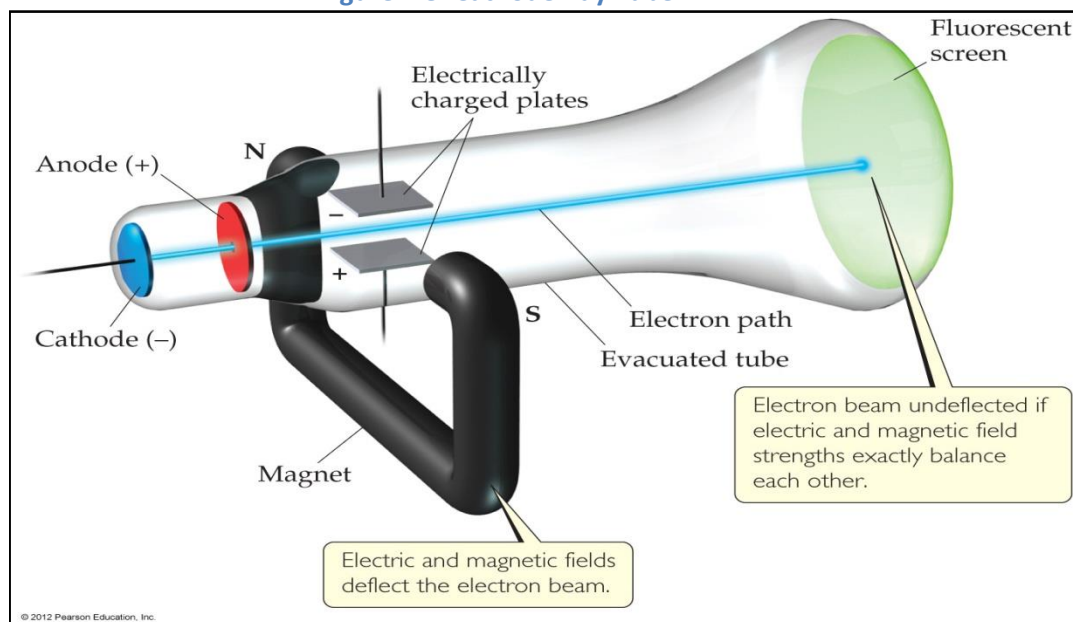
Despite Dalton's postulated indivisibility of atoms, experiments conducted around the beginning of the 20th century showed that atoms consist of particles. These experiments showed that atoms consist of two kinds of particles, namely, **a nucleus**, which is the atom's central core, that is positively charged and contains most of the atom's mass, and one or more electrons. An **electron** is a very light, negatively charged particle that exists in the region around the atom's positively charged nucleus.

1.10.1 Discovery of the Electron

The electron was discovered in 1897 by Joseph John Thomson (1856-1940) during cathode ray experiments (or his studies of conduction of electricity through gases) that provided a value of the ratio of the electron's charge (e) to its mass (m_e) as $\frac{e}{m_e} = -1.76 \times 10^8$ coulombs/gram.

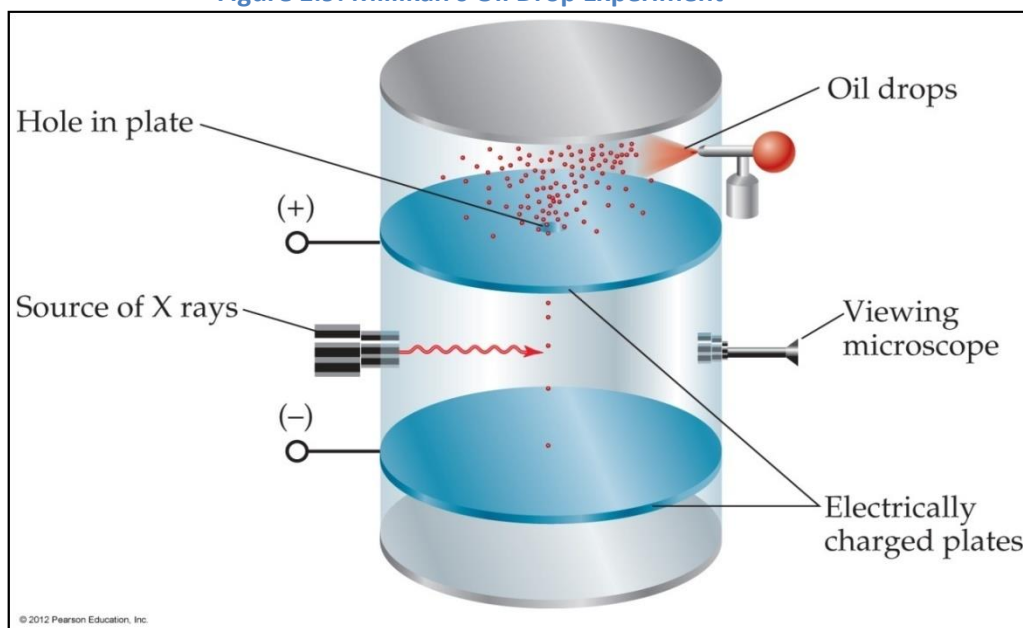
See: http://www-outreach.phy.cam.ac.uk/camphy/electron/electron_index.htm

