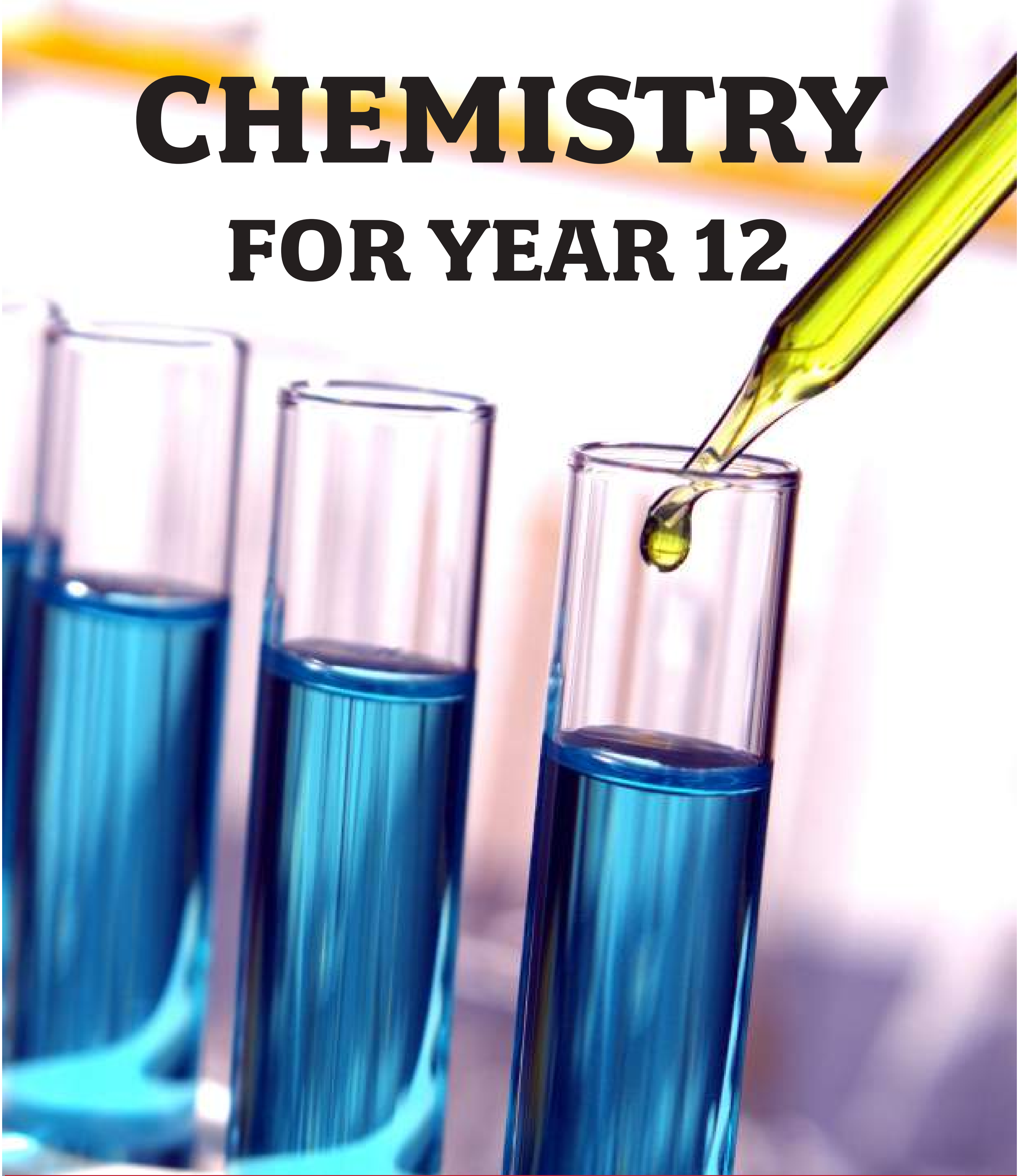
CHEMISTRY

FOR YEAR 12



Curriculum Development Unit
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PREFACE

Welcome to **Chemistry for Year 12**.

This book was prepared as a course material for Year 12 Chemistry and related activities. It covers the **Year 12 Chemistry Syllabus 2016**.

The syllabus is the framework for the learning and teaching of Chemistry. This textbook is a resource for students and teachers and it is hoped that it will be useful in implementing the syllabus and to **complement lessons prepared by teachers.**

Students and teachers are encouraged to use other resource materials as well.

Suggestions for amendments are welcome.

MINISTRY OF EDUCATION

SUVA.

23rd November, 2015

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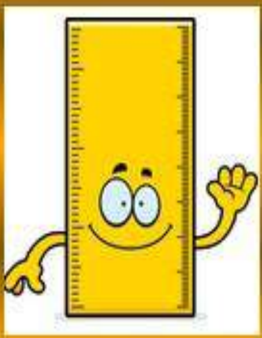
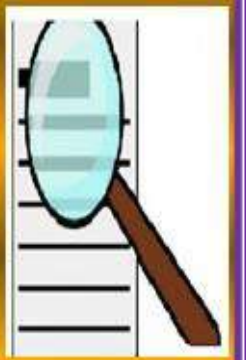


UNCERTAINTIES IN MEASUREMENTS



Strand Outcome

Demonstrate knowledge and skills on the use of uncertainty in measurements and dimensional analysis in calculations.

Sub-strands

1.1 Uncertainties in measurements
1.2 Dimensional Analysis

1.1 Uncertainties in measurements

Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Explain the occurrence of uncertainty and describe its sources in measurements.
- ✓ Distinguish between precision and accuracy in measurements.
- ✓ Distinguish between systematic and random error and describe its effect on precision and accuracy in measurements.
- ✓ Express and compare uncertainty in some common laboratory equipment.

1.1.1 Measurements and Uncertainties

Uncertainty of measurements is the doubt that exists about the result of any measurement. It tells us something about the quality of a particular measurement.

This section considers the uncertainty in experimental measurements. Here are some fundamental things that should be realised about measurements and uncertainties:

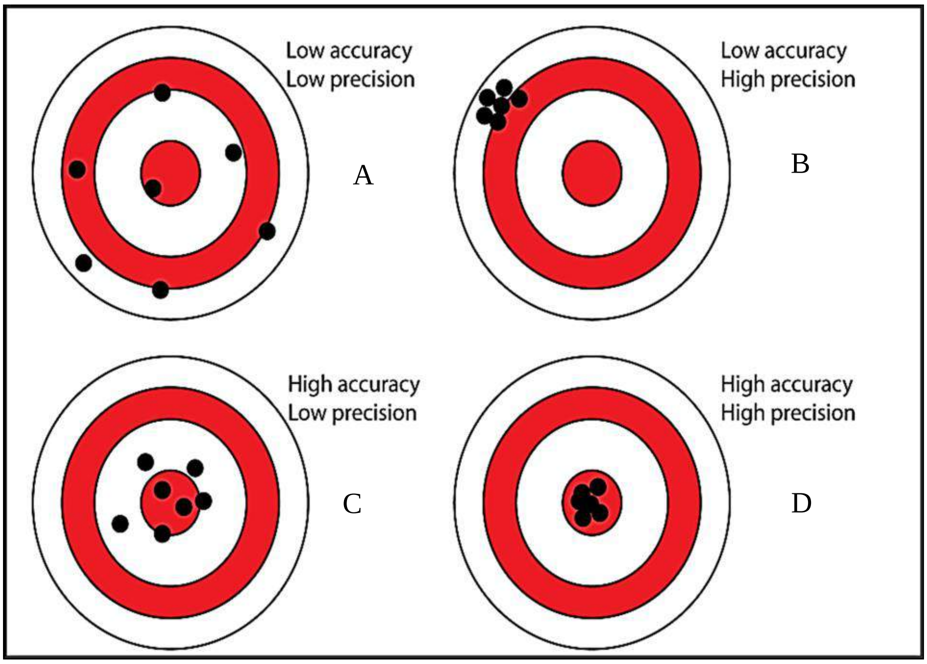
- Measurement is fundamental to science.
- In scientific work, two kinds of numbers are recognised: **exact** numbers (*those whose values are known exactly*) and **inexact** numbers (*those whose values have some uncertainty*).
- **Exact numbers** have defined values or they are integers that result from counting numbers of objects. For example, there are exactly 1000 mL in 1 L, exactly 1000 g in a kilogram, and exactly 100 cm in a meter.

  Note: The number 1 in any conversion factor between units, as in 1 m = 100 cm, is an exact number.
- Numbers obtained by measurement are always **inexact** and has an **uncertainty** or **error** associated with it. This is because there are always in-built limitations in the equipment used to measure quantities (*systematic errors*), and there are differences in how different people make the same measurement (*random errors*).
- The result of any measurement has two essential components:
  i. A **numerical value** giving the best estimate possible of the quantity measured.
  ii. The **degree of uncertainty** associated with this estimated value.
- This uncertainty can be reported either as an explicit ± value or as an implicit uncertainty, by using the appropriate number of significant figures.
- When the uncertainty is explicitly expressed, the numerical value of a ± uncertainty expresses the range of the result. For example, a titre result reported as 1.25 ± 0.05 mL means that the experimenter has some degree of confidence that the true titre value falls in between 1.20 mL and 1.30 mL.
- When significant figures are used as an implicit way of indicating uncertainty, the last digit is considered uncertain. For example, a result reported as just 1.13 implies a minimum uncertainty of ± 0.01 and a range of 1.12 to 1.14.
- Uncertainty values should only have **one significant figure**.
- Two terms are usually used in discussing the uncertainties in measured values: **accuracy** and **precision**.

i. **Accuracy** refers to the closeness of a measured value to a standard (known or true) value. For example, if in the laboratory you obtain a weight measurement of 5 g for a given substance, but the true weight is 10 g, then your measurement is not accurate. In this case, your measurement is not close to the true value.

ii. **Precision** refers to the closeness of two or more measurements to each other. Using the example above, if you weigh a given substance five times, and get 5 g each time, then your measurement is very precise. You can be very precise but inaccurate, as described above. Therefore, precision is independent of accuracy.

- The analogy of a shooter attempting to hit the bulls' eye pictured below illustrates the difference between the two terms.



Source: http://www.antarcticglaciers.org

☑ Precise hits are tightly clustered at any one place (B & D).

☑ Accurate hits fall in the bulls' eye (C & D).

☑ Accurate and precise hits are tightly clustered in the bulls' eye (D).

Note: All measurements have a degree of uncertainty regardless of precision and accuracy.

1.1.2 Types of Errors

Uncertainties in measurements are caused by systematic and random errors.

1. Systematic errors

- When tools are used for measurements, it is assumed that they are always correct and accurate.
- However, measuring tools are not always correct. Systematic errors arise from incorrect procedure, incorrect use of instruments, or failure of some value to be what it is assumed to be.
- Systematic errors are consistent in magnitude and/or direction. If the magnitude and direction of the error is known, accuracy can be improved by carrying out necessary corrections.

<u>Example</u>

i. A balance reading 0.01 g without any mass placed on it.
   - In this case, all measurements will be higher by 0.01 g to its actual mass.
   - This can be corrected by:
     - Calibrate the balance to read zero when there is no mass on it, OR
     - Subtracting 0.01 g after measuring the mass of an object.

ii. An experimenter consistently overshooting the endpoint of a titration because she is wearing tinted glasses and cannot see the first color change of the indicator.

   - This can be corrected by wearing clear glasses instead of tinted glasses.

Note:

- Systematic errors can result in high precision but poor accuracy, and usually the results cannot be averaged, even if the observations are repeated many times.
- They are frequently difficult to discover.
- Systematic errors are not random and therefore can never cancel out. They affect the accuracy but not the precision of the measurement.

2. Random errors

- Also known as human error.
- A random error is determined by the experimenter's skill or ability to perform the experiment and read scientific measurements.
- These errors are random since the results yielded may be too high or too low.
- Random errors vary from measurement to measurement.

Note:

- Often random errors determine the precision of the experiment or limit the precision.
- Unlike systematic errors, random errors vary in magnitude and direction.
- It is possible to calculate the average of a set of individual measurement values, which is likely to be more accurate than most of the individual measurements.
- Random uncertainty can be reduced, but never eliminated.

Example

- As shown on the right, a random error arises when one does not have his/her eyes exactly in level with the bottom of the meniscus when reading a burette.
- Sometimes the eyes will be above the level and the reading will be lower than the true (correct) value, while sometimes the eyes will be below the level and the reading will be higher than the true value (*usually by small amounts*).

Source: http://study.com

Note:

Systematic and random errors refer to problems associated with making measurements. ***Mistakes*** made in the calculations or in reading the instrument ***are not considered in error analysis***. It is assumed that the experimenters are careful and competent.

1.1.3 Uncertainties in measuring instruments

The following method can be used for finding the uncertainty associated with any measuring instruments.

- First find the least count (the smallest printed increment) of the measuring device.
- Most often you will feel that you are able to attain more precision than is indicated by the least count.
- In this case, estimation needs to be carried out. By estimating, the least count is divided into imaginary increments.
- For most measurements, you can divide the least count into 5 imaginary increments. This is called the one-fifth rule.
- There will be some measurements where the one-fifth rule will be too general. Therefore, if you feel that your confidence in the last significant digit of the measurement is greater than

this, you can divide the least count into 10 imaginary increments. This is called the one-tenth rule.

- Similarly, if you feel that your confidence in the last significant figure is lower than this value, than you can divide the least count by half.
- A good judgment must be made when estimating the error.
- In measurements, you must always try be as accurate and precise as possible. A good way to ensure this is to have the lowest possible value for your uncertainty and this can be achieved by using the one-tenth rule.

Important

Two rules can be followed to ensure that correct uncertainty has been obtained.

Rule 1: The uncertainty must be to the same decimal place value as the measured value.

Rule 2: The measured value should be evenly divisible by the uncertainty.

- For digital instruments, the uncertainty is equal to the smallest digital increment and no estimation will be performed. For instance, a top pan balance reading may be represented as 10.25 ± 0.01 g.
- It should also be noted that the greater the number of significant figures a measurement has, the more precise it is.

Example 1

What is the difference between 2.0 g and 2.00 g?

Solution

- For many people, there is no difference but for a scientist, there is difference in the number of significant figures in the two measurements.
- The value 2.0 has two significant figures, while 2.00 have three significant figures.
- This indicates that the measurement 2.00 g is more precise.
- A measurement of 2.0 g indicates that the mass is between 1.9 g and 2.1 g. Thus the mass is 2.0 ± 0.1 g.
- A measurement of 2.00 g indicates that the mass is between 1.99 and 2.01 g. Thus the mass is 2.00 ± 0.01 g.

Example 2

Give the reading (in mL) of the following measuring apparatus with the correct uncertainty.

- The smallest division in this apparatus is 0.1.
- The meniscus of the liquid in the apparatus is between 10.20 and 10.30. Thus the reading of the meniscus of the liquid can be estimated to be 10.25.
- The one-tenth rule can be used to find the uncertainty value.
- Therefore, the uncertainty is 0.1 divided by 10, which is 0.01.
- The final measurement can be written as **10.25 ± 0.01 mL**.

Uncertainity

10.25 ± 0.01 mL

Measured

Best estimate

Recording in this manner show the number you are certain of (10.2 mL), the number you think is the best estimate (0.05 mL), and the range within which you are certain the number falls (0.01 mL), called the uncertainty or range of the reading.

<u>Example 2</u>

Suppose your Chemistry teacher requested you to measure the volume of a quantity of liquid. You are given two different glasswares in which you should make the measurement - a 50 mL graduated cylinder and a 50 mL burette. You are required to use both the glasswares and get the volume of the liquid with their uncertainties.

(a) <u>50 mL graduated cylinder</u>

Source: http://chem.wisc.edu

25.2 ± 0.1 mL

- The smallest division in this apparatus is 1.
- The meniscus of the liquid in the cylinder is between 25 and 26. The volume can be estimated to be 25.2 since it is not exactly between 25 and 26 but is much closer to 25.
- The one-tenth rule can be used to find the uncertainty value.
- Therefore, the uncertainty is 1 divided by 10, which is 0.1.
- The final measurement can be written as:

  25.2 ± 0.1 mL.

(b) <u>50 mL burette</u>

Source: http://chem.wisc.edu

25.18 ± 0.01 mL

- The smallest division in this apparatus is 0.1.
- The meniscus of the liquid in the cylinder is between 25.1 and 25.2. The reading can be estimated to be 25.18 since it is not exactly between 25.1 and 25.2 but is much closer to 25.2.
- The one-tenth rule can be used to find the uncertainty value.
- Therefore, the uncertainty is 0.1 divided by 10, which is 0.01.
- The final volume can be written as:

  25.18 ± 0.01 mL.

<u>Summary</u>

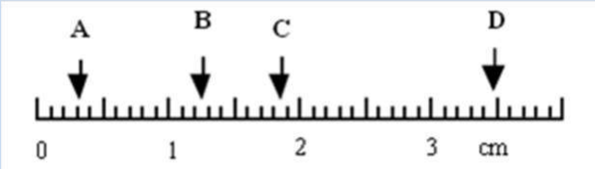
- Uncertainties always exist in measured quantities.
- Uncertainty gives a measure of the precision of an instrument.
- The uncertainty of a measurement should be given with the actual measurement, for example, 40.74 ± 0.05 mL.
- A measured result is only complete if it is accompanied by a statement of the uncertainty in the measurement.
- Measurement uncertainties can come from the measuring instrument, from the item being measured, from the environment, from the operator, and from any other sources.
- The use of good practice can reduce measurement uncertainties.
- Each instrument has an inherent amount of uncertainty in its measurement. Even the most precise measuring instrument available will not be able to give the actual value because to do so would require an infinitely precise instrument.

In ancient times, the body ruled when it came to measuring. The length of a foot, the width of a finger, and the distance of a step were all accepted measurements.

Source: http://www.factmonster.com

Exercise

1. Consider the figure on a measuring instrument given below.



Using this figure, write the measured value and uncertainties for the readings at A, B, C and D.

2. In the laboratory, a student uses a top pan balance to find the mass of a small object. She tells you correctly that the digital readout shows 44.15 g. Write the measurement with its uncertainty.

3. State whether you agree or disagree with the statements below:

i. If a measurement result has uncertainty, it means that the measurement was not performed correctly.

ii. A measured value can be regarded as an estimate of the true value.

iii. The estimated value can be corrected to some extent by adding or subtracting the measurement uncertainty from it.

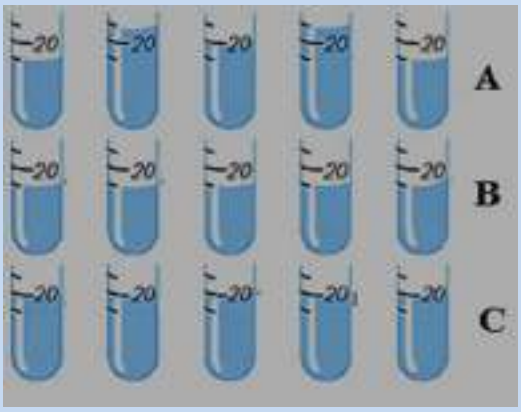
iv. Measurement uncertainty characterises the accuracy of the measurement result.

4. Distinguish between:

   i. Accuracy and precision in measurements.

   ii. Systematic and random errors with suitable examples.

5. To test their measuring skills, a teacher gave three students the same volume of water (20 mL) to measure. To get to the best result, each student decided to measure the same volume of water 5 times. The results obtained from the students are as follows:



Which measurement (A, B or C) has:

i. High precision but low accuracy.

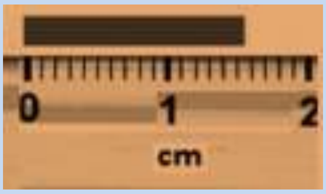
ii. High accuracy and high precision.

iii. Low precision but low accuracy.

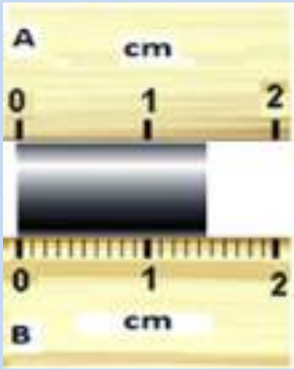
6. Explain if it is possible for a measurement to be accurate but not precise?

7. When a student was measuring a given amount of water from a measuring cylinder, she distorted the cylinder and many of the readings done was estimated by the experimenter. What is the random error and the systematic error in this case?

8. What is the length with its uncertainty of the magnesium ribbon shown in the diagram below?

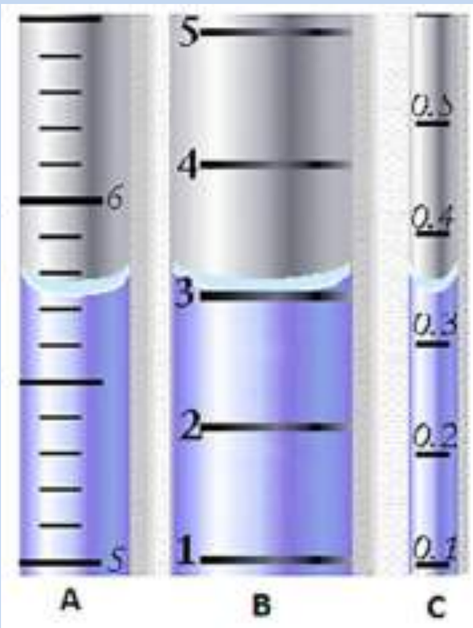


9. Study the diagram given below which shows a piece of magnesium ribbon being measured by two rulers (A and B) and answer the question that follows.