Figure 1.8: Cathode Ray Tube



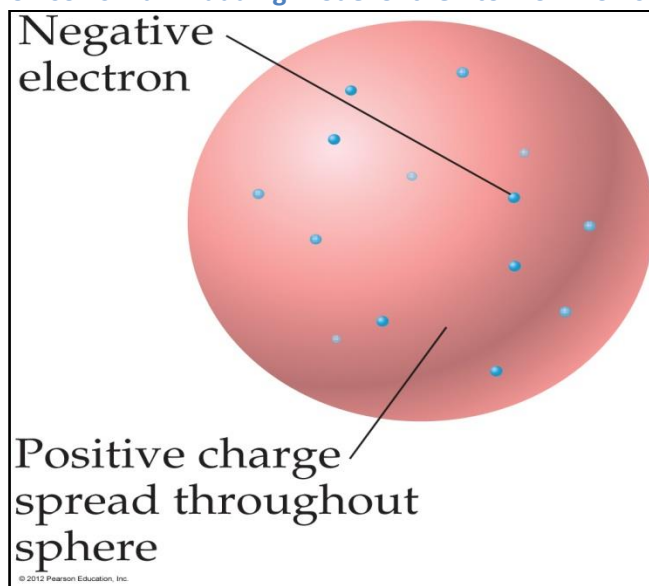
In 1909 Robert Millikan's study of the movement of charged oil drops in the presence and absence of an electric field found the charge of an electron to be 1.602×10^{-19} coulombs so that the ratio $\frac{e}{m_e} = -1.76 \times 10^8$ coulombs/gram was used to calculate the mass of the electron as 9.109×10^{-28} g or 9.109×10^{-31} kg.

Figure 1.9: Millikan's Oil Drop Experiment



The electron was found to be 1840 times lighter than the lightest atom (hydrogen). This clearly shows that the electron is a subatomic particle. Thomson's first proposed model of the atom was that of Plum Pudding shown in Figure 1.10

Figure 1.10: Thomson's Plum Pudding Model of the Atom around 1900

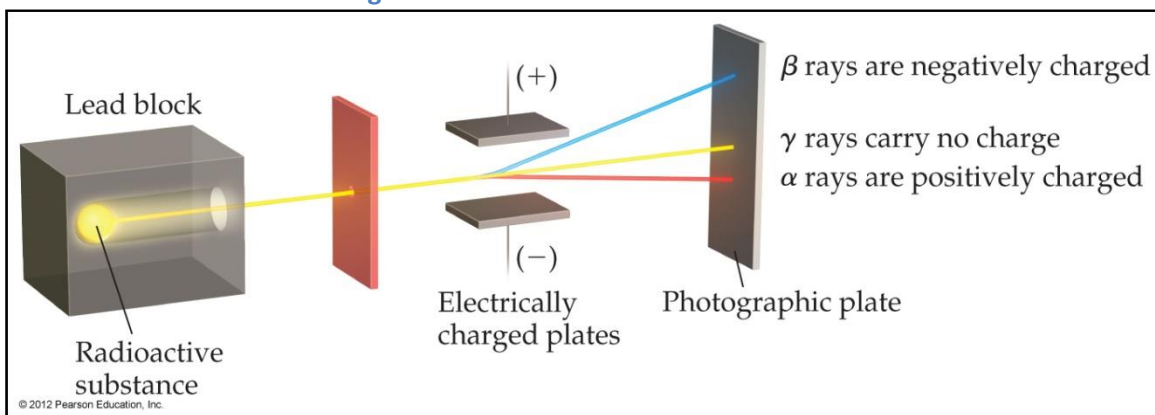


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1.10.2 Radioactivity

Radioactivity is the spontaneous emission of radiation by an atom. It was first observed by Henri Becquerel in 1895. Marie and Pierre Curie also studied it. Figure 1.11 shows the principal forms of a radioactive atom.

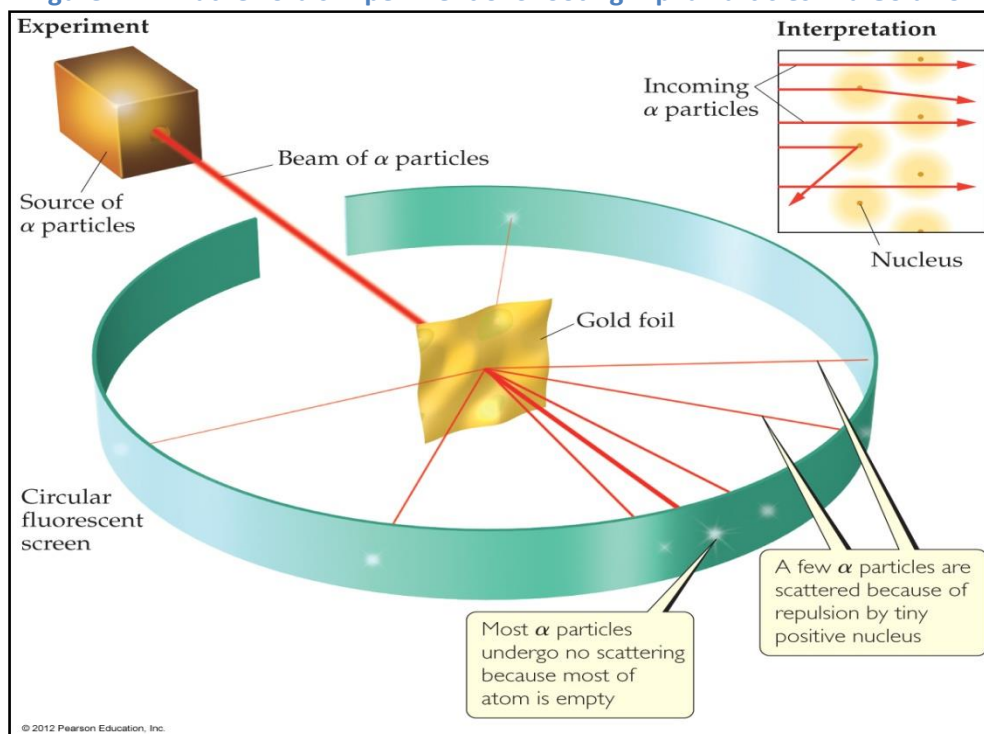
Figure 1.11: Nuclear Radiation



1.10.3 The Nuclear Model of the Atom

In 1911 Ernest Rutherford (1871-1937) put forth the idea of a nuclear model of the atom. This was based on experiments of bombarding thin gold foils with alpha (α) radiation. Most α particles passed through a metal foil as though nothing were there, but a few (about 1 in 8000) were scattered at large angles and sometimes almost backward.

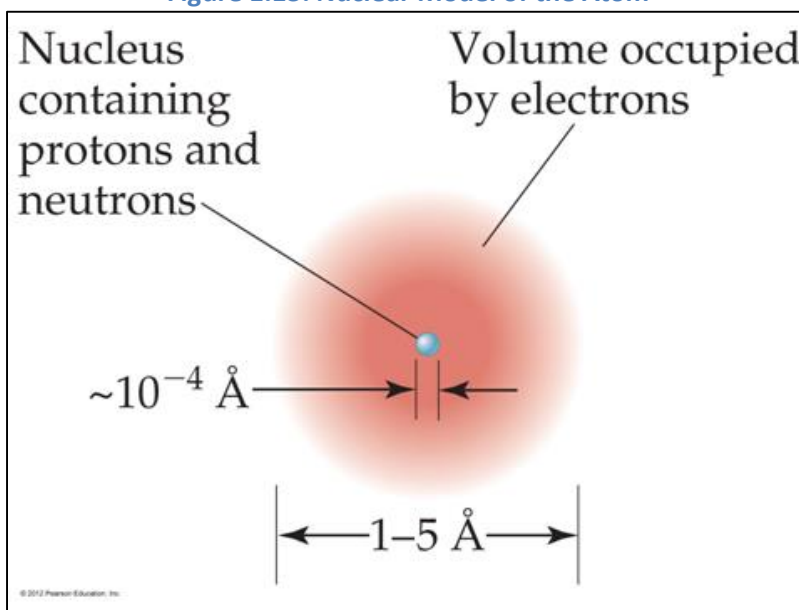
Figure 1.12: Rutherford's Experiment of Shooting Alpha Particles in a Gold Foil



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According to Rutherford's model, most of the mass of the atom ($\geq 99.95\%$) is concentrated in a positively charged centre or nucleus around which negatively charged electrons move. Nuclei have diameters of about 10^{-15} m (10^{-3} pm or 1 fm) whereas atomic diameters are about 10^{-10} m or 1Å which is 100 pm; that is, atomic diameters are a hundred thousand times larger. This situation can be conceptualized by the following example. If a golf ball represented the nucleus of an atom, the atom would be about 3 miles (or close to 5 km) in diameter.

Figure 1.13: Nuclear Model of the Atom



The nuclear model easily explains the results of bombarding gold and other metal foils with alpha particles. Protons were discovered by Rutherford in 1919 while neutrons were discovered by James Chadwick in 1932.

1.11 MODERN VIEW OF ATOMIC STRUCTURE

The nucleus of an atom also has a structure; it is composed of two different kinds of particles, protons and neutrons. An important property of the nucleus is its electric charge. The simplest way to obtain the nuclear charge is by analyzing the x-rays emitted by the element when irradiated by cathode rays. Other experiments involved α particle scattering. From these experiments, researchers discovered that each element has a unique nuclear charge that is an integer multiple of the magnitude of the electron charge. This integer, which is characteristic of an element, is called the atomic number (Z).

1.11.1 Atomic Number

A **proton** is a nuclear particle having a positive charge equal to that of the electron and a mass 1840 times that of an electron. The **atomic number** (Z) is the number of protons in the nucleus of an atom.

The proton was discovered by Rutherford in 1919 during an experiment in which α particles were used to strike nitrogen atoms according to the nuclear reaction ${}^{14}_7\text{N} + {}^4_2\alpha \rightarrow {}^{17}_8\text{O} + {}^1_1\text{p}$.

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At this point, the term atomic number can be used to give a more precise definition of an element as a substance whose atoms all have the same atomic number.

1.11.2 Mass Number

The **neutron** is a nuclear particle having a mass almost identical to that of the proton but has no electric charge. It was discovered by James Chadwick (1891-1974) during experiments in which beryllium nuclei were struck by α particles and neutrons knocked out according to the nuclear reaction ${}^9_4\text{Be} + {}^4_2\alpha \rightarrow {}^{12}_6\text{C} + {}^1_0\text{n}$.

We characterize a nucleus by its atomic number (Z) and its **mass number** (A) that is the total number of protons and neutrons.

A **nuclide** is an atom characterized by a definite atomic number and mass number. The shorthand notation of any nuclide consists of the symbol of the element with the atomic number written as a subscript on the left and the mass number as a superscript on the left. You write the nuclide symbol for the naturally occurring sodium nuclide as an example as

Mass Number and Atomic Number $\rightarrow {}^{23}_{11}\text{Na}$.

For a neutral atom, the number of electrons equals to the number of protons.

1.11.3 Isotopes

All nuclei of atoms of a particular element have the same atomic number but the nuclei may have different mass numbers. **Isotopes** are atoms whose nuclei have the same atomic number but different mass numbers; that is, the nuclei have the same number of protons but different numbers of neutrons.