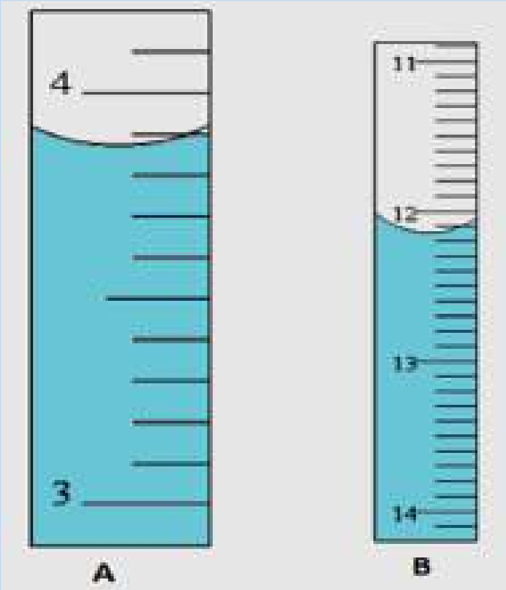


Give the correct length measurements with their uncertainties of the magnesium ribbon for rulers A and B. If you were given a choice, which ruler would you choose to use? Give a reason for your answer.

10. Give the reading (in mL) with their uncertainties of the three measuring cylinders shown below.



11. Give the reading (in mL) from a graduated cylinder (A) and a burette (B) given below with their uncertainties.



Source: http://www.csudh.edu

12. Which of the following quantities can be determined exactly?

A. The number of light switches in the room you are sitting in now.
B. The number of meters in a kilometer.
C. The number of stars in the sky.
D. The number of milliliters in a liter.
E. The number of red blood cells in exactly one litre of blood.

12. A sample of liquid has a measured volume of 24.15 mL. Assume that the measurement was recorded properly.

i. How many significant figures does the measurement have?

ii. Assume the volume measurement was made with a graduated measuring cylinder. How far apart were the scale divisions on the measuring cylinder, in mL?
A. 10 mL B. 1 mL C. 0.1 mL D. 0.01 mL

iii. Which of the digit(s) in the measurement is uncertain?
A. 2 B. 4 C. 1 D. 5

13. For each of the following mistakes during the density of water experiment, determine if the mistake would give a high or low density of water. ***(Assume everything else in the experiment was done correctly).***

A. When weighing a volumetric flask full of water, a student did not notice that the outside of the flask was wet.
B. When weighing an empty beaker, a student did not notice that there was some water inside.
C. Before placing an empty 10 mL graduated cylinder on the balance to get its mass, a student did not notice that the balance read "-1.44 g". When the student weighed the full 10 mL graduated cylinder later in the experiment, she correctly set the balance to "0.00 g".

1.2: DIMENSIONAL ANALYSIS

Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Describe some common units of measurements.
- ✓ Convert a quantity expressed with a single unit of measurement to an equivalent quantity with another single unit of measurement.
- ✓ Convert a quantity expressed as a rate to another rate.

- Dimensional analysis is a problem solving technique that uses the correct cancellation of the units of physical quantities to get the correct set up of the solution to the problem.
- Dimensional analysis will help ensure that the solution to problems yield the proper units.
- It allows for a unit to be converted to another unit of the same dimension.

  For example:

  i. *We can convert time measured in seconds into another unit of time, such as minutes, because we know that there are always 60 seconds in one minute.*

  ii. *1 kilogram of pure water is equal to 1 litre. Using that knowledge, you could convert a volume of pure water to a mass of water, or vice versa.*
- With a known **conversion factor**, it is sometimes possible to convert to a new dimension.

<u>Metric Conversions</u>

Common Metric Conversions			
Prefix	**Symbol**	**Exponential base units (in given units)**	**Base units (in given units)**
Giga	G	10^{9}	1,000,000,000
Mega	M	10^{6}	1,000,000
Kilo	k	10^{3}	1,000
Hecto	h	10^{2}	100
Deca	da	10^{1}	10
Base Unit	-	10^{0}	1
Deci	d	10^{-1}	0.1
Centi	c	10^{-2}	0.01
Milli	m	10^{-3}	0.001
Micro	μ	10^{-6}	0.000001
Nano	n	10^{-9}	0.000000001

Conversion factors commonly used in Chemistry

1 cm^3	1 mL
1 dm^3	1 L
1 kg	1000 g
1 g	1000 mg
1 atm	101.3 kPa
1 atm	760 mmHg
1 mole	6.02×10^{23} particles (atoms, molecules, ions)
1 L	1000 cm^3
1 L	1000 mL
1000 cm^3	1 dm^3
1 x 10^6 cm^3	1 m^3
1 x 10^{15} cm^3	1 km^3
1 μm	1×10^{-6} m
1 cm	1×10^{-2} m
1 cm^3	1000 mm^3
1 cm^3	1×10^{12} $μm^3$
1 km	1000 m
1 m	100 cm
1 m	1000 mm
1 cm	10 mm

Note:

- Any conversion factor can be written in two ways:

  For example: 1 L = 1000 mL can be written as:

$$\frac{1\text{L}}{1000\text{ mL}} \quad \text{or} \quad \frac{1000\text{ mL}}{1\text{L}}$$

  - ✓ The first of these factors is used when we want to convert cm^3 to liters.
  - ✓ The second factor is used when we want to convert liters to cm^3.

- Therefore, the general formula that can be used for most conversions is:

$$\text{Desired unit} = \text{Given unit} \times \frac{\text{Desired unit}}{\text{Given unit}}$$

For other conversions:

- Dimensional analysis begins with finding the appropriate conversion factors.
- Then simply multiply the values together such that the units cancel by having equal units in the numerator and the denominator.
- Units that are on the top and bottom of an expression cancel out.

Solved Examples

1. Convert 25.0 mL to litres.

Solution

$$\text{Desired unit} = \text{Given unit} \times \frac{\text{Desired unit}}{\text{Given unit}}$$

$$= \frac{25.0\ \text{mL}}{1} \times \frac{1\ \text{L}}{1000\ \text{mL}} = \frac{25\ \text{L}}{1000}$$

$= \underline{\mathbf{0.025\ L}}$

1 L = 1000 mL

2. Convert a density of 2.16 g/mL to its equivalent in g/L.

Solution

$$\frac{2.16\ \text{g}}{\text{mL}} \times \frac{1000\ \text{mL}}{1\ \text{L}}$$

1 L = 1000 mL

$= \underline{\mathbf{2160\ g/L}}$

3. What is the mass in kg of a 2 L liquid which has a density of 13.6 g/mL?

Solution

For such problems, set up the calculation so that it will yield a result with a mass in grams.

$$\frac{13.6\ \text{g}}{\text{mL}} \times \frac{1000\ \text{mL}}{1\ \text{L}} = \frac{13600\ \text{g}}{\text{L}}$$

1 L = 1000 mL

$$\frac{13600\ \text{g}}{\text{L}} \times 2\ \text{L} = 27200\ \text{g}$$

$$\frac{27200\ \text{g}}{1} \times \frac{1\ \text{kg}}{1000\ \text{g}} = \underline{\mathbf{27.2\ kg}}$$

1 kg = 1000 g

Note:

Doing such conversions becomes very important when you forget a formula and also to ensure you get to the correct unit.

However, there can be other methods to get to the right answer. This includes using ratios and formulas.

Hint: Use the formula: $\rho = m / V$ to solve Example 3.You should get to the same answer, provided you do the right conversions at the right places!

4. A drug container has a label which gives the concentration of the drug it contains as 1200 mg/mL. Determine the volume (in mL) of the medication that must be given in accordance to a prescription which required 1600 mg of the drug.

Note: In this question, a conversion factor is not necessary since the unit required is the same as in the question. Thus ratio (cross-multiply method) can be used.

$$1200 \text{ mg} : 1 \text{ mL}$$
$$1600 \text{ mg} : \text{X}$$

$$= 1200 \text{ mg} \times \text{X} : 1600 \text{ mg} \times 1 \text{ mL}$$

$$= \frac{1200 \cancel{\text{mg}} \text{X}}{1200 \cancel{\text{mg}}} : \frac{1600 \cancel{\text{mg}} \text{mL}}{1200 \cancel{\text{mg}}}$$

$$\text{X} = \frac{1600 \text{ mL}}{1200}$$
$$= \mathbf{1.33 \text{ mL}}$$

Exercise

1. Do the following conversions:
   i. 15 mm to its equivalent in m.
   ii. 8.2 micrometers to its equivalent in mm.
   iii. Pressure of 7 atm to its equivalent in kPa.
   iv. Pressure of 2280 mmHg to its equivalent in atm.
   v. A density of 9.46 g/L to its equivalent in g/mL.
2. The density of water is 1.00 g/mL. What is the density of water in kg/L?
3. A chemistry student requires 250 milligrams of a chemical for a particular experiment. She has 30 grams of the chemical. How many times can the student carry out the experiment?
4. The density of table salt is 2.16 g/mL. What is the mass of 2 L of this salt?
5. If 61.2 g of sodium chloride is contained in 1 L solution of sodium chloride, how many grams of sodium chloride is contained in 2.75 L of this solution?
6. A sample of sea water contains 6.28 g of sodium chloride per litre of solution. How many milligrams of sodium chloride would be contained in 15.0 mL of this solution?

Strand Outcome

Demonstrate an understanding of the differences in structures and properties of the types of matter.

2.1 Atomic Structure and Bonding

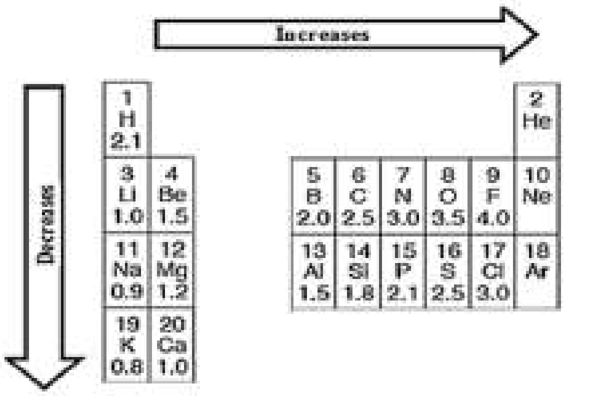
Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Describe and explain the trends of electronegativity of elements in the Periodic Table.
- ✓ Draw Lewis structures of some common molecules.
- ✓ Describe and explain the geometry of molecules using the VSEPR theory.
- ✓ Explain the effect of electronegativity on bond polarity.
- ✓ Explain the polar and non-polar nature of some covalent compounds.
- ✓ Explain and compare the properties of different types of solids based on their structure and bonding.
- ✓ Describe and explain the hydration of ionic solids in water.

2.1.1 Electronegativity

- Electronegativity is a measure of the strength of an atom to attract the shared pair of electrons in forming a covalent bond.
- The Pauling scale is most commonly used as a measure of electronegativity of elements. For example, fluorine is the most electronegative element and is assigned a value of 4.0.
- The trend in electronegativity is such that it increases across the period from left to right and decreases down the group for the first twenty elements.



Electronegativity values of the first twenty elements of the periodic table

Reasons for the observed trends:

- Across the period, electrons are being added to the same energy level of the atom. This increases the nuclear charge since there is now greater attraction between the nucleus and the shared pair of electrons in a bond. This results in increasing electronegativity across the period.
- Down the group, there is addition of new energy levels which increases the distance of the positively charged nucleus to the valence electrons. This means that the inner electrons now shield the outer electrons more. This increasing shielding effect down the group leads to decreasing attraction for the shared pair of electrons in a bond. This results in decreasing electronegativity down the group.

Exercise

1. What do you understand by the term 'electronegativity'?
2. Describe the trend of electronegativity across the period and down the group for the first twenty elements in the Periodic Table.
3. Briefly explain why:
   i. Oxygen is more electronegative than lithium.
   ii. Potassium is less electronegative than lithium.

2.1.2 Lewis Structure Diagrams

Lewis Structures

- Lewis structures are visual representations of the bonds between atoms and illustrate the lone pairs of electrons in molecules.
- Lewis structures are useful because they show which atoms are bonded together and whether any atom possesses lone pair of electrons or not.

Lewis Theory

The Lewis theory is based on the following principles:

1. Valence electrons take part in chemical bonding.
2. Ionic bonds are formed between atoms when electrons are transferred from one atom to another. Ionic bond is a bond between non-metals and metals.
3. Covalent bonds are formed between atoms when pairs of electrons are shared between atoms. A covalent bond is formed between two non-metals.
4. Electrons are transferred or shared so that each atom reaches a more stable electron configuration i.e. the noble gas configuration. This is called the **octet rule.** The octet rule states that there should be 8 valence electrons in the outermost or valence shell of an atom. Note that those atoms which can only have a maximum of 2 electrons in its outermost or valence shell (e.g. hydrogen) is exempted from the octect rule.

Constructing Lewis Structures

Steps

1. Calculate the total number of valence electrons available for bonding by adding up all the valence electrons of individual atoms in the molecule.
2. Identify the central atom and place all other atoms around the central atom.

> **Note:**
> The least electronegative atom in the molecule is usually the central atom. The exceptions are hydrogen and halogens, which are normally at the perimeter of the molecule**.**

3. Place a pair of electrons between all the atoms to form a bond.
4. Distribute the remaining valence electrons to the atoms surrounding the central atom (or atoms) to satisfy the octet rule for these surrounding atoms as follows:

- Subtract the number of electrons used to form the bonds from the total valence electrons from Step 1.
- Distribute the remaining electrons as pairs to all the atoms to satisfy the octet rule.

- If the octet rule is still not satisfied on an atom, it suggests that a multiple bond is present *(Two electrons fewer than an octet suggest a double bond; four fewer than an octet suggests a triple bond or two double bonds.)*
- To obtain a multiple bond, move one or two electron pairs (*depending on whether the bond is to be double or triple*) from a surrounding atom to the bond connecting the central atom.

Note: Atoms that often form multiple bonds are C, N, O and S.

<u>Example 1</u>

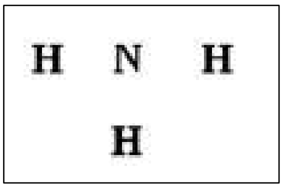
Ammonia (NH_3)

1. Find the total number of valence electrons in the molecule.

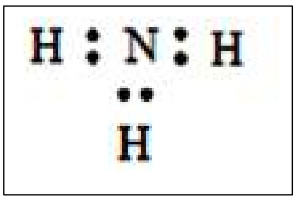
NH_3

1 Nitrogen atom → 1 x 5 valence electrons	= 5 valence electrons
3 Hydrogen atoms → 3 x 1 valence electron	= 3 valence electrons
Total	= 8 valence electrons **or** 4 electron pairs

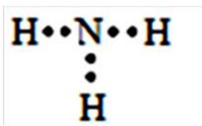
2. Identify the central atom.
   For NH_3, the central atom will be nitrogen (N).
   Note that hydrogen atoms will be at the perimeter of the molecule.



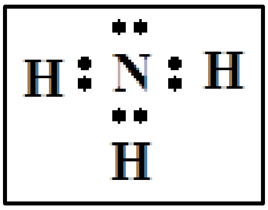
Place a pair of electrons **between** all the atoms to at least form a bond.



Note: The electrons that are shared between two atoms needs to be drawn as shown above and **<u>NOT</u>** like this:



3. Determine which atoms need an octet.
   Out of the four available pairs, 3 pairs of electrons are used up and only one pair is left. Place this electron pair around the nitrogen atom because it has 6 electrons and needs 2 more electrons to gain an octet. The three hydrogen atoms have 2 electrons each and hence have fully filled shells or stable configurations. All the electrons or electron pairs as determined in Step 1 are used up and all the atoms have attained octet, hence the correct Lewis structure has been obtained for NH_3.



<u>Example 2</u>

Carbon dioxide (CO_2)

1. Find the total number of valence electrons in the molecule.

CO_2

1 Carbon atom → 1 x 4 valence electrons	= 4 valence electrons
<u>2 Hydrogen atoms → 2 x 6 valence electrons</u>	<u>= 12 valence electrons</u>
Total	= 16 valence electrons **or** 8 electron pairs

2. Identify the central atom.

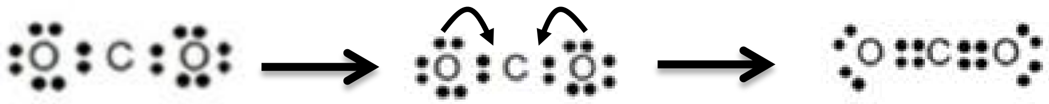
Carbon is the least electronegative atom in the molecule and will be the central atom. The two oxygen atoms will be at the perimeter of the molecule.

O C O

3. Place a pair of electrons **between** all the atoms to form a bond.

4. Determine which atoms need an octet.

Out of the eight available pairs, 4 pairs of electrons are used up and 4 pairs are left. Place these electron pairs around the more electronegative atom to obtain an octet. Since there are two atoms of oxygen, two pairs of electrons can be placed around each of them to obtain octets. This uses up all valence electrons but the carbon atom is still without an octect. This implies that multiple bonds are a possibility. By moving a lone pair of electron from each of the oxygen atom to the form a double bond between carbon and oxygen atom, the octect can be satisfied for all the atoms. All the electrons or electron pairs as determined in Step 1 are used up and all the atoms have attained octet, hence the correct Lewis structure has been obtained for CO_2.



Note: Lewis structures can also be represented in the following forms:

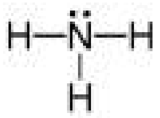
(i) The bonded pair of electrons is shown using horizontal or vertical lines and lone pairs of electrons as dots.

(ii) The bonded pair of electrons is shown using horizontal or vertical lines and the lone pairs of electrons are not explicitly shown but assumed to be present. These are also known as Kekule structures.

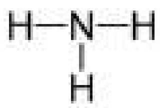
Example

NH_3

(i)

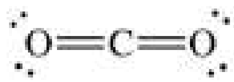


(ii)



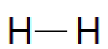
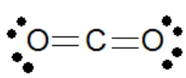
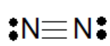
CO_2

(i)



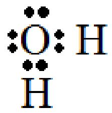
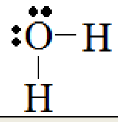
(ii)

Summary

Single Bonds	**Double Bonds**	**Triple Bond**
➢ A single bond is when two electrons (one pair of electrons) are shared between two atoms. ➢ It is depicted by a single horizontal line between the two atoms.	➢ A double bond is when two atoms share two pairs of electrons with each other. ➢ It is depicted by two horizontal lines between two atoms in a molecule.	➢ A triple bond is when three pairs of electrons are shared between two atoms in a molecule. ➢ It is depicted by three horizontal lines between two atoms in a molecule.
Example: Hydrogen (H_2) 	Example: Carbon dioxide (CO_2) 	Example: Nitrogen (N_2) 

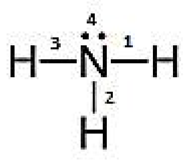
Exercise

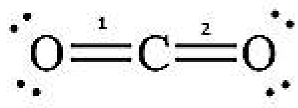
<u>Copy and complete the table below</u> *(The first one is done for you!)*

Name of molecule	Molecular formula	Lewis Structure	
		Electron Dot Diagram	**Line Representation**
Water	H_2O		
Hydrogen	H_2		
Nitrogen	N_2		
Oxygen	O_2		
Chlorine	Cl_2		
Hydrogen sulphide	H_2S		
Nitrogen trichloride	NCl_3		
Methane	CH_4		
Tetrachloromethane	CCl_4		

2.1.3 Shapes of Molecules

- The arrangement of electron groups about the central atom of a molecule is called its **electron group geometry**.
- Electron group geometry of each central atom in a molecule can be determined by simply counting the number of "groups" of electrons around the central atom in the Lewis structure diagram (bonded pairs and lone pairs).
- A "group" of electrons can be a single bond, double bond, triple bond, or a lone pair of electrons.
- For example, in NH_3, there a four electron groups; 3 bonded pair groups and 1 lone pair group. In CO_2, there are two electron groups; two double bonded pair groups. These are labelled in the diagram below.





- Two electron groups are arranged **linearly**, three groups are arranged in a **trigonal planar** fashion and four are arranged **tetrahedrally**.
- The table below summarises the terms given to the electron group geometry of a molecule depending on the number of electron groups.

<u>Electron- Group geometry as a function of the number of electron groups</u>

No. of electron groups	Electron group geometry	Predicted bond angle
2	Linear	180^o
3	Trigonal planar	120^o
4	Tetrahedral	109.5^o

- The shape of a molecule can be related to these three basic arrangements of electron groups.
- When experiments are used to determine the shape of a molecule, the atoms are located, not the electron groups.
- The **molecular geometry** (shape) of a molecule is the arrangement of the atoms in space.
- The molecular geometry of a molecule can be predicted from its electron-group geometry.
- Since electron groups are negatively charged regions, they repel one another. The repulsive forces between bonding and non-bonding electrons determine geometry of the groups of electrons around a central atom. As the negative charges repel one another, the

electron groups arrange themselves around the central atom so that they are as far apart as possible.