Table 1:8: Some Isotopes of Carbon*

Symbol	Number of Protons	Number of Electrons	Number of Neutrons
${}^{11}\text{C}$	6	6	5
${}^{12}\text{C}$	6	6	6
${}^{13}\text{C}$	6	6	7
${}^{14}\text{C}$	6	6	8

*Almost 99% of the carbon found in nature is ${}^{12}\text{C}$

1.12 INTRODUCTION TO THE PERIODIC TABLE OF THE ELEMENTS

In 1869 Dmitri Mendeleev (1834-1907) and J Lothar Meyer (1830-1895) working independently made similar discoveries. They found that when they arranged the elements in order of atomic mass they could place them in horizontal rows, one row under the other, so that the elements in each vertical column have similar properties. A tabular arrangement of elements in rows and columns, highlighting the regular repetition of properties of the elements is called a **periodic table**.

A modern version of the periodic table arranges elements by atomic number.

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1.12.1 Periods and Groups

The basic structure of the periodic table is its division into rows and columns or periods and groups.

A **period** consists of the elements in any one horizontal row of the periodic table while a **group** consists of the elements in any one column of the periodic table.

Groups are usually numbered.

- (i) North America labels the groups with Roman numerals accompanied with the letters A and B.
- (ii) Europe uses a similar convention but some columns have the letters A and B interchanged.
- (iii) To remove this confusion, the International Union of Pure and Applied Chemistry (IUPAC) suggested a convention in which columns are numbered 1 to 18.

Using the North American convention, the A groups are called **main-group (or representative) elements**; the B groups are called **transition elements**. The two rows of elements at the bottom of the table are called **inner transition elements**. The first row of the inner transition element that starts with lanthanum is referred to as the **lanthanides** whereas the second row that starts with actinium is called the **actinides**.

As mentioned earlier, the elements in any one group have similar properties. For example Group IA are **alkali metals**, Group IIA are **alkaline earth metals**, Group VIA are **Chalcogens**, Group VIIA are **halogens** and Group VIII are **noble gases**.

1st period consists of hydrogen (H) and helium (He).

2nd period has 8 elements starting with lithium (Li) and ending with neon (Ne).

3rd period also has 8 elements. It starts with sodium (Na) and ends with argon (Ar).

4th period has 18 elements, potassium (K) to krypton (Kr).

5th period has also 18 elements, rubidium (Rb) to Xenon (Xe).

6th period actually consists of 32 elements but in order for the row to fit on a page part of it appears at the bottom of the table. Otherwise the table would have to be expanded with an additional 14 elements placed after barium (atomic number 56).

7th period though incomplete, also has some of its elements placed as a row at the bottom of the table.

1.12.2 Metals, Non-metals and Metalloids.

If one examines a tri-coloured periodic table like that found in Ebbing and Gammon or Chang, the table can be divided by a heavy and staircase line that starts with boron (B) and ends at astatine (At). On the left of the line are metals while elements on the right are non-metals.

A metal is a substance or mixture that has a characteristic lustre or shine and is generally a good conductor of heat and electricity. Except for mercury, metallic elements are solids at room

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temperature (about 20°C). They are malleable (that is, they can be hammered into sheets) and ductile (that is, they can be drawn into wire).

A **non-metal** is an element that does not exhibit the characteristics of a metal. Most of the non-metals are gases (e.g. chlorine and oxygen) or solids (e.g. phosphorus and sulphur). The solid non-metals are usually hard, brittle substances.

Most of elements bordering the staircase line in the periodic table are **metalloids** or **semimetals**. A metalloid is an element having both metallic and non-metallic properties. These elements, such as silicon (Si) and germanium (Ge) are usually good semiconductors, that is they are elements that, when pure, are poor conductors of electricity at room temperature but become moderately good conductors at high temperature.

Figure 1.14: The Periodic Table of Elements

Periods - horizontal rows

Groups - vertical columns containing elements with similar properties

Elements arranged in order of increasing atomic number

Steplike line divides metals from nonmetals

1A
1
1 H

2A
2
2 He

3A
13
5 B

4A
14
6 C

5A
15
7 N

6A
16
8 O

7A
17
9 F

8A
18
10 Ne

3
Li

4
Be

11
Na

12
Mg

13
Al

14
Si

15
P

16
S

17
Cl

18
Ar

19
K

20
Ca

21
Sc

22
Ti

23
V

24
Cr

25
Mn

26
Fe

27
Co

28
Ni

29
Cu

30
Zn

31
Ga

32
Ge

33
As

34
Se

35
Br

36
Kr

37
Rb

38
Sr

39
Y

40
Zr

41
Nb

42
Mo

43
Tc

44
Ru

45
Rh

46
Pd

47
Ag

48
Cd

49
In

50
Sn

51
Sb

52
Te

53
I

54
Xe

55
Cs

56
Ba

71
Lu

72
Hf

73
Ta

74
W

75
Re

76
Os

77
Ir

78
Pt

79
Au

80
Hg

81
Tl

82
Pb

83
Bi

84
Po

85
At

86
Rn

87
Fr

88
Ra

103
Lr

104
Rf

105
Db

106
Sg

107
Bh

108
Hs

109
Mt

110
Ds

111
Rg

112
Cn

113

114

115

116

117

118

57
La

58
Ce

59
Pr

60
Nd

61
Pm

62
Sm

63
Eu

64
Gd

65
Tb

66
Dy

67
Ho

68
Er

69
Tm

70
Yb

89
Ac

90
Th

91
Pa

92
U

93
Np

94
Pu

95
Am

96
Cm

97
Bk

98
Cf

99
Es

100
Fm

101
Md

102
No

Metals

Metalloids

Nonmetals

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1.13 MOLECULES AND MOLECULAR COMPOUNDS

In describing the composition of chemical substances using the atomic theory, we find that they exist as

- (i) atoms of noble gases
- (ii) molecules of elements or compounds
- (iii) ionic compounds.