- Valence Shell Electron Pair Repulsion (VSEPR) theory states that the pairs of electrons that surround the central atom of a molecule or ion are arranged as far apart as possible to minimize electron-electron repulsion. The VSEPR theory can be used to predict the shapes of molecules by following a simple procedure.
- Firstly note that the order of repulsion of electron groups is:

  Lone pair - lone pair > lone pair - bonded pair > bonded pair - bonded pair
- Therefore, the VSEPR theory is used to predict the shape of molecules by considering repulsions between pairs of electrons (bonding pairs and lone pairs).

Note: **For molecules, Lewis structures predict the number and types of bonds, whereas VSEPR theory predicts the shapes of molecules.**

The following steps are used to predict molecular geometries with the VSEPR model:

1. Sketch the Lewis structure of the molecule or ion.

2. Count the total number of electron groups around the central atom and work out the electron group geometry. Arrange the electron groups around the central atom in the way that minimizes the repulsions among them. Note that the electrons in a multiple bond constitute a single electron group, just like a single bond does. For example, the Lewis structure of CO_2 () has two C=O double bonds and CO_2 has two electron groups.

3. Identify the type of interactions presents (Lone pair- lone pair, lone pair-bond pair, bond pair-bond pair**)** and predict deviations from electron group geometry.

4. Describe the molecular geometry in terms of the angular arrangement of the bonded atoms.

<u>Let's look at an example:</u>

<u>Ammonia (NH_3)</u>

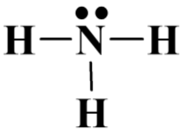
- The Lewis structure of NH_3 (see below) shows a central nitrogen atom surrounded by four separate regions of high electron density. Three of these regions consist of a single pair of electrons forming a covalent bond with a hydrogen atom; the fourth region contains an unshared pair of electrons.
- According to the VSEPR theory, the repulsions in the molecule are minimized by directing three hydrogen atoms and the lone pair to the four corners of a tetrahedron.
- The four regions of high electron density around the nitrogen are arranged in a tetrahedral manner, so we predict that each H - N - H bond angle should be 109.5°. However, the observed bond angle is 107.3°. This small difference between the predicted angle and the

observed angle can be explained by noting that the unshared pair of electrons on nitrogen repels the adjacent bonding pairs more strongly than the bonding pairs repel each other.

- Note that ammonia is not a tetrahedral molecule, i.e. the shape of the ammonia molecule is actually trigonal pyramid. The atoms of ammonia form a pyramidal molecule with nitrogen at the peak and the hydrogen atoms at the corners of a triangular base. The unshared pair of electrons in ammonia contributes to the tetrahedral electron group geometry of the molecule but is not included in the description of its molecular geometry (shape). Its molecular geometry (shape) is based only on the arrangement of atoms, which is trigonal pyramidal.

NH_3

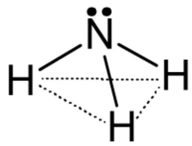
Molecular formula



Lewis Structure

Electron group geometry

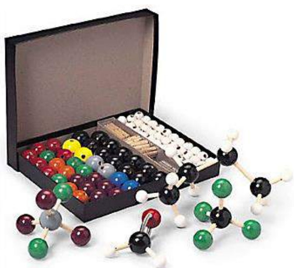
(Tetrahedral)



Molecular geometry (shape)
Trigonal pyramidal

Note:
The illustration above shows that the shape of NH_3 is predicted by first writing the Lewis structure, then using the VSEPR theory to determine the electron-group geometry, and then finally focusing on the atoms themselves to describe the molecular geometry.

It was seen that the trigonal pyramidal molecular geometry of NH_3 is a consequence of its tetrahedral electron group geometry. Thus, ***when describing the shapes of molecules, the molecular geometry or shape is given rather than the electron group geometry.***



Source: www.meta-synthesis.com

Important: This concept is best taught and learnt using models.

Source: http://www.apqua.org

Molecular models of some molecules:

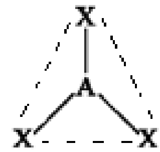
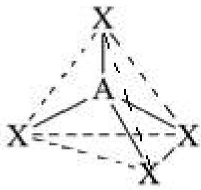
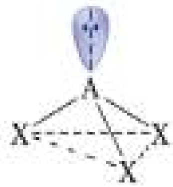
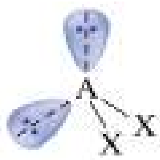
H_2O

BH_3

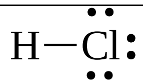
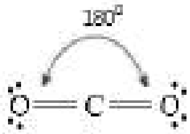
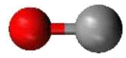
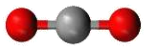
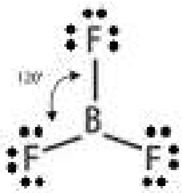
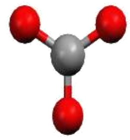
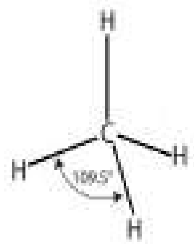
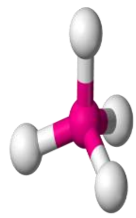
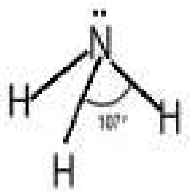
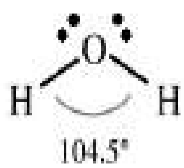
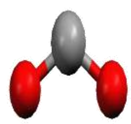
NH_3

Source: https://ruthlearns.wordpress.com

The table below summarizes the possible molecular geometries when there are four or fewer electron groups about the central atom in a molecule (**A** is the central atom).

No. of Electron Groups	**No. of Lone pairs of Electrons**	**No. of Bonded Pairs**	**Electron Group Geometry**	**Molecular Geometry (Shape)**	**General example**
2	0	2	Linear	Linear	X—A—X
3	0	3	Trigonal Planar	Trigonal Planar	
4	0	4	Tetrahedral	Tetrahedral	
4	1	3	Tetrahedral	Trigonal Pyramidal	
4	2	2	Tetrahedral	Bent shape or V-Shape	

The table below describes the shape of some common molecules using the VSEPR theory.

Shape	Description	Example	Bond Angle	Example	Ball and Stick Model
Linear	▪ Two atoms are bonded together. OR ▪ Two atoms are bonded to a central atom with no lone pair of electrons on the central atom. ▪ There is equal repulsion between the bonded pairs of electrons.	HCl CO_2	180°	 	 
Trigonal Planar	▪ Three atoms are bonded to a central atom and the central atom does not have any lone pair of electrons. ▪ There is equal repulsion between the bonded pairs of electrons.	BF_3	120°		
Tetrahedral	▪ Four atoms are bonded to the central atom and central atom does not have any lone pair of electrons so there is equal repulsion between each bonded pairs.	CH_4	109.5°		
Trigonal Pyramid (Pyramidal)	▪ Three atoms are bonded to the central atom and the central atom has a lone pair of electrons. The lone pair has a greater repulsion compared to the bonded pair -bonded pair repulsion thus compressing the bond angles and giving rise to a trigonal pyramidal shape.	NH_3	107°		
V-shape or Bent shape	▪ Two atoms are bonded to the central atom and the central atom has lone pairs of electrons which has a greater repulsive forces compared to the bonded pair -bonded pair repulsion, thus compressing the bond angles and giving rise to a bent shape.	H_2O	104.5°		

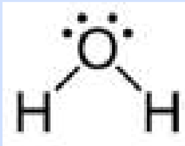
Exercise

1. Explain the basic principles of the VSEPR theory for predicting the shapes of molecules and ions.

2. Predict and justify the electron group geometry and molecular geometry (shape) of the following molecules using the VSEPR theory.

   O_2, BCl_3, CCl_4, NCl_3, H_2S

3. Study the diagram of the water molecule below and answer the questions that follow:



   (i) State the number of electron groups present in the water molecule.
   (ii) Describe the electron group geometry of the water molecule.
   (iii) Describe the molecular geometry (shape) of the water molecule.

5. Using the VSEPR theory, determine the shape of the following molecules:

   i. PCl_3 ii. BeH_2 iii. Cl_2

6. Account for the following:

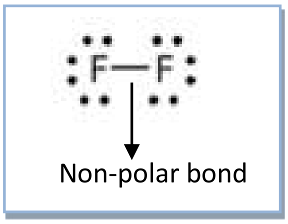
   i. Although both NH_3 and CH_4 have tetrahedral electron group geometry, the shape of NH_3 is trigonal pyramid and that of CH_4 is tetrahedral.
   ii. Although both are triatomic molecules, the H_2S molecule has a bent shape while CO_2 molecule has a linear shape.

7. In VSEPR Theory, under which of the following conditions will the electron group geometry be the same as the molecular geometry?

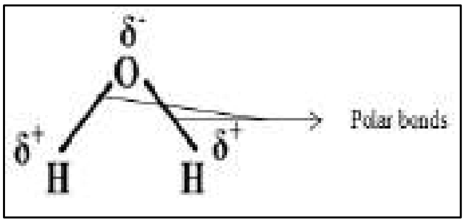
   A. In molecules with at least one lone pair of electrons on the central atom.
   B. In molecules with no lone pairs of electrons on the substituent atoms.
   C. In molecules with more than one octet around the central atom.
   D. In molecules with no lone pairs of electrons on the central atom.

2.1.4 Electronegativity and Bond Polarity

- **Bond polarity** is a measure of how equally the electrons in a bond are shared between the two atoms of the bond.
- Therefore, the concept of electronegativity is used to estimate whether a given bond in a molecule will be **non-polar covalent** or **polar covalent**.
- As the difference in electronegativity between the two atoms of a bond increases, so does the bond polarity.
- When two atoms of equal electronegativity are bonded, both will have the same tendency to attract the shared pair of electrons. **Non-polar covalent bonds** will be formed in this case since there is equal sharing of the shared pair of electrons. To get a bond like this, two similar atoms should be bonded. For instance, this sort of bond can be found in H_2, Cl_2 or F_2 molecules.



- Atoms with high electronegativity when covalently bonded to atoms of lower electronegativity have a greater share of the electrons than the other atom. This is because the more electronegative atom attracts the shared pair of electrons away from the less electronegative atom. Thus, some of the electron density around the less electronegative atom is pulled towards the more electronegative atom, leaving a partial positive charge (δ^+) on the less electronegative atom and a partial negative charge (δ^-) on the more electronegative atom. Such a bond is described as a **polar covalent bond**.
- For example, a water molecule (H_2O) has two hydrogen atoms bonded to one oxygen atom as shown on the right.



- Oxygen being more electronegative pulls the shared pair of electrons closer to itself gaining a δ^- charge and hydrogen atoms gains a δ^+.
- An **ionic bond** results when an atom with far greater electronegativity combines with an atom of lower electronegativity, and the result is the complete transfer of valence electrons to the higher electronegative atom. This transfer results in the formation of positive and negative ions.
- Forces that hold the atoms together in a molecule, such as the covalent and ionic bonds are known as the **intramolecular forces**.

Example:

Compound	F_2	HF	LiF
Electronegativity difference	4.0 - 4.0 = 0	4.0 - 2.0 = 1.9	4.0 - 1.0 = 3.0
Type of bond	Nonpolar covalent	Polar covalent	Ionic

The examples above illustrate that the greater the difference in electronegativity between two atoms, the more polar their bond. Note that C-H bonds are considered non-polar because of their relatively equal electronegativities; C (2.5) and H (2.1).

2.1.5 <u>Molecular polarity</u>

- Molecular polarity or the polarity of a molecule will result from the type of bond(s) it contains (either polar or non-polar), as well as the shape of the molecule.
- For most molecules, there is a direct correlation between molecular polarity and number and types of polar or non-polar covalent bonds it contains.
- For other molecules, they may have polar bonds, but in a symmetrical arrangement which then gives rise to a non-polar molecule. Examples of such molecules include CCl_4, CO_2, and BF_3.
- The two types of non-polar molecules include:

1. Molecules whose atoms have equal electronegativities (non-polar bonds).

   Example: Cl_2

2. Molecules whose atoms have different electronegativities (polar-bonds), but in symmetrical arrangements so the bond polarities cancel out. Example: CO_2

Summary: Determining polarity of molecules

1. Draw the Lewis structure.
2. Determine the shape using the VSEPR Theory:
   (a) Determine the electron group geometry.
   (b) Determine the molecular geometry (shape).
3. Determine the molecular polarity:
   (a) Polar bonds and asymmetrical shape – polar molecule (e.g. NH_3).
   (b) Polar bonds and symmetrical shape – non polar molecule (e.g. CCl_4).
   (c) Non polar bonds and symmetrical shape – non polar molecule (e.g. CH_4).

A table showing the polarity of some molecules:

Molecule	Polarity	Reason
CCl_4	Non-polar	▪ The C-Cl bond is polar. However, the tetrahedral shape of the molecule is symmetrical. ▪ For this reason, the bond polarities cancel out, making the molecule non-polar.
CO_2	Non-polar	▪ The C=O bond is polar. However, the linear shape of the molecule is symmetrical. ▪ For this reason, the bond polarities cancel out, making the molecule non-polar.
H_2O	Polar	▪ The O-H bond is polar. ▪ Since the bent shape is asymmetrical, the bond polarities do not cancel out making the molecule polar.
NH_3	Polar	▪ The N-H bond is polar. ▪ Since the trigonal pyramidal shape is asymmetrical, the bond polarities do not cancel out making the molecule polar.
BF_3	Non-polar	▪ The B-F bond is polar. However, the trigonal pyramid shape of the molecule is symmetrical. ▪ For this reason, the bond polarities cancel out, making the molecule non-polar.
Cl_2	Non-polar	▪ The Cl-Cl bond is non-polar and the shape is symmetrical, thus giving rise to a non-polar molecule.

*For your practical, see the experiment on **Polar Molecules**; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Exercise

1. Out of the following pairs, which bond is more polar? Justify your answers.

   (a) B−Cl or C−Cl
   (b) P−F or P−Cl
   In each case, give reasons in terms of electronegativity difference and indicate which atom has the partial negative charge.

2. Consider a Cl_2 molecule. Explain if the bond between the Cl atoms is polar or non- polar.

3. Explain whether the H_2S molecule is polar or non-polar.

4. Account for the following statements.
   i. The C-Cl bond in a carbon tetrachloride molecule (CCl_4) is polar but the overall molecule is non-polar.
   ii. The BCl_3 molecule is non-polar.
   iii. N_2 is a non-polar molecule.

2.1.6 Intermolecular Forces

- In contrast to **intramolecular forces that hold atoms** together in molecules, **intermolecular forces hold molecules** together in a compound.
- Intermolecular forces are generally much weaker than covalent bonds.
- For instance, a lot of energy is required to overcome the intramolecular forces and break both O–H bonds in a water molecule. However, it takes a little energy to overcome the intermolecular attractions and convert liquid water to water vapor at 100°C.
- Another example of intermolecular forces is the **van der Waals forces** which exist between molecules such as iodine (I_2) and chlorine (Cl_2).

2.1.7 Bonding in Solids

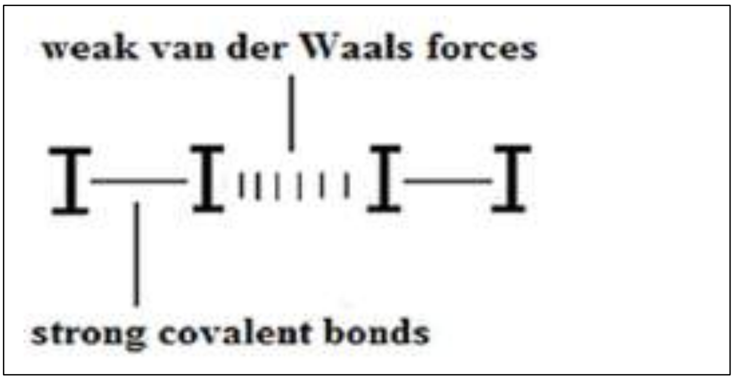
The major types of solids are discrete molecular, ionic, metallic and covalent solids.

Structure and properties of some solids

1. Discrete molecular solids

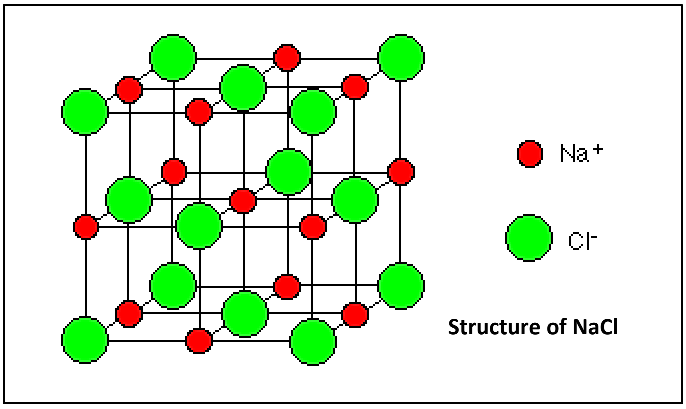
- Molecular solids have simple structures and are generally soft.
- The molecules are held together by weak intermolecular forces known as van der Waals forces or hydrogen bonds.
- Such solids sublime easily or have low melting and boiling point since not much energy is required to break the weak intermolecular forces of attraction.
- They do not conduct electricity because there are no ions or free moving electrons in the structure.
- It should be noted that the intramolecular bonds (bonds within molecules) are strong covalent bonds which cannot be broken easily.
- Examples of molecular solids include: I_2 and solid carbon dioxide (dry ice).

Intermolecular and intramolecular forces in iodine (I_2)



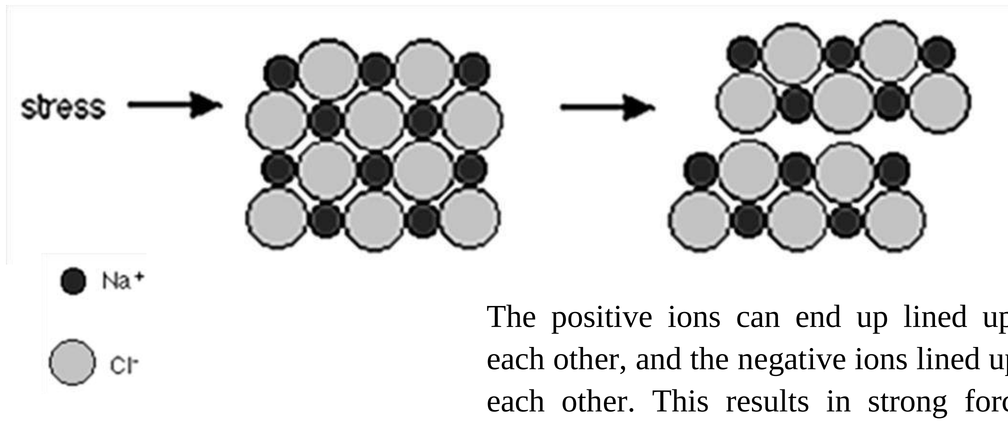
2. Ionic solids

- Ionic compounds are made up of cations and anions as shown for NaCl on the right.



Source: *http://chemwiki.ucdavis.edu*

- In solid state, the ions are in fixed positions arranged in a large 3-dimensional ionic lattice.
- The strong electrostatic forces of attraction between oppositely charged ions hold the lattice together.
- Ionic compounds have high melting and boiling points because a lot of energy is required to break the strong electrostatic attractions between the positive and negative ions.
- Ionic compound do not conduct electricity in solid state because the ions are fixed in the ionic lattice and are not able to move.
- Ionic compounds conduct electricity in molten/solution form since upon dissociation the ions are free to move and hence can conduct electricity. *(See notes on hydration).*
- Ionic compounds are not soluble in organic solvents because the attractions between the solvent molecules and the ions are not big enough to overcome the attractions holding the ionic lattice together.
- Ionic solids are crystalline brittle. If an attempt is made to distort the crystals, it will shatter (break down). This is because an external force (stress) causes the planes of ions to shift. This results in strong repulsive forces which occur when ions of the same charge are brought side-by-side. Thus the crystal repels itself to pieces as shown below.

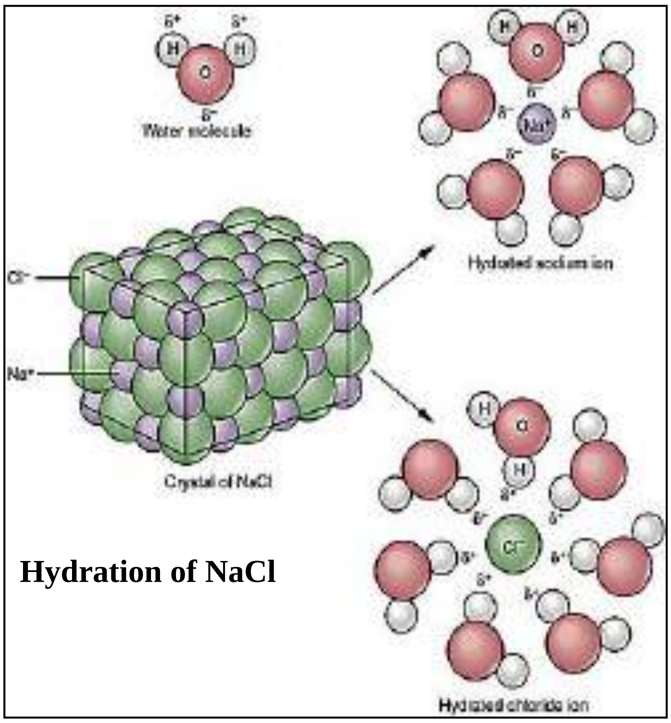


The positive ions can end up lined up with each other, and the negative ions lined up with each other. This results in strong forces of repulsion and breaking of the crystal.

Source: *http://chemwiki.ucdavis.edu*

Hydration of ions

- The ability of ions and other molecules to dissolve in water is due to polarity.
- The breaking down of ionic lattice or crystals is known as the solvation process, whereby the salt dissolves and the ions are free to move around in the water solvent. Since water is the solvent, the solvation process in this case is called **hydration**.
- For example, in the illustration on the right, sodium chloride is shown in its crystalline form, as well as in its dissolved form in water.



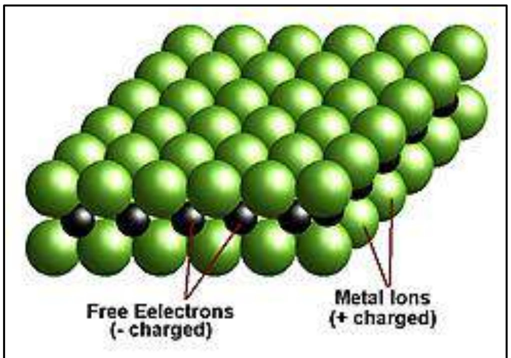
Hydration of NaCl

Source: http://cnx.org/

- When an ionic solid is mixed with water, the partial positive ends of polar water molecule are attracted to the negative ions of the solid, while the partial negative ends are attracted to the positive ions of the solid.
- Since the force of attraction between water and the ion is greater than the attraction between positive and negative ions, the ions separate. Thus, NaCl dissociates into its ions by hydration.
- The attraction of polar water molecules to the ions results in hydrated ions.

3. Metallic solids

- Metals have metallic bond in which the positive metal ions are strongly attracted to the valence electrons of the atoms. This gives rise to high melting and boiling point of the metals.
- These valence electrons form a ‘sea of electrons’ around the positive charges. Since these electrons are free to move, metals can conduct electricity in solid state.
- The electrons are also able to absorb and re-emit light of all wavelengths, making the metals lustrous.



Structure of metals

Source: http://www.green-planet-solar-energy.com

- The free electrons also allow the layers of charges to move without creating any repulsive forces or breaking of the metallic bond. Therefore, when stress is applied, the structure of metals does not change, but the shape of the metal changes. This is why metals are malleable and ductile.

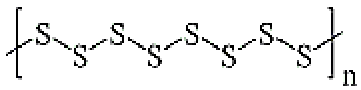
3. Covalent Network Solids

- These solids are made up of atoms held together by a network of strong covalent bonds. They can be classified as linear, 2-dimensional and 3-dimensional solids.

1. Linear solids

(i) Plastic sulphur

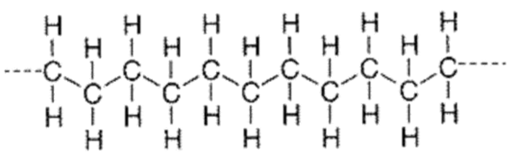
Plastic sulphur, a non-crystalline or amorphous form of sulphur, is formed when molten sulphur is poured into cold water. Dark brown or black ribbons of sulphur are formed, which are very soft and can be stretched easily. Plastic sulfur consists of very long chains of sulfur atoms aligned and held by van der Waals forces. It is unstable and changes into rhombic sulphur on slight heating or even at room temperature.



Structure of linear polymer of sulphur
Source: http://chemwiki.ucdavis.edu

(ii) Polyethene

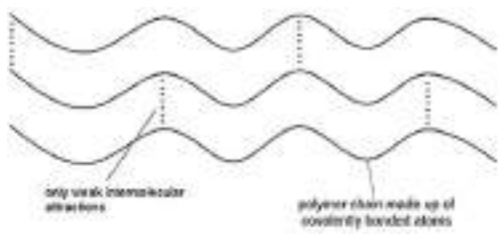
Polyethene (also known as polythene or polyethylene) is an amorphous solid that consists of more or less randomly oriented chains of ($-CH_2-CH_2-$) linkages. One chain is held to its neighbors in the structure by van der Waals forces. Polyethene is the simplest example of a linear polymer.



A polyethene chain

(Linear and covalently bonded)

Source: http://www.chemheritage.org/

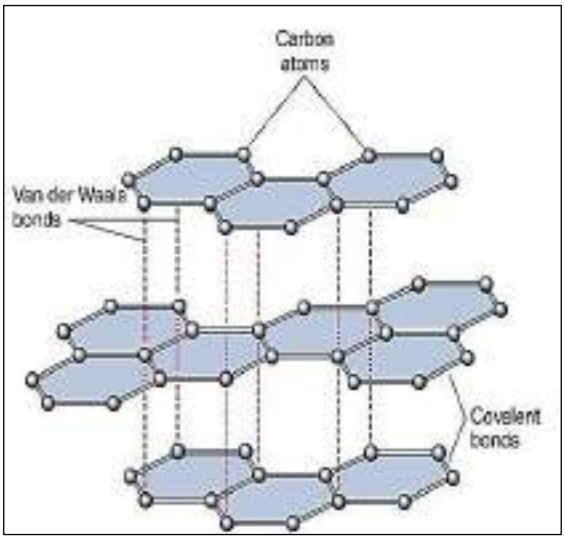


Van der waals forces between polyethene chains

Source: https://www.studyblue.com

2. 2-Dimensional solids

Graphite



Structure of Graphite

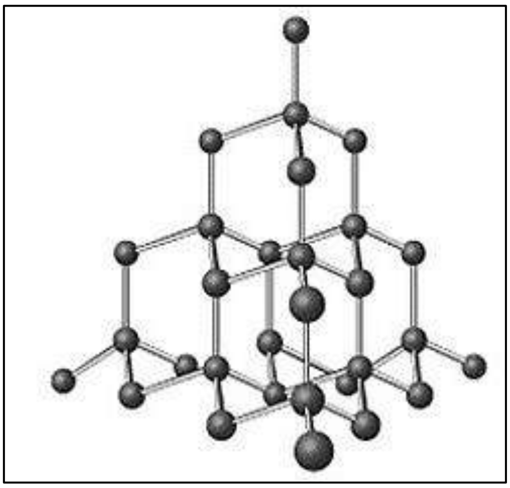
Source: http://chemistrycarbons.blogspot.com

- Each carbon atom is bonded to three adjacent carbon atoms by strong covalent bonds.
- Graphite has high melting point because a lot of energy is required to break the strong covalent bonds.
- Between the layers are delocalized electrons which are able to move freely, allowing graphite to conduct electricity.
- The carbon atoms are arranged in layers of interconnected hexagonal rings.
- The layers of carbon atoms are held together by weak van der Waals forces of attraction. These layers can slide over each other making graphite soft and slippery.

3. 3-Dimensional Solids

i. Diamond

- Each carbon atom is bonded to four other carbon atoms by strong covalent bonds in a three dimensional network. For this reason, diamond is extremely hard and has high melting and boiling point.
- Diamond is a non-conductor of electricity because there are no free moving electrons in the structure as all the valence electrons are involved in bonding.



Structure of Diamond

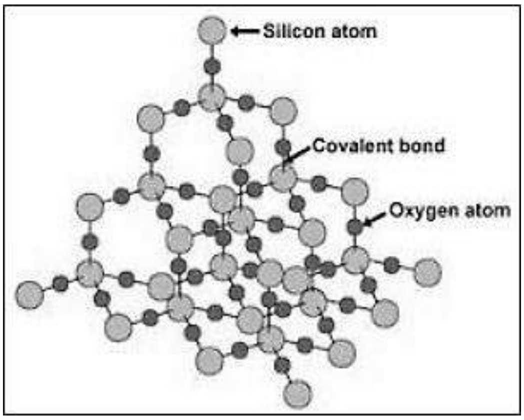
Source: https://www.learner.org

Diamonds are the very hardest natural substance. The only thing that can scratch a diamond is another diamond.

Source:http://www.brilliantearth.com

ii. Silicon dioxide (SiO_2)

- Also known as silica.
- Each silicon atom is covalently bonded to four oxygen atoms and each oxygen atom is covalently bonded to two silicon atoms in a giant covalent structure.
- For this reason, silicon dioxide is very hard and has high melting and boiling point.
- It does not conduct electricity because there are no free moving electrons in the structure.

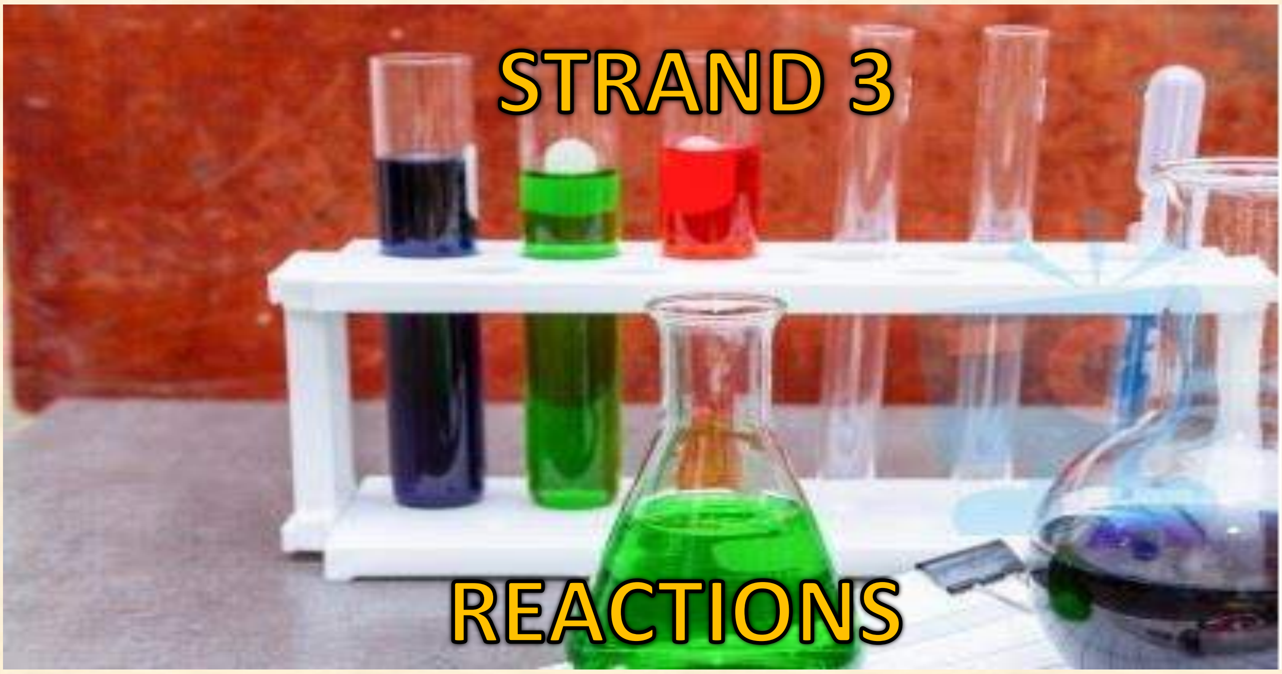


Structure of silicon dioxide

Source: http://www.green-planet-solarnergy.com

Exercise

1. With the help of a diagram, explain the polarity of a water molecule.
2. Give two properties of ionic compounds.
3. Using a suitable diagram, explain why potassium chloride (KCl) is able to dissolve in water. What is this process known as?
4. What is the difference between intermolecular and intramolecular bonding? Which one is stronger and why?
5. Account for the following properties:
   i. Solid iodine can sublime at relatively low temperature.
   ii. Diamond is hard whereas graphite is soft and slippery.
   iii. Graphite can conduct electricity but diamond cannot.
   iv. Metals can conduct electricity in solid state.
   v. Ionic compounds have a high melting and boiling point.
   vi. Metals are lustrous.
   vii. Diamond and graphite have a high melting and boiling point.
   viii. Silicon dioxide is hard and has high melting and boiling point.
6. Briefly explain why ionic compounds cannot conduct electricity in solid state but can conduct electricity in liquid state (in molten and solution).
7. Plastic bottles are usually made up of polyethene. Describe the type(s) of molecular forces present in such a material.
8. In terms of its structure and bonding, explain why plastic sulphur can be stretched easily.



Strand Outcome

Demonstrate an understanding of the chemical principles that involve changes during chemical reactions.

Sub-strands

3.1 Quantitative Chemistry
3.2 Oxidation-Reduction
3.3 Physical Chemistry

3.1 QUANTITATIVE CHEMISTRY

Achievement indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Relate quantities between reactants and products by performing stoichiometric calculations.
- ✓ Determine the empirical and molecular formula of a compound from mass percent composition and molar mass data.
- ✓ Determine the water of crystallisation of a hydrated salt.
- ✓ Prepare a standard solution.
- ✓ Determine the concentration of a solution by volumetric analysis (titration).
- ✓ Describe gravimetric analysis of ions.

3.1.1 Quantitative Chemistry

Quantitative chemistry enables one to calculate known quantities of materials. For instance, calculating how much product can be made from a known starting material or calculating the quantity of a given component present in a sample.

Relationship between reactants and products

To understand quantitative chemistry, it is important to understand the various terms used in the quantitative analysis of a situation and their application.

- Some important terms are as follows:

1. Avogadro's Number (6.02 x 10^{23})

- The mass of a single atom or molecule is so small that it is impossible to measure it in the laboratory.
- Scientists have chosen a number of atoms which have a mass in grams equivalent to the mass of one atom.
- The Avogadro's number is chosen as the number of atoms in twelve gram of $^{12}_{6}C$. This number has been found to be 6.02×10^{23}.
- Therefore, Avogadro's number is defined as the number of elementary particles (which can be atoms, ions or molecules) in one mole of a substance.
- The symbol for Avogadro's number is N_A and its unit is (per mole) mol^{-1}.

Amadeo Avogadro
Source:http://www.glogster.com

2. The Mole

- The mole measures the **amount** of a substance containing the same number of particles as there are atoms in exactly 12 grams of Carbon ($^{12}_{6}C$), which is 6.02×10^{23} atoms.
- The symbol for moles is *n* and its unit is mol.

1 mole = 6.02×10^{23} particles (atoms, ions or molecules)

Generally: **Number of particles = moles × Avogadro's Number**

Number of molecules = moles of molecules × Avogadro's number

Total number of ions = moles of compound × number of ions in the compound × Avogadro's number

Number of atoms in a compound = moles × number of atoms in the compound × Avogadro's number

Example 1

- 1 mole of carbon (C) contains 6.02×10^{23} carbon atoms.
- 1 mole of ammonia (NH_3) contains 6.02×10^{23} ammonia molecules.
- 1 mole of sodium chloride (NaCl) contains 6.02×10^{23} sodium ions (Na^+) and 6.02×10^{23} chloride ions (Cl^-).

Example 2

Find the number of hydrogen atoms in 3 moles of water (H_2O).

Solution

In H_2O, there are two hydrogen atoms. Thus;

$$\text{No. of hydrogen atoms} = \text{moles} \times \text{number of hydrogen atoms} \times \text{Avogadro's Number}$$
$$= 3 \times 2 \times 6.02 \times 10^{23}$$
$$= \underline{3.61 \times 10^{24}}$$

Exercise

1. Use Avogadro's number to find the number of:

   (a) Helium atoms in 0.5 moles of the gas.
   (b) H atoms in 0.65 moles of hydrogen gas.
   (c) Sodium atoms in 6 moles of sodium (Na).
   (d) Oxygen atoms in 2 moles of water.
   (e) Water molecules in 2.5 moles of water (H_2O).
   (f) H atoms in 4 moles of ammonia (NH_3) molecule.
   (g) Sodium ions in 6 moles of sodium chloride (NaCl).
   (h) Chloride ions in 3 moles of sodium chloride (NaCl).
   (i) Carbon atoms in 4.5 moles of carbon dioxide (CO_2) molecules.

2. Change the following quantities to the amount present in moles of the particles.

   (a) 4.0×10^{23} molecules of H_2O.
   (b) 0.5×10^{23} molecules of CO_2.
   (c) 3.6×10^{25} atoms of hydrogen.
   (d) 1.5×10^{23} ions of K^+.

3. Molar mass

- Molar mass is the mass of one mole of any chemical compound.
- The symbol for molar mass is *M*.
- The unit for molar mass is grams per mole (g mol^{-1}).
- Molar mass gives the amount of atoms or molecules or ions present in one mole of the substance.
- For an element, the molar mass is the mass of 1 mole of atoms of that element.
- For a covalent molecular compound, the molar mass is the mass of 1 mole of molecules of that compound.
- For an ionic compound, the molar mass is the mass of 1 mole of the formula units.

Note: The molar mass of any substance is its atomic mass, molecular mass, or formula mass in grams per mole (g mol^{-1}).

Example

Substance	Atomic mass (amu)	Molecular mass (amu)	Formula mass (amu)	Molar mass (g mol^{-1})
Carbon (C)	12			12
CO_2		44		44
NaCl			58.5	58.5

How to find the molar mass of an element

- The molar mass of an element is found by looking at the atomic mass of that element on the periodic table.
- For example, the mass of one atom of Carbon is 12 amu. Therefore, the molar mass of carbon is equal to 12 g mol^{-1}.
- In the case of hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine, the elements are diatomic, which means that each molecule of the element has two atoms of that element joined together.
- For example, the formula of hydrogen is H_2 and nitrogen is N_2.
- The molar mass of diatomic molecules is calculated by multiplying the molar mass of the atoms by two. For instance, the molar mass of nitrogen (N_2) would be 28 g mol^{-1} (2×14 g mol^{-1}).
- For any chemical compound that is not an element, the molar mass can be found from the chemical formula. The molar mass will be equal to the sum of the molar masses of all the atoms in one molecule of that compound.

Example 1

In carbon dioxide (CO_2), there is one atom of carbon and two atoms of oxygen. Therefore, the molar mass will be equal to the molar mass of carbon plus the molar mass of oxygen.

$$1 \times M\,(\mathrm{C}) + 2 \times M\,(\mathrm{O}) = (1 \times 12\ \mathrm{g\ mol^{-1}}) + (2 \times 16\ \mathrm{g\ mol^{-1}}) = \underline{\mathbf{44\ g\ mol^{-1}}}$$

Example 2

For table salt (NaCl), the molar mass will be equal to the molar mass of sodium plus the molar mass of chlorine.

$$1 \times M\,(\mathrm{Na}) + 1 \times M\,(\mathrm{Cl}) = (1 \times 23\ \mathrm{g\ mol^{-1}}) + (1 \times 35.5\ \mathrm{g\ mol^{-1}}) = \underline{\mathbf{58.5\ g\ mol^{-1}}}$$

Exercise

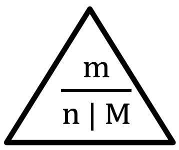
Find the molar masses of the following compounds:

1. Sodium fluoride (NaF)
2. Potassium hydroxide (KOH)
3. Ammonia (NH_3)
4. Water (H_2O)

Mole Calculation

- In most calculations, the mass of a substance needs to be converted to an amount (moles).

  Conversions are done using the following relationship:



Where:

$\text{moles } (n) = \dfrac{\text{mass (m)}}{\text{molar mass } (M)}$	$\text{mass (m)} = \text{moles } (n) \times \text{Molar mass } (M)$

Example

1. What is the amount of copper atoms in 5.00 g of copper?

$$n = \frac{\mathrm{m}}{M}$$
$$= \frac{5.0\ \mathrm{g}}{64\ \mathrm{g\ mol^{-1}}}$$
$$= \underline{\mathbf{7.8 \times 10^{-2}\ mol}}$$

2. What is the amount of magnesium ions in 109 g of magnesium chloride ($MgCl_2$)?

$$n = \frac{\mathrm{m}}{M}$$
$$= \frac{109\ \mathrm{g}}{95\ \mathrm{g\ mol^{-1}}}$$
$$= \underline{\mathbf{1.14\ mol}}$$

3. What is the mass of 2.35 moles of argon atoms?

$$\mathrm{m} = n \times M$$
$$= 2.13\ \mathrm{mol} \times 40\ \mathrm{g\ mol^{-1}}$$
$$= \underline{\mathbf{85.2\ g}}$$

4. Find the mass of 4.8 mol of ammonia gas (NH_3)?

$$\mathrm{m} = \mathrm{n} \times M$$
$$= 4.8\ \mathrm{mol} \times 17\ \mathrm{g\ mol^{-1}}$$
$$= \underline{\mathbf{81.6\ g}}$$

4. Stoichiometry in Chemical Reactions

- The quantitative relationship among reactants and products is called **stoichiometry.**
- In order to use stoichiometry to do calculations based on chemical reactions, it is important to first understand the relationship that exist between products and reactants and why they exist. This requires understanding of how to balance chemical equations.
- In stoichiometry, balanced equations make it possible to compare different elements in a chemical reaction.
- From the balanced equation, mole ratios are determined which actually helps in determining the required quantities of the desired products or reactants.