Molecular or ionic substances are concisely described using chemical formulae that combine the symbols for the constituent elements with whole-number subscripts to show the type and number of atoms contained in the smallest unit of the substance. If we consider the formula of aluminium

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oxide, Al_2O_3 , as an example, this means that the compound is composed of aluminium atoms and oxygen atoms in the ratio 2:3.

1.13.1 Molecular Substances

A **molecule** is a definite group of atoms that are chemically bonded together. A molecular formula gives the exact number of different atoms of an element in a molecule. For example hydrogen peroxide has the formula H_2O_2 , water's formula is H_2O , ammonia's formula is NH_3 , carbon dioxide's formula is CO_2 and ethanol's formula is $\text{C}_2\text{H}_6\text{O}$.

A **structural formula** is a chemical formula that shows how the atoms are bonded in a molecule. H-O-H shows the two hydrogen atoms are bonded to oxygen in water. $\text{CH}_3\text{CH}_2\text{OH}$ is a form of a condensed structural formula of ethanol showing that the hydrogen atoms and the hydroxyl (OH) group are bonded to carbon atoms joined by a single bond.

Some elements are molecular substances and represented by molecular formulae. Examples are chlorine, Cl_2 and sulphur, S_8 .

Some elements like carbon in graphite and diamond do not have simple molecular structure but consist of very large, indefinite number of atoms bonded together. These elements are represented simply by their atomic symbols.

An important class of molecular substances is the polymers. **Polymers** are very large molecules that are made up of a number of smaller molecules repeatedly linked together. **Monomers** are the small molecules that are linked together to form the polymer.

1.14. IONIC SUBSTANCES

Although many substances are molecular, others are composed of ions. An **ion** is an electrically charged particle obtained from atoms or chemically bonded group of atoms by adding or removing electrons. Although isolated atoms are normally electrically neutral, during formation of certain compounds atoms can become ions. Metal atoms tend to lose electrons, whereas non-metals tend to gain electrons.

An atom that picks up an extra electron or more than one electron becomes a negatively charged ion, called an **anion**. An atom that loses an electron or more than one electrons becomes a positively charged ion called a **cation**. Examples of ions are a sodium cation, Na^+ , a chloride anion, Cl^- and calcium cation, Ca^{2+} .

Some ions consist of two or more atoms chemically bonded but having an excess or deficiency of electrons so that the unit has an electric charge. An example is the sulphate ion, SO_4^{2-} . The superscript 2- indicates an excess of two electrons on the group of atoms.

An **ionic compound** is a compound composed of cations and anions. Sodium chloride consists of an equal number of sodium ions and chloride ions.

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The formula of an ionic compound is written by giving the smallest possible integer number of different ions in the substance, except the charges on the ions is omitted so that the formulae merely indicate the atoms involved. For example, sodium chloride's formula is written NaCl (not Na^+Cl^-). Iron (III) sulphate is a compound consisting of iron (III) ions, Fe^{3+} , and sulphate ions, SO_4^{2-} , in the ratio 2:3. Its formula is written $\text{Fe}_2(\text{SO}_4)_3$.

Although ionic substances do not contain molecules, we speak of the smallest unit of such a substance. The **formula unit** of a substance is the group of atoms or ions explicitly symbolized in the formula. For example, the formula unit of water, H_2O is the molecule. The formula unit of iron (III) sulphate, $\text{Fe}_2(\text{SO}_4)_3$, consists of two Fe^{3+} and three SO_4^{2-} ions. The formula unit is the smallest unit of such substances.

1.15 NAMING SIMPLE COMPOUNDS

At the beginning of the study of chemistry, there was no system for naming compounds. The names give to compounds at that time are referred to as **common names**. We will begin systematic naming of compounds with naming inorganic **binary compounds**. Binary compounds are composed of two elements.

1.15.1 Binary Ionic Compounds (Type I)

Binary ionic compounds contain a positive ion (cation) always written first in the formula and a negative ion (anion). The following rules are used for naming them:

1. The cation is always named first and the anion second.
2. A monatomic (meaning "one-atom") cation takes its name from the element. For example, Na^+ is called sodium.
3. A monatomic anion is named by taking the root of the element and adding -ide. Thus the Cl^- ion is called chloride.

Some common cations and anions and their names are given in Table 1.9.

Table 1.9: Common Monatomic Cations and Anions

Cation	Name	Anion	Name
H^+	Hydrogen	H^-	Hydride
Li^+	Lithium	F^-	Fluoride
Na^+	Sodium	Cl^-	Chloride
K^+	Potassium	Br^-	Bromide
Cs^+	Caesium	I^-	Iodide
Be^{2+}	Beryllium	O^{2-}	Oxide
Mg^{2+}	Magnesium	S^{2-}	Sulphide
Ca^{2+}	Calcium	N^{3-}	Nitride
Ba^{2+}	Barium	P^{3-}	Phosphide
Al^{3+}	Aluminium		

The rules for naming binary ionic compounds are illustrated by examples in Table 1.10

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Table 10: Examples of Binary Ionic Compounds (Type I)

Compound	Ions Present	Name
NaCl	Na^+ , Cl^-	Sodium chloride
CaS	Ca^{2+} , S^{2-}	Calcium sulphide
Li_3N	Na^+ , N^{3-}	Lithium nitride
CsBr	Cs^+ , Br^-	Caesium bromide
MgO	Mg^{2+} , O^{2-}	Magnesium oxide

1.15.2 Binary Ionic Compounds (Type II)

There are metals that form more than one type of positive ion and thus form more than one type of ionic compound with a given anion. For example the compound FeCl_2 contains Fe^{2+} ions while FeCl_3 contains Fe^{3+} ions. In this case, the *charge on the metal ion must be specified*. Systematic names of above two compounds are iron (II) chloride and iron (III) chloride, respectively, where the *Roman numeral indicates the charge of the cation*.