Note:

The coefficients from the balanced chemical equations specify the mole of the reactant or product.

The following steps can be followed when doing calculations for chemical reactions

- Write the balanced chemical equation.
- Calculate the amount (moles) of known substance.
- Find the relationship thorough mole ratios between what you are given and what you are trying to find out.
- Calculate the amount of unknown substances either in moles or grams as required.

Example 1

Determine the amount (in mol) of magnesium chloride formed when 2.55 grams of magnesium (Mg) combines with sufficient chlorine (Cl_2) to form magnesium chloride ($MgCl_2$).

Solution

i. Write balanced equation for the reaction:

$$Mg + Cl_2 \rightarrow MgCl_2$$

ii. Find the amount in moles of Mg:

$$n = \frac{m}{M}$$

$$= \frac{2.55\ \text{g}}{24\ \text{g mol}^{-1}}$$

$$= \underline{0.11\ \text{mol}}$$

iii. Compare mole ratio:

$$Mg : MgCl_2$$
$$1 : 1$$
$$0.11 : 0.11$$

Thus 0.11 mol of $MgCl_2$ will be formed.

Example 2

Propane (C_3H_8) undergoes combustion reaction. The equation for the combustion of propane is shown below.

$$C_3H_8 + 5O_2 \rightarrow 4H_2O + 3CO_2$$

If 21 g of propane is burnt, how many grams of H_2O is produced?

Yes! It is balanced

Solution

- ✓ Ensure that the equation given is balanced. **Check**.
- ✓ Since you cannot calculate from mass of reactant to mass of products directly, you must convert mass of C_3H_8 to moles of C_3H_8 then from moles of C_3H_8 find moles of H_2O. Then find mass of H_2O from moles of H_2O.

Does it look confusing? Read again. It is very simple and easy to follow. Look at the calculation!!!

Calculation

1. Find moles of C_3H_8:

$$n = \frac{m}{M}$$

$$= \frac{21\text{ g}}{44\text{ g mol}^{-1}}$$

$= \underline{0.48\text{ mol}}$

2. Compare mole ratio:

$$C_3H_8 + 5O_2 \rightarrow 4H_2O + 3CO_2$$

$C_3H_8 : H_2O$

1 : 4

0.48 : X

$\underline{1.92\text{ mol} = X}$

Cross-multiple method is used to find X, which is the amount (in mol) of water produced in the reaction.

3. Mass of water:

$$n = \frac{m}{M}$$

$m = n \times M$

$= 1.92\text{ mol} \times 18\text{ g mol}^{-1}$

$= 34.56\text{ g}$

❖ **Therefore, the mass of water produced is 34.56 g**

For your practical, see the experiment on ***Stoichiometry of the lead-sulphur reaction****, Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Exercise

1. According to the following equation, what mass of carbon dioxide is produced in the complete combustion of 36.5 g of ethanol?

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

2. What mass of iron is produced when 50.0 g of carbon monoxide (CO) reacts with iron(III) oxide? The reaction equation is as follows:

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

3. When copper carbonate is heated, it decomposes. The reaction is:

$$CuCO_3 \rightarrow CuO + CO_2$$

   i. How many moles of copper (II) oxide forms when 247 g of copper carbonate is completely decomposed?
   ii. How many moles of copper carbonate must be decomposed to produce 11 g of carbon dioxide?
   iii. If 318 g of copper (II) oxide is produced, what mass of carbon dioxide is liberated?

4. What mass of iron would be needed to produce 34.5 g of iron(III) oxide?

$$4Fe + 3O_2 \rightarrow 2Fe2O3$$

5. Magnesium burns in air to form MgO as follows:

$$2Mg + O_2 \rightarrow 2MgO$$

   i. Calculate the amount (in moles) of Mg used in the reaction if 2.4 g of Mg was burnt in air.
   ii. How many moles of MgO were produced?
   iii. Calculate the mass of MgO produced in the reaction.

6. What mass of carbon dioxide would be formed if 20.0 g of methane gas is burnt?

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

7. Sodium reacts with water to form sodium hydroxide and hydrogen gas. What mass of sodium would be needed to make 100 g of sodium hydroxide?

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

8. Consider this reaction:

$$4NH_3 + 6NO \rightarrow 5N_2 + 6H_2O$$

   How many grams of each reactant were there if 13.7 moles of N_2 is produced?

5. Empirical Formula (EF)

- The empirical formula expresses the **simplest** whole number ratio of atoms or ions in a compound.
- For instance, empirical formula of ethane (C_2H_6) is CH_3.
- When analysis is carried out to find out what or how much of a particular element is present in compound, the result is often expressed as a percentage composition. From the percentage composition the empirical formula can be calculated.

Steps for calculating the empirical formula

1. Find the masses of each element in hundred grams of the compound *(if percentage composition of the elements is given).*

2. Convert the masses to amount (in moles).

3. Divide each answer in Step 2 by the smallest answer in Step 2 to get the lowest whole number ratio.

4. Where required, make small approximations if needed to get a simple whole number ratio. This can be done by multiplying the empirical formula obtained by a suitable whole number as such:

   - If the final ratio has 0.5 in it, then multiply by 2.
   - If the final ratio has 0.33 in it, then multiply by 3.
   - If the final ratio has 0.25 or 0.75 in it then multiply by 4.
     (See Example 2 below)

Example 1

Find the empirical formula of a compound which contains 85.7 % carbon and 14.3% hydrogen.

1. *Consider 100 g of the organic compound:*

   The mass of carbon in the organic compound is 85.7 g and the mass of hydrogen is 14.3 g.

2. *Amount(in mol)*

   C: $n = \frac{m}{M} = \frac{85.7\ \text{g}}{12\ \text{g mol}^{-1}} = 7.14\ \text{mol}$

   H: $n = \frac{m}{M} = \frac{14.3\ \text{g}}{1\ \text{g mol}^{-1}} = 14.3\ \text{mol}$

3. *Ratio of atoms*

   C: H

   $\frac{7.14}{7.14} : \frac{14.3}{7.14}$

   1 : 2 **Therefore the empirical formula of the compound is CH_2.**

Example 2

1.00 g sample of phosphorus powder was burned in air and reacted with oxygen gas to give 2.29 g of a phosphorus oxide. Determine the empirical formula of the phosphorus oxide.

Solution:

1) Calculate amount (in mol) of phosphorous and oxygen:

P $n = \frac{\text{m}}{M} = \frac{1.00\text{ g}}{31\text{ g mol}^{-1}} = 3.2 \times 10^{-2}$ mol

O $n = \frac{\text{m}}{M} = \frac{1.29\text{ g}}{16\text{ g mol}^{-1}} = 8.1 \times 10^{-2}$ mol

Mass of oxygen = mass of phosphorous oxide - mass of phosphorous

= 2.29 g - 1.00 g
= 1.29 g

2) Ratio of atoms

P : O

$\frac{0.032}{0.032} : \frac{0.081}{0.032}$

1 : 2.5

Ratio is $P_1O_{2.5}$

In the final ratio, O is 2.5 so multiply by 2 to both the atoms to get a whole number ratio: $2(P_1O_{2.5}) = P_2O_5$.

Thus the empirical formula is **P_2O_5.**

Exercise

1. The molecular formula for glucose is $C_6H_{12}O_6$. What is its empirical formula?
2. Determine the empirical formula of a compound composed of:
   i. 80% copper and 20% oxygen
   ii. 53% aluminium and 47% oxygen
   iii. 1.6% hydrogen, 22.2% nitrogen and 76.2% oxygen
   iv. 6.7% hydrogen, 40% carbon and 53.3% oxygen
3. 15 g of a substance yields 4.1 g of carbon and 10.9 g of oxygen. Determine the empirical formula of this substance?
4. Calculate the empirical formula of the oxide of sulphur which is 40 % sulphur by weight.
5. Calculate the empirical formula of the following hydrocarbons.
   i. A hydrocarbon that contains 90% carbon
   ii. A hydrocarbon which contains 82.7% carbon and 17.3% hydrogen
   iii. A hydrocarbon which contains 80% carbon and 20% hydrogen

6. Molecular Formula (MF)

- The molecular formula expresses **the actual number of atoms** of each type in the compound.
- For some substances, the molecular formula of a compound is the same as the empirical formula and for some it can be a multiple of the empirical formula.

Example

Molecular Formula	Empirical formula
C_2H_4	CH_2
$C_6H_{12}O_6$	CH_2O
H_2O	H_2O

Molecular formula = X (Empirical formula) or MF = X (EF)

Where:

$$X = \frac{\text{Molar mass of molecular formula}}{\text{Molar mass of empirical formula}}$$

Note: X determines how many times greater the molecular mass is compared to the mass of the empirical formula.

Example

What is the molecular formula of an organic compound which has a molecular mass of 180 g mol^{-1} and its empirical formula is CH_2O.

Solution

i.) Calculate molar mass of empirical formula (CH_2O).

EF molar mass = C (1) + H (2) + O (1)
= 12 g mol^{-1} (1) + 1 g mol^{-1} (2) + 16 g mol^{-1} (1)
= **30 g mol^{-1}**

ii.) Find n.

$$X = \frac{\text{Molar mass of molecular formula}}{\text{Molar mass of empirical formula}}$$

$$= \frac{180\ \text{g mol}^{-1}}{30\ \text{g mol}^{-1}}$$

= 6

iii.) Molecular formula = X (Empirical formula)
= 6 (CH_2O)
= **$C_6H_{12}O_6$**

Exercise

1. The empirical formula of a hydrocarbon is C_3H_7. Its molar mass is 86 g mol^{-1}. What is the molecular formula of this hydrocarbon?

2. Find the molecular formula of an acid with a molar mass of 98 g mol^{-1} and the percentage composition of the elements in the acid is as such:
   Hydrogen = 2%
   Sulphur = 33%
   Oxygen = 65%

3. The compound ethylene glycol is often used as antifreeze. It contains 38.70% carbon, 9.75% hydrogen and the rest is oxygen. The molecular mass of ethylene glycol is 62 g mol^{-1}. What is the molecular formula of ethylene glycol?

4. A compound having a molecular mass of 194 g mol^{-1} was found to have an elemental analysis of 49.48% carbon, 5.20% hydrogen, 16.47% oxygen, and 28.85% nitrogen. Determine the molecular formula of this compound?

5. Determine the empirical formula and the molecular formula of a drug which has the percentage composition of: 60.60% C, 7.07% H, and 32.30% O. The molecular mass of the drug is 198 g mol^{-1}.

6. Caffeine has percentage composition of 49.50% carbon, 5.20% hydrogen, 28.90% nitrogen and 16.50% oxygen. Calculate the empirical formula of caffeine. If the molar mass of caffeine is 194 g mol^{-1}, determine its molecular formula.

7. Water of Crystallisation

- When crystals of hydrate salts are formed, they do so with a definite number of molecules of water, chemically combined in a definite proportion. This water is called water of crystallisation.
- Water of crystallisation is the number of water molecules, chemically combined in a definite molecular proportion, with the salt in its crystalline state.
- This water is responsible for the geometric shape and colour of the crystals.
- A substance containing water of crystallisation is called a hydrous substance or a hydrate.
- This water can be removed by heating. The salt obtained is said to have become anhydrous.
- The percentage water of crystallisation can be found using the formula:

$$\%\text{ of water of crystallisation} = \frac{\text{mass of water}}{\text{mass of hydrated salt}} \times \frac{100}{1}$$

- Examples of some common hydrated salt include:
  - Hydrated copper sulphate: $CuSO_4.5H_2O$
  - Hydrated sodium carbonate: $Na_2CO_3.10H_2O$
  - Hydrated iron sulphate: $FeSO_4.7H_2O$

Example 1: Calculating theoretical value

Calculate the percentage of water in hydrated magnesium sulphate salt crystals ($MgSO_4.7H_2O$).

Solution

Molar mass of hydrated magnesium sulphate

= 1(Mg) + 1(S) +11(O) + 14(H)

=1(24 g mol^{-1}) + 1 (32 g mol^{-1}) + 11(16 g mol^{-1}) + 14 (g mol^{-1})

= 24 g mol^{-1} + 32 g mol^{-1}+ 176 g mol^{-1} + 14 g mol^{-1}

= **246 g mol^{-1}**

$MgSO_4.7H_2O$ has 7 moles of water

Molar mass of water

= 14(H) + 7(O)

= 14(1 g mol^{-1}) + 7(16 g mol^{-1})

= 14 g mol^{-1} + 112 g mol^{-1}

= **126 g mol^{-1}**

Therefore:

$$\% \text{ of water of crystallization} = \frac{\text{mass of water}}{\text{mass of hydrated salt}} \times \frac{100}{1}$$

$$= \frac{126 \text{ g mol}^{-1}}{246 \text{ g mol}^{-1}} \times \frac{100}{1}$$

= **51.22%**

Example 2: Calculating experimental value

2.46 g of hydrated magnesium sulphate ($MgSO_4.XH_2O$) was heated to a constant mass of 1.20 g. Using these information, determine:

i. The percentage water of crystallisation in this salt.

ii. The value of X.

iii. The formula of the hydrated salt.

Solution

Mass of hydrated magnesium sulphate = 2.46 g

Mass of anhydrous magnesium sulphate = 1.20 g

Mass of water of crystallisation = 1.26 g

> **Note:**
>
> Mass of water = Mass of hydrated salt - Mass of anhydrous salt
>
> = 2.46 g - 1.26 g
>
> = 1.26 g

i. $\% \text{ of water of crystallisation} = \frac{\text{mass of water}}{\text{mass of hydrated salt}} \times \frac{100}{1}$

$= \frac{1.26 \text{ g}}{2.46 \text{ g}} \times \frac{100}{1}$

= **51.2%**

ii. Compare mole ratio:

$MgSO_4 : H_2O$

$\frac{1.20 \text{ g}}{120 \text{ g mol}^{-1}} : \frac{1.26 \text{ g}}{18 \text{ g mol}^{-1}}$

$\frac{0.01}{0.01} : \frac{0.07}{0.01}$

1 : 7

Therefore, the value of X is **7**.

iii. The formula of the hydrated salt is **$MgSO_4.7H_2O$**.

*For your practical, see the experiment on **Water of Crystallisation**; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Exercise

1. Calculate the percentage water of crystallization in the following:
   i. $CuSO_4.5H_2O$
   ii. $Na_2CO_3.10H_2O$
2. 10.222 g sample of hydrated barium iodide (BaI_2) is heated to dry off the water of crystallization. The dry sample has a mass of 9.520 g. What is the formula of the hydrate?
3. 2.07 g sample of washing soda ($Na_2CO_3.XH_2O$) was strongly heated to drive off the water of crystallization. 0.77 g of the anhydrous salt was obtained. Find the value of X and write down the formula of washing soda.
4. An experiment was performed to find the percentage water of crystallisation in a sample of Magnesium sulphate ($MgSO_4.XH_2O$). The following data was obtained:
   Mass of crucible + lid = 20.00 g
   Mass of crucible + lid + magnesium sulphate (before heating) = 22.50 g.
   Mass of crucible and lid + magnesium sulphate (after heating) = 21.22 g.
   Calculate:
   i. Mass of hydrated magnesium sulphate.
   ii. Mass of anhydrous magnesium sulphate
   iii. Mass of water driven off.
   iv. Percentage water of crystallisation in the compound.
   v. The value of X.
5. Strong heating of 6.72 g of magnesium sulphate ($MgSO_4$. XH_2O) produced 4.12 g of anhydrous magnesium sulphate. Calculate the value of X and hence write the formula of the hydrated salt.
6. 6.25 g of hydrated copper (II) sulphate, $CuSO_4.XH_2O$, was gently heated in a crucible until the mass remaining was a constant 4.00 g. Determine the value of X.
7. Calcium chloride ($CaCl_2$) is used as a drying agent since it can absorb water from the atmosphere. In an experiment, 5.00 g anhydrous calcium chloride was used as a drying agent until it could no longer absorb any water. The hydrated crystals had a mass of 9.86 g. Calculate the formula of the hydrated salt.

8. Concentration

Concentration is a measure of the amount of solute in a fixed quantity of solvent. In other words it is a measure of the strength of a solution. Concentration can be calculated in two ways:

1. $$\text{Concentration (g L}^{-1}) = \text{ Mass of solute(g)/Volume of solution (L)}$$

2. $$\text{Concentration (mol L}^{-1}) = \text{Amount of solute (mol) / Volume of solution (L)}$$

Example

4.0 g of sodium hydroxide (NaOH) is dissolved in 2.0 L of solution. Calculate its concentration in:

a) g L^{-1}
b) mol L^{-1}

Solution

i. Molar mass (*M*) of NaOH

= 1 × 23 g mol^{-1} + 1 × 16 g mol^{-1} + 1 × 1 g mol^{-1}

= 40 g mol^{-1}

ii. Moles (n)

$$n = \frac{\text{m}}{M}$$

$$= \frac{4.0\text{ g}}{40\text{ g mol}^{-1}}$$

= 0.10 mol

a. **c** in **g L^{-1}**

$$= \frac{4.0\text{ g}}{2.0\text{ L}}$$

= 2.0 g L^{-1}

b. **c** in **mol L^{-1}**

$$= \frac{0.10\text{ mol}}{2.0\text{ L}}$$

= 5.0×10^{-2} mol L^{-1}

Exercise

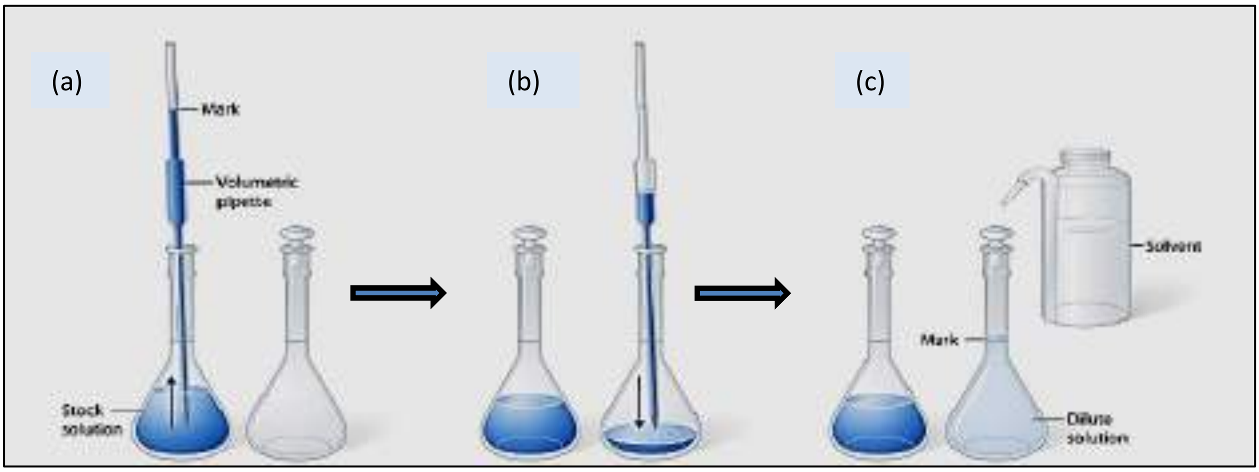
1. What is the concentration in mol L^{-1} of 250 mL of magnesium chloride solution containing 8.4 g of the salt?
2. Calculate the mass of silver nitrate ($AgNO_3$) needed to prepare a solution of 0.02 mol L^{-1}.
3. A student dissolved 2.0 g of sodium carbonate in 50 mL of water. Find the concentration of the solution in:
   i. mol L^{-1}
   ii. g L^{-1}

9. Diluting Solutions

- Dilution is the process of decreasing the concentration of a solution or making a concentrated solution less concentrated.

The formula for solving a dilution problem is: $C_1V_1 = C_2V_2$
where:
C_1 - represents the concentration of initial solution (stock solution).
V_1- represents the volume of initial solution (stock solution).
C_2 - represents the concentration of final solution (diluted solution).
V_2 - represent the volume of final solution (diluted solution).

- The diagram below shows the basic steps which can be taken to perform a dilution.



Source: http://2012books.lardbucket.org/

(a) A volume containing the desired amount of solute is measured from a stock solution of known concentration. (b) The measured volume of stock solution is transferred to a second volumetric flask. (c) The measured volume in the second flask is then diluted with solvent up to the graduation mark.

Example

Find the concentration of a solution made by diluting 45.0 mL of 4.2 mol L^{-1} potassium hydroxide solution to 150 mL?

Solution

✓ First identify the following:

C_1 = 4.2 mol L^{-1}

V_1 = 45.0 mL

C_2 = ?

V_2 = 150 mL

✓ Use the formula:

$$C_1V_1 = C_2V_2$$

$$(4.2 \text{ mol L}^{-1})\ (0.045 \text{ L}) = C_2\ (0.15 \text{ L})$$

$$0.189 \text{ mol} = C_2\ (0.15 \text{ L})$$

$$C_2 = \frac{0.189 \text{ mol}}{0.15 \text{ L}}$$

$$= \underline{\mathbf{1.26\ mol\ L^{-1}}}$$

Alternative method

1. Find amount of KOH (in mol) in the original solution:

   $n = c \times V$

   $= 4.2 \text{ mol L}^{-1} \times 0.045 \text{ L}$

   $= \underline{0.189 \text{ mol}}$

2. Find the new concentration of KOH:
   $c = n/V$
   $= 0.189 \text{ mol} / 0.15 \text{ L}$
   $= \underline{\mathbf{1.26\ mol\ L^{-1}}}$

Exercise

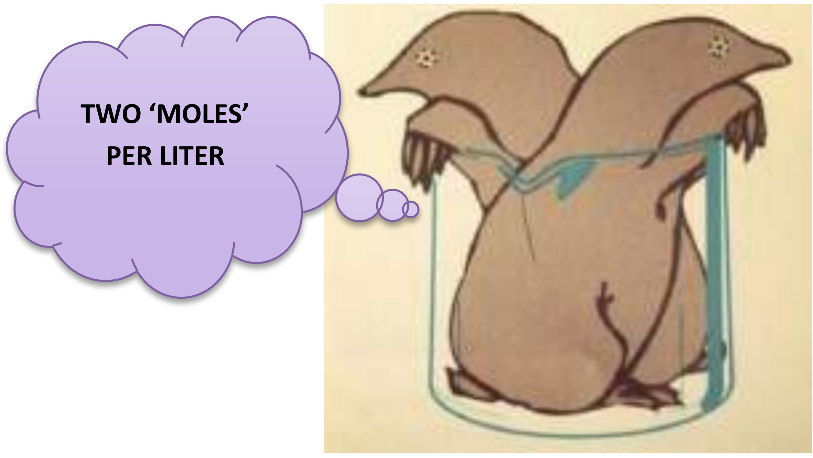
1. What will be the new concentration of a sodium chloride (NaCl) solution, if you dilute 45 mL of 4.2 mol L^{-1} sodium chloride solution to 250 mL?

2. A student took 45 mL of a potassium hydroxide solution and diluted it to 250 mL. The concentration of the new solution was found to be 4.2 mol L^{-1}. What was the initial concentration of the potassium hydroxide solution?

3. A chemist was required to prepare 0.20 mol L^{-1} of glucose solution. What volume of this solution can be made if he has 50.0 mL of 0.50 mol L^{-1} glucose solution in his pharmacy?

4. During an experiment, a student required 100 mL of 0.10 mol L^{-1} hydrochloric acid (HCl) solution. However, the school laboratory only had 0.20 mol L^{-1} of HCl. Calculate the volume of the 0.20 mol L^{-1} HCl solution that must be diluted to give 100 mL of the 0.10 mol L^{-1} HCl solution.

5. Most laboratories keep frequently used stock solutions which are often of high concentration. When required, these stock solutions are diluted and used. Calculate the volume (in mL) of 1 mol L^{-1} sodium hydroxide (NaOH) aqueous solution needed to make 100 mL of 0.5 mol L^{-1} NaOH aqueous solution.



Source: http://www.biogag.com

3.1.2 <u>Quantitative Analysis</u>

- Quantitative analysis is the determination of the amount of a substance present in a given sample.
- A variety of ways can be used for quantitative analysis.
- **Gravimetric** and **volumetric** analysis are two chemical quantitative analysis methods that can be used to determine the amount of a substance in a sample.

1. <u>Gravimetric analysis</u>

- In this method, the ion needed to be analysed is converted into an insoluble salt of known composition that can be separated from the sample and reweighed.
- Some conditions for accurate gravimetric analysis are:

    1. The ion to be analysed must be completely precipitated.
    2. The precipitate must be a pure compound.
    3. The precipitate must be easily filtered out.

> ***<u>Note</u>**: Precipitation reactions occur when cations and anions in aqueous solution combine to form an insoluble ionic solid called a precipitate.*

<u>Steps of gravimetric analysis</u>

1. Prepare a solution containing known weight of the sample.
2. Separate the desired ion by precipitation.
3. Weigh the isolated precipitate.
4. Find the amount of the particular constituent in the sample from the weight and composition of the precipitate obtained.

Example 1: Determination of sulphate in a mineral ore

- To analyse the content of sulphur in a chunk of mineral ore, the ore is treated with concentrated nitric acid and potassium chlorate.
- The purpose of the concentrated nitric acid and potassium chlorate is to convert all the sulfur to sulfate (SO_4^{2-}).
- The nitrate and chlorate are then removed by treating the solution with concentrated HCl.
- Finally, the sulfate is reacted with barium (Ba^{2+}) to form a precipitate ($BaSO_4$).
- Therefore the sulphate is weighed as $BaSO_4$.

Example 2: Determination of chloride content in a compound

1. Place some sample of the compound into a small vial with lid on and dry in the oven. Afterwards, cool it in a desiccator. This is to ensure that correct mass of the dried sample is obtained.
2. Weigh the dried sample into beaker and dissolve it.
3. Add a precipitating agent (Ag^+ and Pb^{2+}) to the solution.

    - To analyse the content of chloride in a compound, a cation must be found that forms an insoluble compound with chloride and must be pure and easily filtered.

> According to the solubility rule, Ag^+ and Pb^{2+} form insoluble chlorides. Therefore, silver chloride (AgCl) could be used to determine the % Cl^-, because it is insoluble, can be formed pure and is easily filtered.

4. Ensure that complete precipitation has occurred.
5. Filter the solution, ensuring that all the precipitate has been transferred from the beaker to the filter.

> Note that it is important that the precipitate is quantitatively transferred to the filter. This is because if any remains in the beaker, the mass obtained will be inaccurate.

6. Dry and weigh the precipitate.
7. Using stoichiometry determine the mass of chloride.
8. Find percentage by mass of chloride by dividing the mass of chloride by the mass of the compound.

9. Given below is a sample calculation.

Sample Calculation

Mass of sample of unknown chloride after drying: 0.09 g

Mass of AgCl precipitate: 0.23 g

Moles of AgCl

$$\frac{0.23\ \text{g}}{143.5\ \text{g mol}^{-1}} = 1.60 \times 10^{-3}\ \text{mol}$$

Moles of Cl^-

$$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(l)}$$

According to the equation above, one mole of AgCl contains one mole of Cl^-.

Therefore: n (Cl^-) = 1.60×10^{-3} mol

Mass of Cl^-

m = n x *M*

= 1.60×10^{-3} mol $\times$ 35.5 g mol^{-1}

= 5.68×10^{-2} g

Percentage of Cl^-

$$= \frac{\textbf{mass of chlorine ions}}{\textbf{mass of sample}} \times 100\%$$

$$= \frac{5.68 \times 10^{-2}\ \text{g}}{0.09\ \text{g}} \times 100$$

= **63.11% Cl^- in unknown chloride sample**

Exercise

1. The concentration of ammonium sulphate in soil can be determined by dissolving the salt and after filtering, adding acidified barium chloride to the solution. The insoluble barium sulphate ($BaSO_4$) that forms is separated, dried and then weighed. In an experiment to determine the concentration of ammonium sulphate in soil, it was found that 1 kg of soil yielded 0.016 g of $BaSO_4$.
   i. What type of analysis is carried out above?
   ii. Calculate the amount (in mol) of $BaSO_4$ in 0.116 g.
   iii. What mass of ammonium sulphate is preset in 1 kg of soil?

2. Volumetric Analysis

- Titration is the most common volumetric analysis method.

What is titration?

- Titration is a technique in which a solution of known concentration is used to determine the unknown concentration of another solution.
- The solution whose concentration is accurately known is called a **standard solution**.
- The **end point** is the point at which the indicator changes colour, signifying the end of the titration. This may not be the exact point where the reactants have completely reacted, but simply shows that a particular pH change has occurred.
- Indicators must be chosen carefully otherwise it will not be helpful in the titration. They must be chosen on the basis of pH at the equivalence point of the two reagents.

Note

Do not confuse end point with equivalence point.

The endpoint and the equivalence point are not exactly the same. The equivalence point is a theoretical concept and is determined by the stoichiometry of the reaction. This is where the titration should really end. The endpoint is the color change from the indicator and this is where the titration ends in practice. Therefore, the end point is a very close approximation of the equivalence point. The accuracy of the titration depends on how closely the experimental end point can be brought to the theoretical stoichiometric equivalence point.

- Two common indicators are phenolphthalein and methyl orange.

Indicator	Colour in Acid	Colour in Base	Colour at end-point	pH at colour change
Methyl Orange	Red	Yellow	Orange	3.1 - 4.5
Phenolphthalein	Colourless	Pink	Pale pink	8.0 - 10.0

- A primary standard is a salt or compound that is used to prepare a standard solution.

The requirements of a good primary standard:

- ☑ It should readily dissolve in water.
- ☑ It should be cheap and readily available.
- ☑ It should be of high purity.
- ☑ It should be stable in air, at room temperature and at moderately high temperature.
- ☑ It should have a high molecular mass so that small errors in weighing will be insignificant when doing calculations.

Equipment used for titration

1. Volumetric flask (or Standard flask)

- It is used for preparing the standard solution.
- The volumetric flask should only be rinsed with distilled water before preparing the standard solution to remove any impurities.

2. Standard pipette

- It is used for transferring an **exact volume** of the standard solution into the conical flask for titration.
- It should be rinsed with distilled water first to remove impurities and then with the solution it has to contain to prevent dilution.

An **aliquot** is an accurately measured sub-volume of the standard solution. It is usually taken out with a pipette from a volumetric flask.

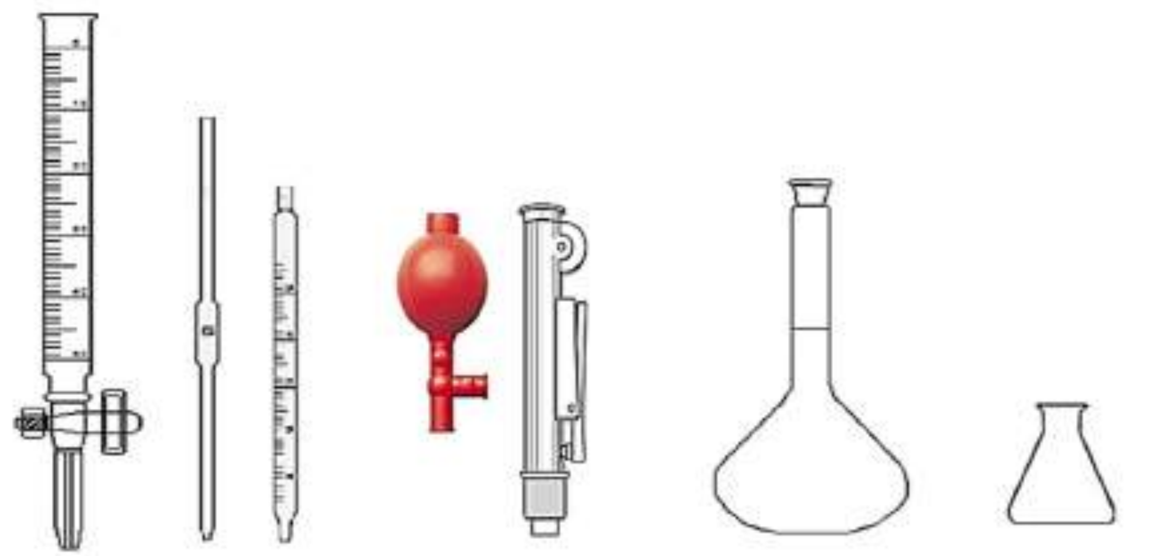
3. Burette

- It is used for transferring small volumes of liquids into the conical flask during titration.
- It should be rinsed with distilled water to remove impurities and then with the solution it has to contain to prevent dilution of the solution being titrated.

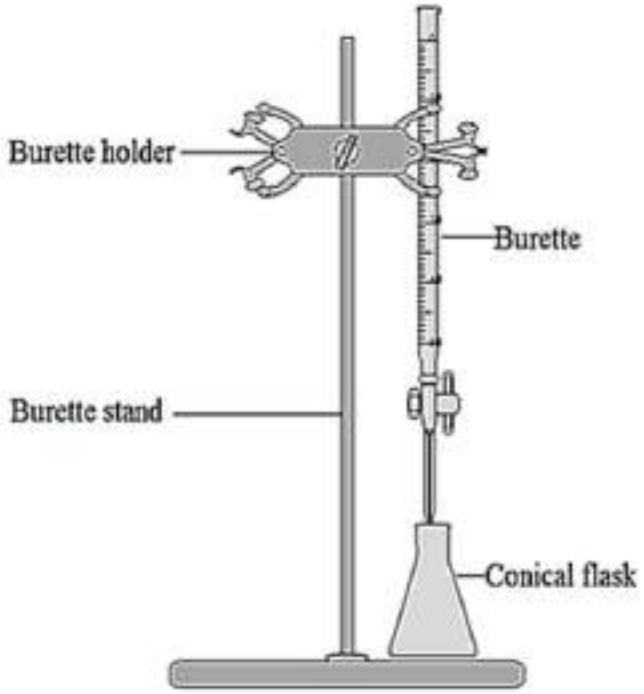
A **titre** is the volume delivered from a burette during titration sufficient to complete neutralisation.

4. Conical flask

- Holds the standard solution during titration. It should only be rinsed with distilled water to remove impurities.



Burette   Pipettes   Pipette fillers   Volumetric flask   Conical flask



A titration set-up

Preparing a standard solution

i. Accurately weigh out the required mass of the substance in a beaker on a top pan balance.
ii. Add distilled water from a wash bottle to dissolve it. Using a glass rod, stir until all the solid has dissolved.
iii. Using a clean funnel, transfer the solution to the volumetric flask. Rinse out the beaker and the glass rod properly into the volumetric flask with the wash bottle.
iv. Add water to just below the line on the volumetric flask. Top up with a Pasteur pipette or dropper to ensure that the bottom of the meniscus is on the line.
v. Put the lid on the volumetric flask and turn the volumetric flask over a couple of times to mix the solution.
vi. Label the solution with the name of the solution prepared, concentration and the date.

Performing titration

Pilot titration

- This titration is done rapidly at first to determine the approximate volume of the solution needed from the burette to reach end-point.
- The results from this titration should not be used for calculation purposes. The volume is just used as a guide for later titration.

Steps

1. Use a pipette to deliver a known amount of the solution into the conical flask which has been cleaned appropriately.
2. Add a few drops of indicator and put the conical flask on a white tile to see the colour change of the indicator more clearly.
3. Fill the burette with the titrant. Record the initial burette reading. (*Calculate the reading expected at the end point - this comes from the pilot titration*).
4. Add titrant from the burette to the flask, rapidly at first but stop about 1 mL before the expected endpoint reading.
5. Rinse the walls of the flask with a little bit of distilled water from the wash bottle.

> **Note:** You are concerned with the volume of titrant added to the flask, not the total volume of the flask; therefore, the added water will not affect your calculations.

6. Resume delivery of the titrant from the burette one drop at a time until the end point is seen.
7. Record the final burette reading.
8. Repeat the titration until at least three concordant results are obtained. *Concordant results are results which agree with each other.*

> **Note:** It is important to repeat the titration several times to check that the titre value is consistent to give reliable calculations and results. At least **three** titre values that fall within the range of 0.2 - 0.4 mL will be a good titration. The average titre values must be within 0.2 - 0.5 mL of the expected outcome. When done correctly and carefully, a titration will yield very precise results.

9. Calculate the average amount of the titrant needed to reach the endpoint.
10. Carry out the relevant calculation to determine the concentration of the solution.

Some skills required during titrations

1) Remove the air bubble from the nozzle of the burette before taking the initial reading. This will prevent taking inaccurate volume reading if the bubble escapes during a titration.
2) Ensure that there should not be any leakage from the burette during titration.
3) Use a funnel to fill the burette to prevent spillage.
4) Burette should be read to two decimal places to ensure accuracy. Burettes are usually graduated to 0.1 mL. The second decimal place of the burette reading is estimated, for example 12.25 mL.
5) The beaker should not touch the funnel when filling the burette.
6) Remember to remove the funnel from the burette before you actually begin titration to prevent accident as the funnel can fall during titration.
7) Take all readings from eye level to get accurate results.
8) Always read lower meniscus for colourless solution and upper meniscus for coloured solutions.
9) To expel the last drop of solution from the pipette, simply touch the inner surface of the conical flask with the nozzle of the pipette.
10) Swirl the conical flask continuously when adding solution from the burette.
11) It is not necessary to bring the burette reading to zero before each titration.
12) Before beginning the actual titration, always calculate the expected end point volume from the pilot titration.

The basic steps in performing acid-base titration calculations are:

1. Write the balanced chemical equation for the reaction.

   Note: A reaction between an acid and a base will give salt and water. However, if the base used is made up of carbonates, then CO_2 will also be produced.

2. Extract all the relevant data from the question.
3. Calculate the amount (in mol) of the substance whose concentration is known.
4. From the mole ratio of the substance with known concentration and the substance with unknown concentration, find the moles of the substance with unknown concentration.
5. Use the moles of the substance and volume of the solution with unknown concentration to find its concentration.

Example

In a titration experiment, it was found that 11.60 mL of 3.0 mol L^{-1} sulfuric acid was required to neutralize 25.0 mL of sodium hydroxide solution. Find the concentration of the sodium hydroxide solution used.

Solution

1. **Balanced chemical equation**

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

2. **Relevant information provided**

H_2SO_4	NaOH
• V= 11.60 mL= 0.0116 L • c = 3.0 mol L^{-1}	• V= 25.0 ml = 0.025 L • c = ?

Note: Convert all volumes to liters (L) for consistency during calculation.

3. **Calculate the number of moles of H_2SO_4**

$$c = \frac{n}{V}$$

$$n = c \times V$$

$= 3.0 \text{ mol L}^{-1} \times 0.0116 \text{ L}$

= **0.0348 mol**

It is recommended that in steps 3 and 4 you keep all the values you get in your calculator (or up till 4-5 decimal place value) and ***only round off your final answer*** *to the correct number of significant figures.*

4. **Compare mole ratio**

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

H_2SO_4 : NaOH

1 : 2

0.0348 : X

X= 0.0696 mol

Therefore 0.0696 mol of NaOH was used.

5. **Concentration of NaOH**

$$c = \frac{n}{V}$$

$$= \frac{0.0696 \text{ mol}}{0.025 \text{ L}}$$

= **2.78 mol L^{-1}**

For your practical, see the experiment on ***Preparation of a standard sodium carbonate solution; Standardisation of hydrochloric acid solution and Analysis of vinegar****; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji*

Exercise

1. Which of the following statements provide the correct information to accurately prepare a 0.2 mol L^{-1} sodium carbonate (Na_2CO_3) solution?

   A. Dissolve 10.6 g of anhydrous sodium carbonate in 500 mL of distilled water.
   B. Dissolve 28.6 g of anhydrous sodium carbonate in 1L of distilled water.
   C. Dissolve 21.6 g of anhydrous sodium carbonate in distilled water and make it up to 1 litre with more distilled water.
   D. Dissolve 57.2 g of anhydrous sodium carbonate in 500 mL of distilled water.

2. A student wished to find the concentration of some dilute hydrochloric acid (HCl) by titration. He pipetted 20 mL portions of 0.01 mol L^{-1} sodium hydroxide into conical flasks and titrated each of them with dilute HCl using a suitable indicator. Which of the following procedures is most likely to result in an incorrect value for the concentration of the acid?

   A. Rinsing the conical flasks with distilled water but not drying them before use.
   B. Rinsing the burette with only distilled water but not drying it before use.
   C. Rinsing the pipette with the sodium hydroxide solution before use.
   D. Rinsing the burette with distilled water and then with hydrochloric acid before use.

3. 2.65 g of sodium carbonate (Na_2CO_3) was dissolved in water and made up to 500 mL in a standard flask. 20 mL portions of this solution required 18.50 mL of a solution of hydrochloric acid when titrated using methyl orange indicator.

   (a) Write a balanced equation for the reaction between sodium carbonate and the hydrochloric acid.
   (b) Calculate:
      i. The concentration of the standard sodium carbonate solution in mol L^{-1}.
      ii. The concentration of the hydrochloric acid in mol L^{-1}.

4. 25.0 mL of 0.40 mol L^{-1} sodium hydroxide solution was placed in a well cleaned conical flask. Some phenolphthalein indicator was added to it. A pilot titration was carried out which indicated that about 17 mL of hydrochloric acid (HCl) was required for neutralisation. Three more titrations gave the titre readings as: 14.96 mL, 15.02 mL and 15.00 mL.

(a) Write a balanced equation for this reaction.
(b) Calculate the average titre volume in L.
(c) Calculate the amount (in mol) of hydroxide ions used in the titration.
(d) Calculate the concentration of the hydrochloric acid.

5. A standard solution of sodium carbonate was prepared by dissolving 10.6 g of the anhydrous salt in distilled water. The solution was then transferred to a standard flask and the total volume was adjusted to the 500 mL mark.