Another system for naming these ionic compounds in the older literature *is where the ion with higher charge has a name ending in -ic and the one with lower charge has a name ending in -ous*. In this course, only systematic names should be used.

Table 1.11: Names of Common Type II Ions

Ion	Systematic Name	Older Name
Fe^{3+}	Iron (III)	Ferric
Fe^{2+}	Iron (II)	Ferrous
Cu^{2+}	Copper (II)	Cupric
Cu^+	Copper (I)	Cuprous
Co^{3+}	Cobalt (III)	Cobaltic
Co^{2+}	Cobalt (II)	Cobaltous
Sn^{4+}	Tin (IV)	Stannic
Sn^{2+}	Tin (II)	Stannous
Pb^{4+}	Lead (IV)	Plumbic
Pb^{2+}	Lead (II)	Plumbous
Hg^{2+}	Mercury (II)	Mercuric
Hg^+	Mercury (I)	Mercurous
Ag^+	Sliver [†]	
Zn^{2+}	Zinc [†]	
Cd^{2+}	Cadmium [†]	

[†]Although these are transition metals, they form only one type of ion and a Roman numeral is not used in names of their ionic compounds.

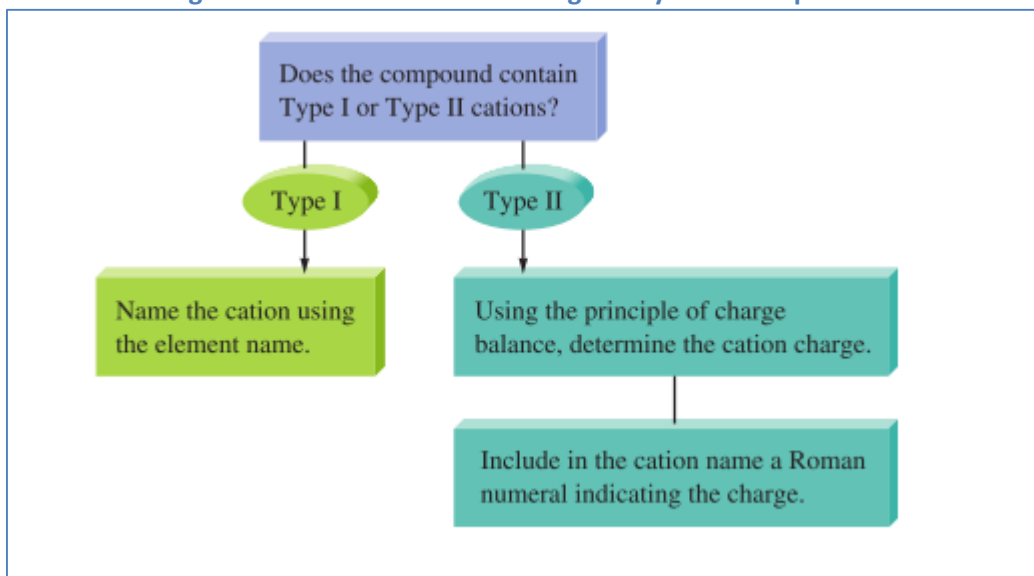
Table 1.12: Examples of Names of Ionic Compounds (Type II)

Compound	Name	Comments
CuCl	Copper (I) chloride	\because the anion is Cl^- , cation must be Cu^+ [Copper (I)] which requires a Roman numeral
HgO	Mercury (II) Oxide	\because the anion is O^{2-} , cation must be Hg^{2+} [Mercury (II)] which requires Roman numeral
Fe_2O_3	Iron (III) Oxide	The three O^{2-} ions carry a total charge of 6—, so two Fe^{3+} ions [Iron (III)] are needed to give 6+ charge

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Before you can name ionic compounds you need to be able to write and name ions. The simplest ions are monatomic ions. A monatomic ion is an ion formed from a single atom.

Figure 1.15: Flowchart for Naming Binary Ionic Compounds



It should be noted that use of Roman numerals in a systematic name is required only in cases where more than one ionic compound forms between a given pair of elements.

1.15.3 Ionic Compounds with Polyatomic Ions

Polyatomic ions contain more than one atom. These names must be memorized. Most important ones are given in Table 1.13 below

Table 1.13: Common Polyatomic Ions

Ion	Name	Ion	Name
Hg_2^{2+}	Mercury (I)	NCS^- or SCN^-	Thiocyanate
NH_4^+	Ammonium	CO_3^{2-}	Carbonate
NO_2^-	Nitrite	HCO_3^-	Hydrogen carbonate (or bicarbonate)
NO_3^-	Nitrate	ClO^- or OCl^-	Hypochlorite
SO_3^{2-}	Sulphite	ClO_2^-	Chlorite
SO_4^{2-}	Sulphate	ClO_3^-	Chlorate
HSO_4^-	Hydrogen Sulphate (common name is bisulphate)	ClO_4^-	Perchlorate
OH^-	Hydroxide	$\text{C}_2\text{H}_3\text{O}_2^-$	Acetate
CN^-	Cyanide	MnO_4^-	Permanganate
PO_4^{3-}	Phosphate	$\text{Cr}_2\text{O}_7^{2-}$	Dichromate
HPO_4^{2-}	Hydrogen phosphate	CrO_4^{2-}	Chromate
$\text{H}_2\text{PO}_4^{3-}$	Dihydrogen phosphate	O_2^{2-}	Peroxide
		$\text{C}_2\text{O}_4^{2-}$	Oxalate
		$\text{S}_2\text{O}_3^{2-}$	Thiosulphate

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Example 2.7 Naming Compounds Containing Polyatomic Ions (Zumdahl&Zumdahl 8th Ed. pp. 62)

1. Give the systematic name for each of the following compounds
2. Given the following systematic names, write the formula for each compound.