   (a) How many moles of sodium carbonate were dissolved in 500 mL of the solution?
   (b) Calculate the concentration of the sodium carbonate solution in mol L^{-1}.

   25 mL samples of the sodium carbonate solution were then titrated against a solution of hydrochloric acid. It was found that an average of 20.00 mL of the acid was required to react completely with 25.0 mL of the sodium carbonate solution.

   (c) Find the number of moles of the acid which reacted with one mole of carbonate ions.
   (d) Calculate the concentration, in mol L^{-1} of the hydrochloric acid.

6. 10.00 mL samples of 0.05 mol L^{-1} aliquot of a standard potassium carbonate (K_2CO_3) solution was pipetted into a conical flask. This sample was titrated with hydrochloric acid (HCl) solution of unknown concentration. The indicator, methyl orange showed that the end-point had been reached after an average titre of 25.00 mL of HCl was added.

   (a) What is the function of the indicator?
   (b) Determine the moles of potassium carbonate in the 10.00 mL sample.
   (c) Write a balanced equation for the reaction between hydrochloric acid and potassium carbonate.
   (d) How many moles of HCl are there in the 25.00 mL titre?
   (e) Determine the concentration of HCl used in this titration.

7. 20.0 mL of 0.21 mol L^{-1} sodium carbonate was titrated with hydrochloric acid solution. When the end point was reached, an average titre of 31.22 mL HCl had been added. What is the concentration of HCl used?

8. During the preparation of a standard solution of anhydrous sodium carbonate (Na_2CO_3), a student obtained the following results:

   Mass of beaker and anhydrous sodium carbonate = 131.10 g
   Mass of empty beaker = 128.45 g

   The student then dissolved the sodium carbonate in enough water to form 100.00 mL of solution.

   (a) Calculate the concentration of the solution prepared in: g L^{-1} and mol L^{-1}.
   (b) The standard solution was then titrated against a hydrochloric acid solution of unknown concentration using methyl orange indicator. It was found that 20.0 mL of the sodium carbonate solution was neutralised by 5.16 mL of the acid.

i. What piece of apparatus would be used to measure the 20.0 mL of standard solution into a conical flask for the titration?

ii. What piece of apparatus would be used to measure the volume of acid necessary to neutralize the standard solution?

iii. Briefly describe how it was known when the two solutions were neutralized?

iv. Write an equation that occurred between the hydrochloric acid and sodium carbonate.

v. Calculate the concentration of the hydrochloric acid solution in mol L^{-1}.

9. In a titration experiment, a student obtained the following titre values: 15.10 mL, 14.85 mL,15.20 mL, 15.10 mL and 14.80 mL. Which three best values should the student select for her titre values? What is the average titre in L?

10. During an experiment, a student needed to find out the concentration of sodium hydroxide (NaOH) solution which is in their school laboratory. He opted to use titration method. Find the concentration of the sodium hydroxide solution if 27.50 mL of 0.2 mol L^{-1} hydrochloric acid (HCl) was needed to titrate 25.0 mL of the sodium hydroxide solution.

11. During a titration experiment, a student placed 50.0 mL of 0.20 mol L^{-1} sodium hydroxide (NaOH) solution in a conical flask. He filled the burette with sulfuric acid (H_2SO_4) of unknown concentration. In the pilot titration, the student determined that the volume of acid needed to neutralize the sodium hydroxide was about 21.00 mL. Following that, he repeated the experiment three times and the following results were obtained.

Trial	**Volume of H_2SO_4 (mL)**
1	20.20
2	19.90
3	20.10

Determine the concentration of the sulphuric acid used.

12. 25.0 mL of barium hydroxide solution, $Ba(OH)_2$ was placed in a conical flask. The burette was filled with 0.062 mol L^{-1} nitric acid (HNO_3). A titration was carried out very quickly and it was found that about 40.00 mL of nitric acid was required to neutralize the barium hydroxide solution. Following that, three more titrations were performed to determine the volume of nitric acid required in the titration. The results obtained are as follows:

Trial	**Initial volume of HNO_3 (mL)**	**Final volume of HNO_3 (mL)**
1	48.90	9.80
2	47.50	8.50
3	49.00	10.10

Determine the concentration of barium hydroxide solution if the reaction equation is as follows:

$$\mathbf{Ba(OH)_2 + 2HNO_3 \rightarrow Ba(NO_3)_2 + 2H_2O}$$

3.2 OXIDATION-REDUCTION

Achievement indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Distinguish between oxidation and reduction reactions.
- ✓ Distinguish between some common oxidising and reducing agents.
- ✓ Balance redox equations.
- ✓ Describe and explain the electrolytic processes in the production of aluminium and copper metals.

The terms oxidation and reduction (abbreviated REDOX) have been used by chemist for many years.

- REDOX reactions take place around us every moment and include many diverse processes.
- In fact, they are directly linked to the origin of life. For instance, oxidation of nutrients forms energy which enables human beings, animals, and plants to survive.
- Corrosion of metals, combustion of fuels and smelting of mineral ores to their metals are examples of redox reactions.

Oxidation-Reduction Terminology

Term	Transfer of atoms	Transfer of electrons	Change in oxidation number
Oxidation	Gain of oxygen or loss of hydrogen	Loss of electrons	Increase in oxidation number
Reduction	Loss of oxygen or gain of hydrogen	Gain of electrons	Decrease in oxidation number
Oxidant (Oxidising agent)	Substance that loses oxygen or gains hydrogen	Substance that gains electron or an electron acceptor	Substance whose oxidation number has decreased
Reductant (Reducing Agent)	Substance that gains oxygen or loses hydrogen	Substance that loses electron or an electron donor	Substance whose oxidation number has increased

To make the definitions easy to remember in terms of electron transfer:

OIL RIG

OI**L** - **O**xidation is **L**oss of electrons
RI**G** - **R**eduction is **G**ain of electrons

Examples of common oxidising agents

Oxidising Agent			
Name	**Appearance**	**Use**	**Reduction Equation**
Oxygen (O_2)	- Colourless gas	- Combustion - Oxidation of metals	$O_2 \rightarrow O^{2-}$
Chlorine (Cl_2)	- Greenish yellow pungent gas	- Bleaching - Disinfectant	$Cl_2 \rightarrow 2Cl^-$
Permanganate ion (MnO_4^-)	- Purple coloured solution	- Breathalyzer test - Water treatment and disinfection - Synthesis of organic compounds - Oxidation of alcohols	$MnO_4^- \rightarrow Mn^{2+}$
Dichromate ion ($Cr_2O_7^{2-}$)	- Orange solution	- Chrome plating to protect metals from corrosion and to improve paint adhesion. - photographic screen printing - Breathalyzer test - Wood treatment - Sulfur dioxide test - Oxidation of alcohols	$Cr_2O_7^{2-} \rightarrow Cr^{3+}$
Hydrogen peroxide (H_2O_2)	- Colourless liquid	- Bleaching - Disinfectant/ Antiseptic	$H_2O_2 \rightarrow H_2O$
Dilute acids (H^+)	- Colourless	- Oxidation of metals	$H^+ \rightarrow H_2$

Examples of common reducing agents

Reducing Agent			
Name	**Appearance**	**Use**	**Oxidation Equation**
Metal e.g. Zn, Mg, Fe	- Silvery shiny metals	- Formation of metal oxides.	$Zn \rightarrow Zn^{2+}$ $Mg \rightarrow Mg^{2+}$ $Fe \rightarrow Fe^{2+}$
Carbon	- Black solid- Charcoal	-Used in the smelting process in the production of metals.	$C \rightarrow CO_2$
Sulphur Dioxide	- Colourless gas - Irritating smell - Gives colourless sulphite ion in solution	- Used as a preservative as it delays the oxidation of food by bacteria, as a fumigant and bleaching agent.	$SO_2 \rightarrow SO_4^{2-}$ $SO_3^{2-} \rightarrow SO_4^{2-}$
Ferrous ion	- Pale green solution	- Formation of iron oxides and hydroxides.	$Fe^{2+} \rightarrow Fe^{3+}$
Carbon monoxide (CO)	- Is a colourless and odorless	- Important industrial gas, which is widely used as a fuel.	$CO \rightarrow CO_2$

Example

In the reaction represented by the equation: $2Al^{+} + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$, determine the species that is oxidized.

i. $Al \rightarrow Al^{3+} + 3e^-$

Losing electrons so Oxidation

ii. $Cu^{2+} + 2e^- \rightarrow Cu$

Gaining electrons so Reduction

- Thus the species that is oxidised is Al.

3.2.2 Oxidation number (Oxidation state)

- An oxidation number is a number that is assigned to an element in a chemical reaction to show the total number of electrons which have been removed from an element (a positive oxidation state) or added to an element (a negative oxidation state) to get to its present state.
- Assigning oxidation numbers simplifies the process of working out what is being oxidized and what is being reduced in redox reactions.
- Oxidation state is commonly used to determine the changes in redox reactions and is mostly numerically similar to valence electrons.

Rules for assigning oxidation number

1. The oxidation number of an atom is zero in a neutral substance that contains only one type of element.
   Example: For O_2 and Mg, the oxidation number is 0.
2. The oxidation number of each oxygen atom in a compound is -2, except in peroxides (e.g. H_2O_2) where the oxidation number is -1.
3. The oxidation number of each hydrogen atom in a compound is +1except in metallic hydrides (example in LiH, NaH, CaH_2, and $LiAlH_4$) where it is -1.
4. The oxidation number of Group I metals is +1.
5. The oxidation number of Group II metals is +2.
6. The sum of the oxidation number in a neutral molecule is equal to zero (0).
   Example: For $C_6H_{12}O_6$, the oxidation number is 0.
7. The oxidation number of an atom in a monoatomic ion is equal to the charge on the ion.
   Example for: Na^+ = +1
   Cl^- = -1
8. The sum of the oxidation numbers in a polyatomic ion is equal to the charge on the ion. Example for H_3O^+, the oxidation number is +1.

Note: Always include '+' or '-' sign before the number to indicate oxidation states.

Example

Calculate the oxidation number of Sulphur in:

i) SO_3
ii) SO_4^{2-}
iii) H_2SO_4

Solution

(i) Sulphur in SO_3

$1(S) + 3(O) = 0$

$S + 3(-2) = 0$

$S - 6 = 0$

$\underline{S = +6}$

(ii) Sulphur in SO_4^{2-}

$1(S) + 4(O) = -2$

$S + 4(-2) = -2$

$S + -8 = -2$

$\underline{S = +6}$

(iii) Sulphur in H_2SO_4

$2(H) + 1(S) + 4(O) = 0$

$2(+1) + S + 4(-2) = 0$

$+2 + S + -8 = 0$

$S - 6 = 0$

$\underline{S = +6}$

Exercise

1. Find the oxidation number of chromium (Cr) in:

   i. $K_2Cr_2O_7$ ii. CrO_4^{2-} iii. $Cr_2O_7^{2-}$

2. Find the oxidation number of carbon in:

   i. CO_2 ii. CO_3^{2-}

3. Find the oxidation number of hydrogen in:

   i. H_2O ii. NaH iii. H_2O_2

4. Identify the following reactions as either oxidation or reduction:

   i. $Cr_2O_7^{2-} \rightarrow Cr^{3+}$
   ii. $MnO_4^{-} \rightarrow Mn^{2+}$

5. The oxidation state of chlorine in $HClO_4$ is :

   A. -1
   B. +7
   C. -5
   D. -7

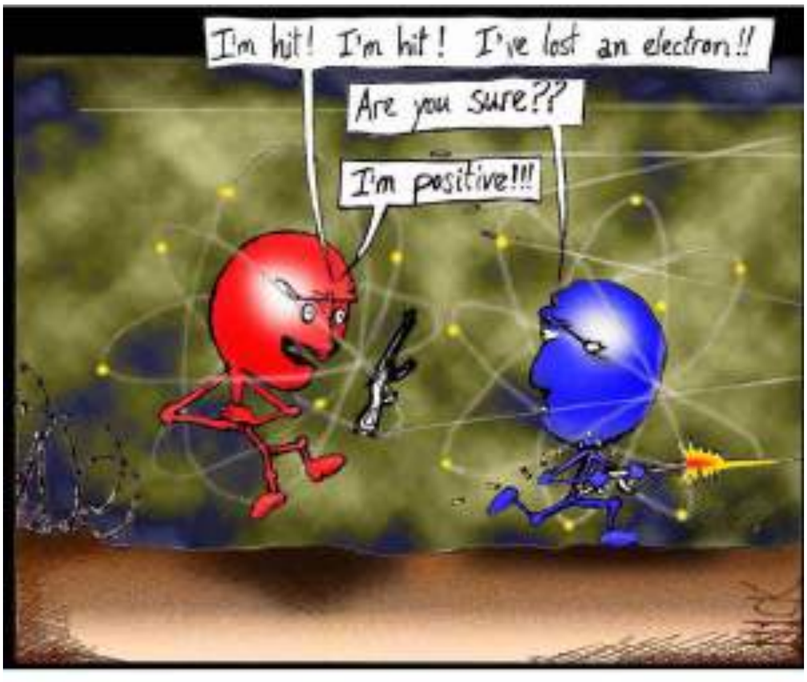
3.2.3 Balancing REDOX equations

Before actually beginning to balance redox reactions, for each reaction you should be able to:

i. Identify the species which is reacting but should not be involved in the redox equation.
ii. Decide what the products are for each reacting species.
iii. Work out which part is oxidation and which is reduction.
iv. Formulate two half-equations.

Rules for balancing redox equations

1. Obtain the half-equations.
2. *For each half-equation:*
   i. Balance all atoms except oxygen and hydrogen.
   ii. Balance oxygen by adding water molecules (H_2O).
   iii. Balance hydrogen by adding hydrogen ions, H^+.
   iv. Balance charges by adding electrons.
3. Add the balanced half-equations so that the electrons cancel out.
4. Identify and cancel out terms common to both sides of the equation.



Another casualty in the War of the Atoms.

Source: http://start.sd34.bc.ca/

Example

Balance the following redox equation.

$SO_2 + NO_3^- \rightarrow SO_4^{2-} + NO$

Solution

Reductant: SO_2

Oxidant: NO_3^-

Reducing half-equation	Oxidising half-equation
$NO_3^- \rightarrow NO$	$SO_2 \rightarrow SO_4^{2-}$
Balancing	Balancing
$NO_3^- \rightarrow NO + 2H_2O$	$SO_2 + 2H_2O \rightarrow SO_4^{2-}$
$NO_3^- + 4H^+ \rightarrow NO + 2H_2O$	$SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+$
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	$SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$

Add both the equations

$[NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O] \times \mathbf{2}$

$[SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-] \times \mathbf{3}$

Multiply both equations by the smallest possible whole number **to cancel out the electrons.**

$2NO_3^- + 8H^+ + 6e^- \rightarrow 2NO + 4H_2O$

$3SO_2 + 6H_2O \rightarrow 3SO_4^{2-} + 12H^+ + 6e^-$

$2NO_3^- + \cancel{8H^+} + \cancel{6e^-} \rightarrow 2NO + \cancel{4H_2O}$

$3SO_2 + \cancel{6H_2O} \rightarrow SO_4^{2-} + \cancel{12H^+} + \cancel{6e^-}$

$2H_2O \qquad 4H^+$

Cancel out terms common to both sides of the equation and identify the "balance" remaining as shown below.

Combine the two equations:

$2NO_3^- + 3SO_2 + 2H_2O \rightarrow 2NO + 3SO_4^{2-} + 4H^+$

Check that the number /types of atoms and the charge are equal in both sides of the equation.

*For your practical, see the experiment on **Electron transfer in oxidation and reduction**; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Exercise

1. Consider the equations shown below and answer the questions that follow.

   a) $Br^- \rightarrow Br_2$

   b) $MnO_4^- \rightarrow Mn^{2+} + H_2O$

   c) $Fe^{3+} \rightarrow Fe^{2+}$

   d) $I_2 \rightarrow I^-$

   i. Balance each of the ion-electron equations.
   ii. State whether each equation shows oxidation or reduction.

2. Balance the following half equations.

   a) $SO_3^{2-} \rightarrow SO_4^{2-}$
   b) $SO_2 \rightarrow SO_4^{2-}$
   c) $NO_3^- \rightarrow NO_2$
   d) $MnO_4^- \rightarrow MnO_2$
   e) $Cr_2O_7^{2-} \rightarrow Cr^{3+}$

3. Balance the equation shown below which shows sulphur dioxide reacting with acidified dichromate ions.

   $$SO_2 + Cr_2O_7^{2-} \rightarrow SO_4^{2-} + Cr^{3+}$$

4. Balance the following equations. In each case, give the balanced half-equations and combine the balanced half-equations to give the overall reaction equation.

   a) $Fe^{2+} + MnO_4^- \rightarrow Fe^{3+} + Mn^{2+}$
   b) $H_2S + SO_2 \rightarrow S + H_2O$
   c) $Cu + NO_3^- \rightarrow NO_2 + Cu^{2+}$
   d) $Cl^- + H_2O_2 \rightarrow Cl_2 + H_2O$

5. The quantity of ethanol in breath can be found by breathing out through a tube-containing potassium dichromate and an acid.
   i. Write the balanced ion-electron half equation for the conversion of acidified dichromate ions ($Cr_2O_7^{2-}$) to chromium (III) ions (Cr^{3+}).
   ii. Is the conversion of dichromate ion to chromium (III) ion, oxidation or reduction?

6. Consider the ion-electron equation shown below.

   $$MnO_4^- + xH^+ + ye^- \rightarrow Mn^{2+} + zH_2O$$

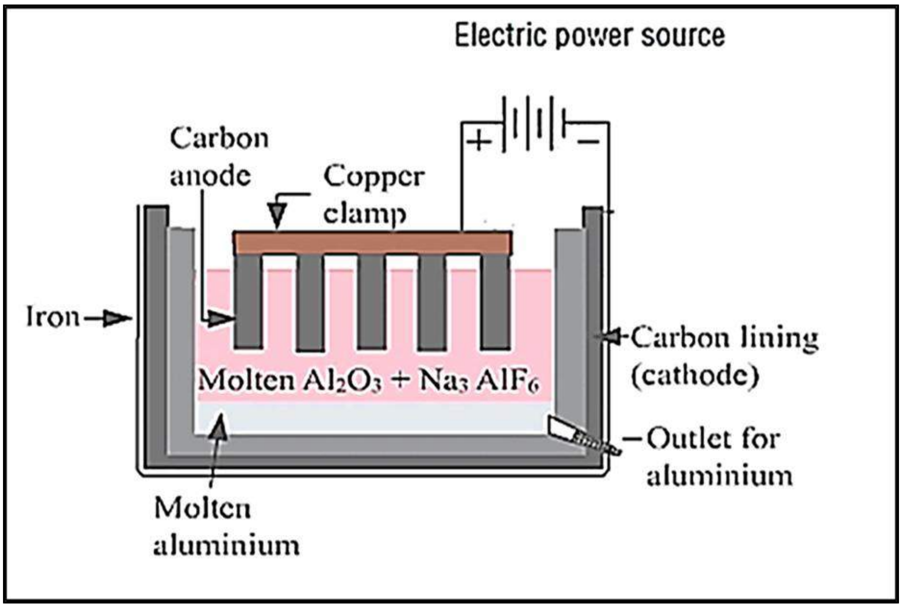
   Determine the numerical values of x, y and z.

3.2.4 Application of redox reactions

1. Production of Aluminum

- Aluminum is a light, conductive and corrosion resistant metal with a strong affinity for oxygen.
- Due to these properties, it is now a widely used material, with applications in the aerospace, architectural construction and marine industries. It also has many domestic uses.
- Aluminium occurs naturally as the mineral bauxite which is primarily a mixture of $Al_2O_3.3H_2O$, Fe_2O_3 and SiO_2.
- Alumina (also known as aluminium oxide, Al_2O_3) is extracted from bauxite and purified through electrolytic reduction as shown below.

Electrolytic reduction of Alumina



Source: http://www.askiitians.com/

- Alumina (Al_2O_3) melts at around 2072 ^oC which is too high for practicality. Therefore, it is dissolved in molten cryolite (Na_3AlF_6), forming an ionic solution. Molten cryolite acts as a relatively low melting solvent for alumina and also helps to conduct the current.
- This solution is decomposed by electrolysis, using a consumable carbon anode.
- The reduction is carried out in rows of cells made of iron. The iron tanks are lined with carbon, which acts as the cathode. Carbon rods dipping in solution act as anodes.

- At the anode, oxygen is liberated, and at the high temperature of the cell (~1000 ^oC) oxidises the carbon anodes to carbon dioxide (some carbon monoxide will also be formed). For this reason, the anodes have to be replaced during the reduction process and that is why it is said to be consumable.
- At the cathode, alumina is reduced to aluminum. The aluminum ions accept electrons and form a layer of molten aluminum metal.