Solution

1.

Formula	Name	Comments
a. Na_2SO_4	Sodium sulfate	
b. KH_2PO_4	Potassium dihydrogen phosphate	
c. $\text{Fe}(\text{NO}_3)_3$	Iron(III) nitrate	Transition metal—name must contain a Roman numeral. The Fe^{3+} ion balances three NO_3^- ions.
d. $\text{Mn}(\text{OH})_2$	Manganese(II) hydroxide	Transition metal—name must contain a Roman numeral. The Mn^{2+} ion balances three OH^- ions.
e. Na_2SO_3	Sodium sulfite	
f. Na_2CO_3	Sodium carbonate	

2.

Name	Formula	Comments
a. Sodium hydrogen carbonate	NaHCO_3	Often called sodium bicarbonate.
b. Cesium perchlorate	CsClO_4	
c. Sodium hypochlorite	NaOCl	
d. Sodium selenate	Na_2SeO_4	Atoms in the same group, like sulfur and selenium, often form similar ions that are named similarly. Thus SeO_4^{2-} is selenate, like SO_4^{2-} (sulfate).
e. Potassium bromate	KBrO_3	As above, BrO_3^- is bromate, like ClO_3^- (chlorate).

1.15.4 Binary Covalent Compounds (Type III)

Binary covalent compounds are formed between two non-metals. Although they do not contain ions, they are named very similarly to binary ionic compounds. The following rules are used for naming them:

1. The first element in the formula is named first using the full element name.
2. The second element is named as if it were an anion.
3. Prefixes are used to denote number of atoms present. Table 1.14 shows the prefixes.
4. The prefix mono- is never used for naming the first element.

Table 1.15 shows the application of the rules for several covalent compounds of Nitrogen and Oxygen.

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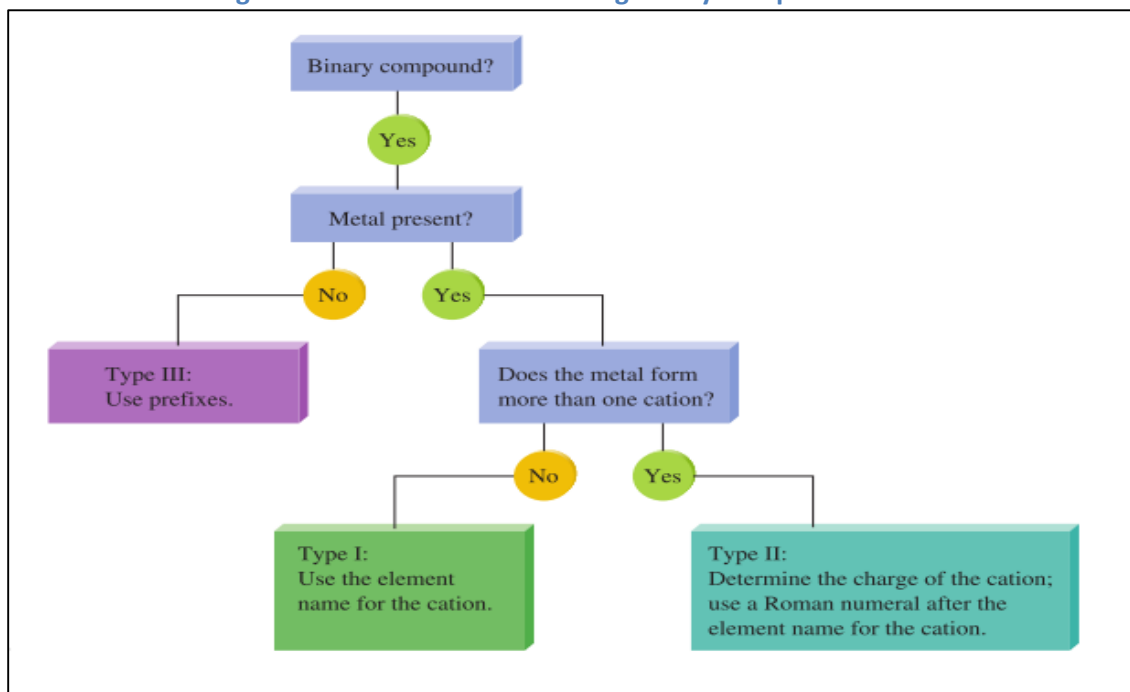
Table 1.14: Prefixes Used to Indicate Number of Chemical Names

Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
septa-	7
octa-	8
nona-	9
deca-	10

Table 1.15: Naming Several Covalent Compounds of Nitrogen and Oxygen

Compound	Systematic Name	Common Name
N ₂ O	Dinitrogen monoxide	Nitrous oxide
NO	Nitrogen monoxide	Nitric oxide
NO ₂	Nitrogen dioxide	
N ₂ O ₃	Dinitrogen trioxide	
N ₂ O ₄	Dinitrogen tetroxide	
N ₂ O ₅	Dinitrogen pentoxide	

Figure 1.16: Flowchart for Naming Binary Compounds



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In the above examples it should be noted that to avoid awkward pronunciations, the final o or a of the prefix is dropped when the element begins with a vowel. For example, N_2O_4 is called dinitrogen tetroxide NOT dinitrogen tetraoxide.

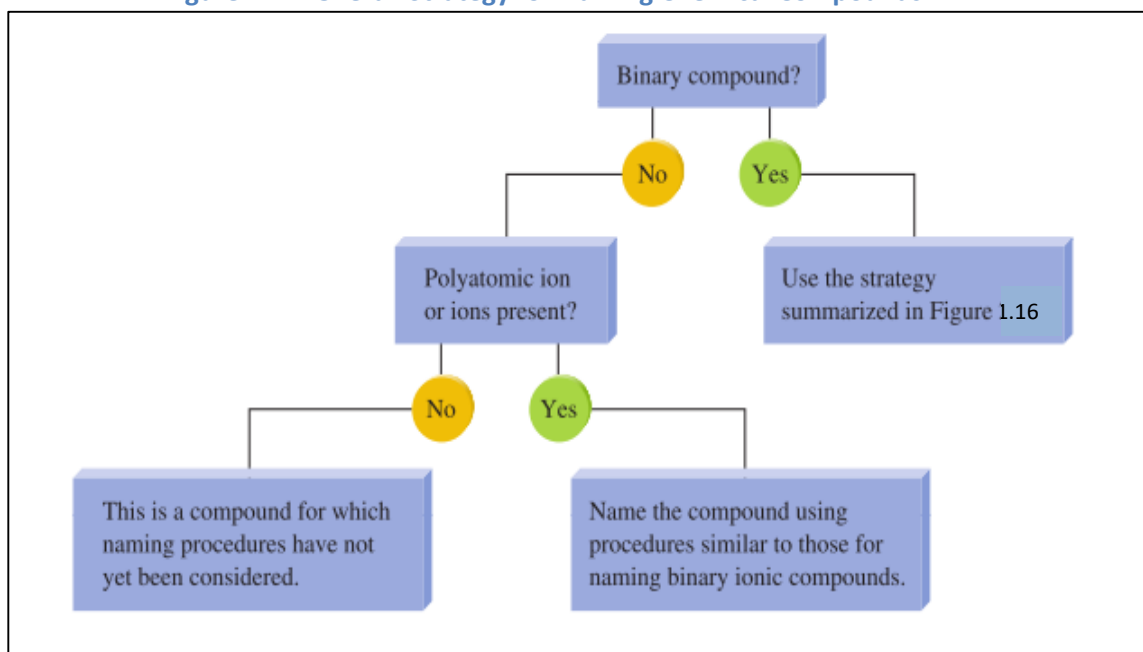
The observation is that the prefix mono is never used for the first element. For example CO is called carbon monoxide NOT monocarbon monoxide.

Finally some compounds are always referred to by their common names. Three common examples are shown in Table 1.16.

Table 1.16: Compounds with no Systematic Names

Compound	Name	Systematic NAME NEVER USED
H_2O	Water	Dihydrogen oxide
NH_3	Ammonia	Nitrogen trihydride
H_2O_2	Hydrogen peroxide	Dihydrogen dioxide

Figure 1.17: Overall Strategy for Naming Chemical Compounds



1.15.5 Names of Acids

An acid is a molecule in which one or more H^+ ions are attached to an anion. When dissolved in water acids produce a solution containing free H^+ ions (protons) and the free anions.

The rules for naming acids depend on whether the anion contains oxygen. If the name of the anion ends in *-ide*, the acid is named with the prefix *hydro-* and the suffix *-ic*. Table 1.17 shows of important acids ending in *-ide*.

When the anion contains oxygen (that is, it is an oxyanion), the acidic name is from the root name of the anion with a suffix *-ic* or *-ous* depending on the name of the anion.

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1. If the anion name ends in *-ate*, the suffix *-ic* is added to the root name. For example, H_2SO_4 contains the sulphate anion (SO_4^{2-}) and is called sulphuric acid; H_3PO_4 contains the phosphate anion (PO_4^{3-}) and is called phosphoric acid; and $\text{HC}_2\text{H}_3\text{O}_2$ contains the acetate anion ($\text{HC}_2\text{H}_3\text{O}_2^-$) and is called acetic acid.
2. If the anion has an *-ite* ending, the *-ite* is replaced by *-ous*. For example H_2SO_3 contains sulphite (SO_3^{2-}), is named sulphurous acid; and HNO_2 contains nitrite (NO_2^-), is named nitrous acid.

The application of these rules can be seen in the names of the acids of the oxyanions of chlorine in Table 1.18 while Table 1.19 shows the important oxyacids.

Table 1:178: Names of Important Acids that do not Contain Oxygen

Acid's Formula	Name of gaseous molecule of acid	Acid Name (as an aqueous solution)
HF	Hydrogen fluoride	Hydrofluoric acid
HCl	Hydrogen chloride	Hydrochloric acid
HBr	Hydrogen bromide	Hydrobromic acid
HI	Hydrogen iodide	Hydroiodic acid
HCN	Hydrogen cyanide	Hydrocyanic acid
H_2S	Hydrogen sulphide	Hydrosulphuric acid

Table 1.18: Names of Acids of the Oxyanions of Chlorine

Acid	Anion	Name
HClO_4	Perchlorate	Perchloric acid
HClO_3	Chlorate	Chloric acid
HClO_2	Chlorite	Chlorous acid
HClO	Hypochlorite	Hypochlorous acid

Table 1:20: Names of Important Oxygen-Containing Acids

Acid's Formula	Anion	Acid Name (as an aqueous solution)
HNO_3	Nitrate	Nitric acid
HNO_2	Nitrite	Nitrous acid
H_2SO_4	Sulphate	Sulphuric acid
H_3PO_4	Phosphate	Phosphoric acid
$\text{HC}_2\text{H}_3\text{O}_2$	Acetate	Acetic acid

Figure 1.18 below shows an overall strategy of naming acids.

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Figure 1.18: Flowchart for Naming Acids. An acid is best considered as one or more H^+ ions attached to an anion.