Summary

Ionisation of Alumina

$2Al_2O_3 \rightarrow 4Al^{+3} + 6O^{2-}$

- **Reaction at the cathode:**

$4Al^{+3} + 12e^- \rightarrow 4Al$ } Reduction

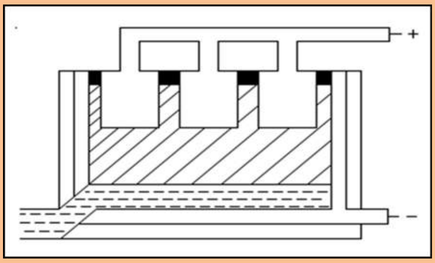
- **Reaction at the anode:**

$6O^{2-} \rightarrow 3O_2 + 12e^-$ } Oxidation

$C + O_2 \rightarrow CO_2$

Exercise

1. Briefly explain why alumina (Al_2O_3) is dissolved in molten cryolite (Na_3AlF_6).
2. What is another name of alumina?
3. In the purification of alumina, why the graphite (carbon) anode is usually replaced from time to time.
4. Briefly describe the process(s) which occurs during the electrolysis of molten alumina (Al_2O_3)?
5. With the help of balanced half-equations, briefly describe the processes occurring at the anode and cathode during the electrolysis of alumina.
6. Give some uses of aluminium based on its properties.
7. Label the diagram below showing the electrolysis of alumina.



Labels list

Graphite anode; Graphite cathode;
molten aluminium; molten electrolyte

2. Production of Copper

- Copper is extracted from its ore.
- An ore is a rock or mineral that has enough required metal in it to make it worth extracting from.
- For copper, it is worth extracting when there is about 2 kg of copper per 1000 kg of ore.
- The method used to extract copper from its ores depends on the nature of the ore.
- Some of the sources of copper include: Chalcocite (Cu_2S), malachite ($CuCO_3.Cu\ (OH)_2$), azurite ($2CuCO_3.Cu\ (OH)_2$), cuprite (Cu_2O) and chalcopyrite or copper pyrite ($CuFeS_2$).
- The major source of copper is chalcopyrite.
- After copper is mined from its ores, it is purified through electrodeposition.

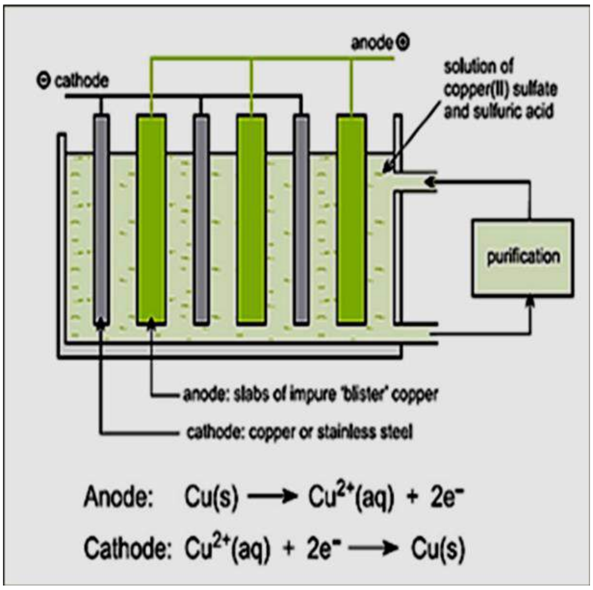
Purification of copper by electrodeposition

- The electrolytic refining of copper produces the high quality and high purity copper required by industries.
- The blister copper anodes are immersed in an electrolyte containing copper sulphate and sulphuric acid.

> The name 'blister' copper comes from the fact that this copper has bubbles of sulphur dioxide on the surface.

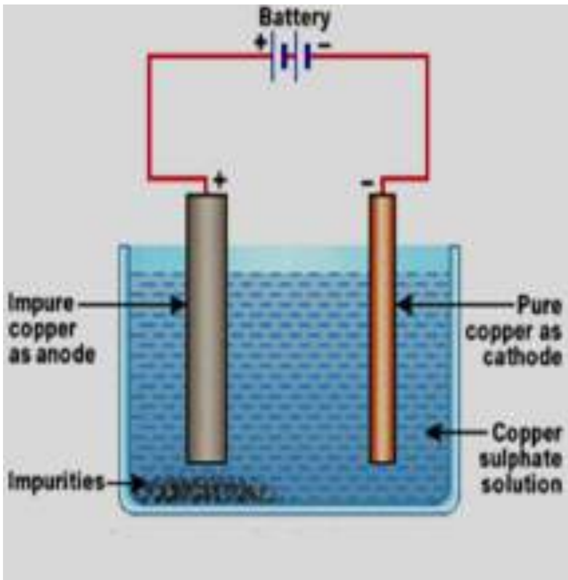
- Pure copper cathodes are arranged between the blister copper anodes and a current of over 200 A passes through the solution.

Industrial set-up for refining of copper



Source: http://www.essentialchemicalindustry.org

Laboratory set-up for refining of copper



Source:http://look4chemistry.blogspot.com

- When the electric current is passed through the electrolyte ($CuSO_4$), the Cu^{2+} in the electrolyte is reduced to copper metal (Cu) and gets deposited at the cathode.
- The anode dissolves to replace the Cu^{2+} in the solution.
- Periodically, the cathodes are removed and pure copper is scraped off.

Summary

- At the anode: $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$ } Oxidation
- At the cathode: $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ } Reduction

What happens to the impurities?

Gold, silver, platinum and tin are insoluble in the electrolyte and so do not deposit on the cathode. They form a valuable sludge that collects under the anode. Soluble impurities such as iron, arsenic, antimony, bismuth and nickel dissolve in the electrolyte.

Exercise

1. Briefly explain why the name 'blister copper' is given to the name of the copper used in electrodeposition.
2. Draw a diagram of the electrodeposition of copper and label the following components:

   Blistered copper, pure copper metal, acidified copper(II) sulphate solution, anode, cathode, cell

3. Write down the reaction half-equation which occurs at the anode during the electrolytic refining of copper.
4. Write down the reaction half-equation which occurs at the cathode during the electrolytic refining of copper.
5. What happens to the impurities during the electro deposition process?

Research activity

* **Collect and present information on aluminium, gold and copper exploration and extraction in Fiji with respect to the following:**
  - **Location of extraction or exploration site(s).**
  - **Type of ore extracted.**
  - **Extraction methods for the metals (if any).**

3.3 PHYSICAL CHEMISTRY

Achievement indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Describe and explain the effect of different factors on the rate of reactions using collision theory.
- ✓ Distinguish between exothermic and endothermic reactions.
- ✓ Calculate heat changes in thermochemical reactions.
- ✓ Explain the factors that affect an equilibrium using the Le Chatelier's principle.
- ✓ Apply the Le Chatelier's Principle to the Haber process.
- ✓ Describe the equilibrium reactions in the aqueous systems.
- ✓ Describe and compare the properties, strengths and reactions of acids and bases.

3.3.1 Rates of Chemical Reactions

- The rate of a reaction indicates how rapidly the products are formed from the reactants.
- Knowing about reaction rates is very important in chemical industries.
- For instance, chemical companies need to make their products as quickly and as cheaply as possible. Knowing the rate of the reactions helps the companies to achieve this.
- The rate of a reaction can be found by:
  - i. Measuring how quickly the reactants are used up.
  - ii. Measuring how quickly the products are formed.

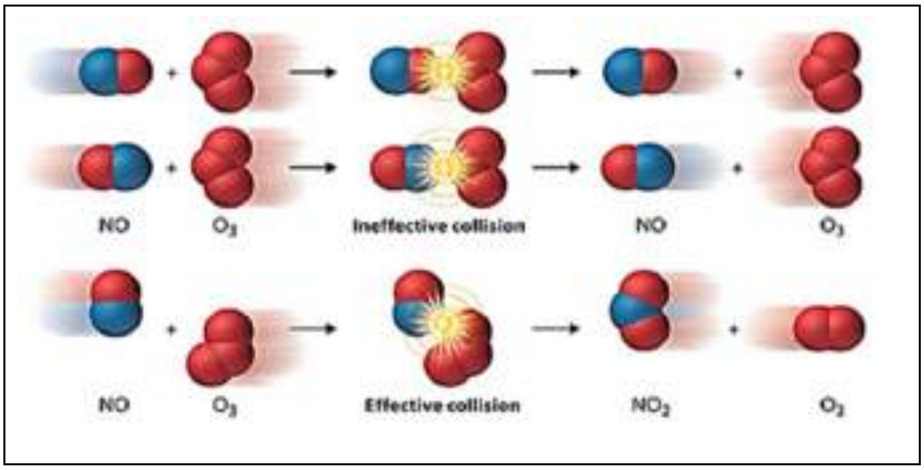
Collision theory and rates of reaction

- Particles in a gas or liquid are constantly moving from place to place and colliding with each other.
- Most chemical reactions occur by the collision of reacting particles.
- The collision theory states that the rate of a chemical reaction is proportional to the number of collisions between reactant particles.
- Effective collisions are those that result in a chemical reaction.

The requirements for an effective collision are:

1. The reactants must frequently collide with each other. The more the collision, the faster the reaction.
2. The particles must have sufficient energy to start the reaction (*activation energy).* This means that the total kinetic energy of the colliding particles must be greater than the activation energy, which is the minimum energy required for a reaction to occur.
3. The reacting particles must collide in proper orientation to form new products. This enables the existing forces or bonds between particles to be broken so that new bonds can form.

An illustration of effective and ineffective collision



Source: http://www.askiitians.com

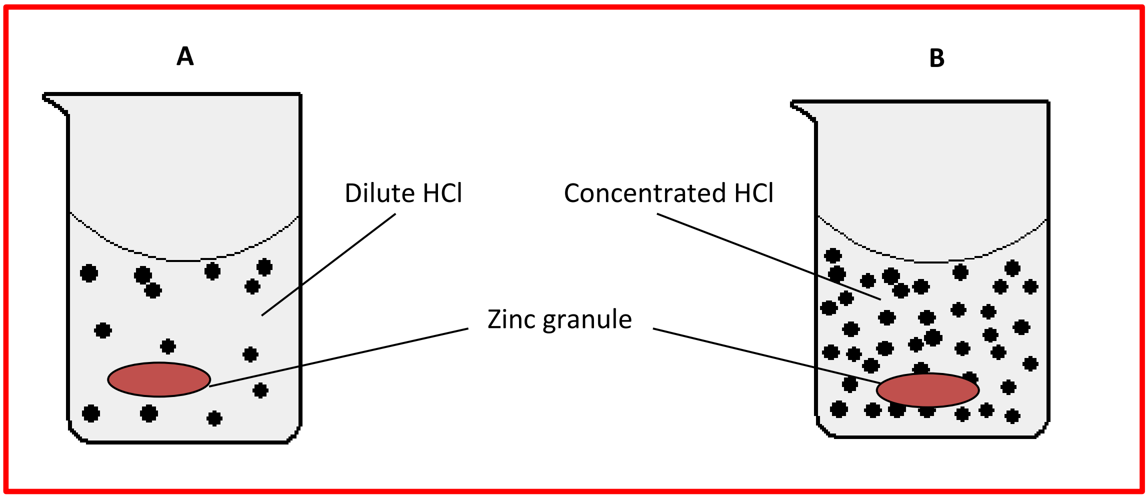
Factors affecting reaction rates

- The rate of a chemical reaction depends on factors such as temperature, concentration of the reactants, the surface area of any solid reactants, the use of catalysts and pressure.

1. Concentration

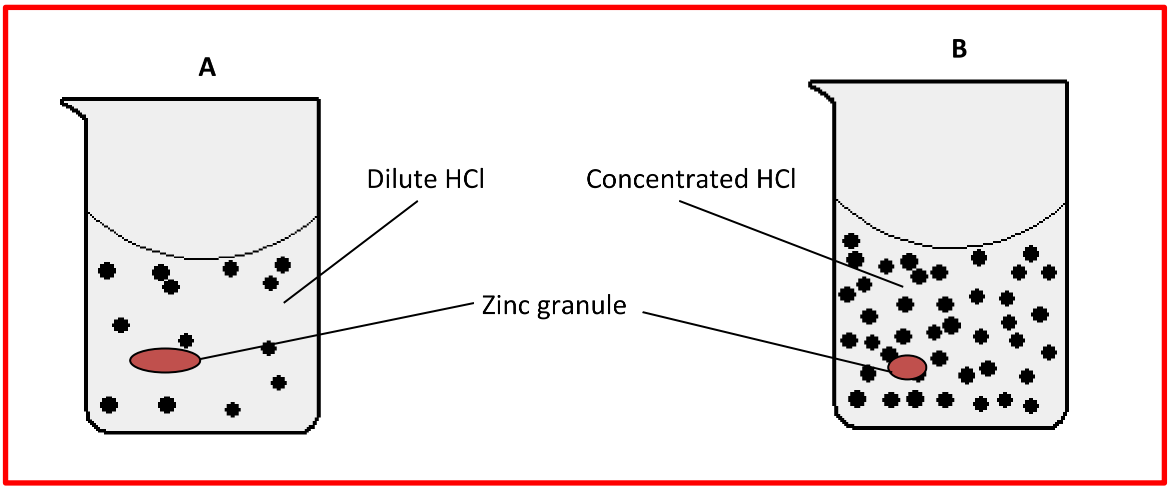
Increase in concentration of the reactants increases the rate of chemical reactions. This is because the more the reacting particles, the more the collisions occurring between them, thus higher chances of effective collision.

Example



This illustration shows that zinc granules (of same size) being placed in two different beakers of HCl solutions (dilute and concentrated solutions).

After 10 minutes



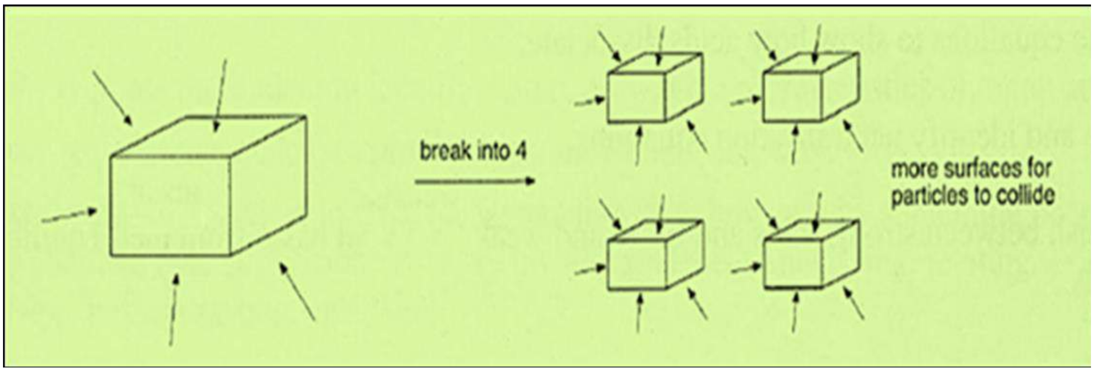
The illustration above shows that the zinc granule reacted faster in concentrated HCl than in the same volume of dilute HCl at a given time. This is because the concentrated HCl had more HCl particles to collide with the zinc granule, thus more chances of effective collision, which increased the rate of the reaction.

2. Temperature

Increase in temperature increases the rate of chemical reactions. This is because when temperature increases, the reacting particles gain energy and move faster, resulting in greater chances of effective collision between them to form products. For most reactions, increasing the temperature by 10 ^{o}C will approximately double the reaction rate.

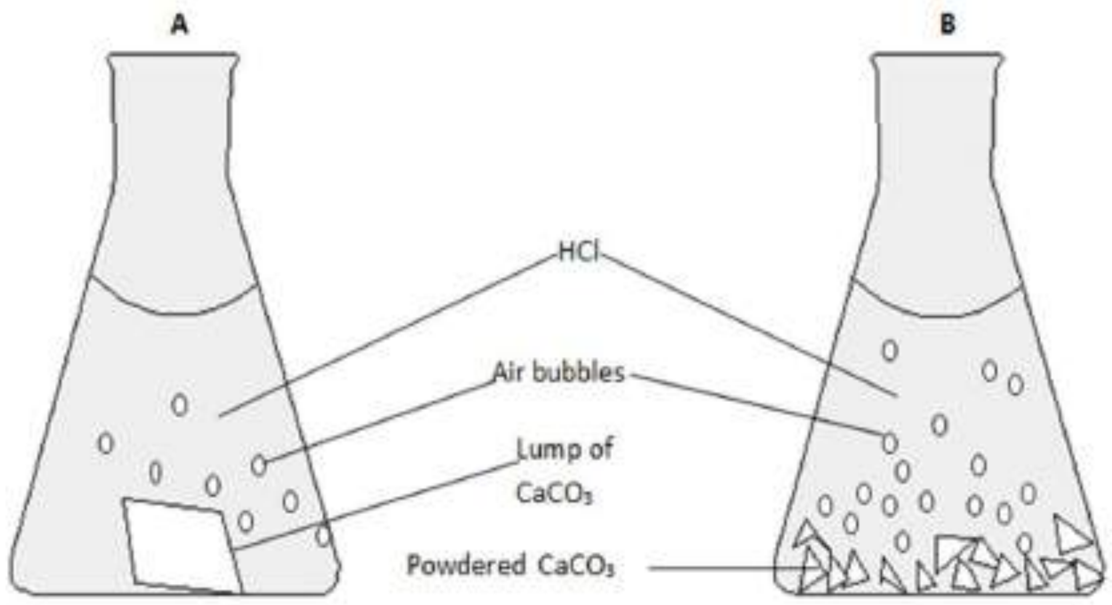
3. Surface Area

A solid in a solution can only react when particles in the solution collide with the surfaces of the solid. Increase in surface area of the reacting particles increases the rate of chemical reactions. This is because when the surface area is increased, more particles will be exposed to collide amongst each other at a given time and create more effective collision.



Surface area of a solid can be increased by breaking it up into smaller pieces.

Example



This illustration shows that a lump of calcium carbonate ($CaCO_3$) reacts in HCl much slower than powdered $CaCO_3$ of the same mass. This is characterised by a slow release of air bubbles (carbon dioxide gas).This is because powdered $CaCO_3$ has a larger surface area, thus greater number of effective collisions between the $CaCO_3$ particles and HCl, which increases the rate of reaction.

4. Catalyst

A catalyst is a substance that increases the rate of chemical reaction by providing an alternative pathway for the reaction; a pathway that has lower activation energy. This allows greater number of colliding particles to have a kinetic energy greater than the new activation energy. The catalyst is not used up during the reaction.

5. Pressure

If the reactants are gases, at higher pressure, there will be more particles per unit volume. Therefore more particles will collide per second and the rate of reaction will increase.

*For your practical, see the experiment on **Temperature and reaction rate**; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Exercise

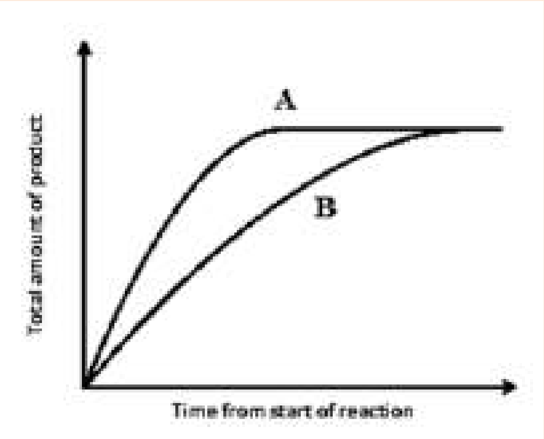
1. Give two ways to measure rate of a reaction.

2. List the factors which can affect a reaction rate.

3. Briefly describe what the collision theory states about chemical reactions.

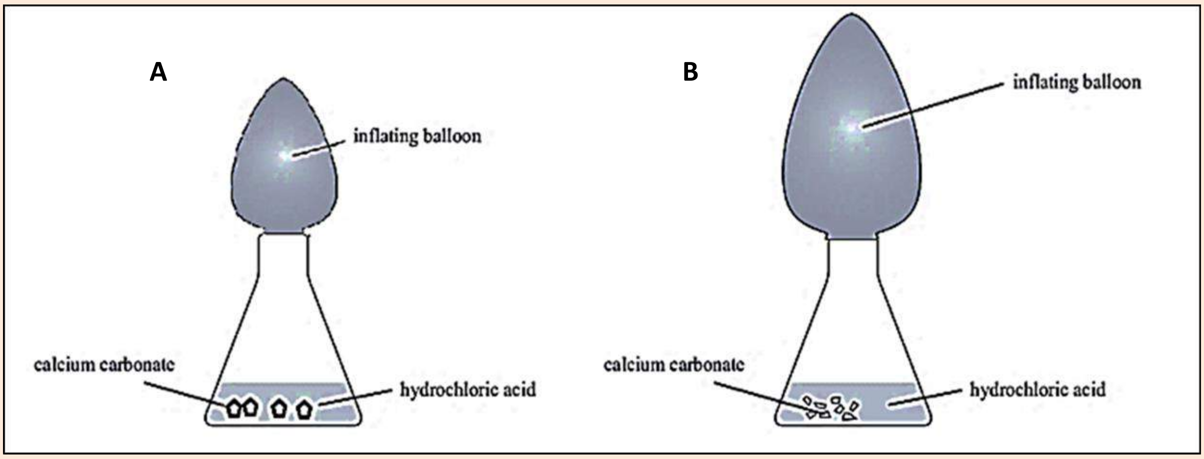
4. Some samples of zinc were placed in some hydrochloric acid solution at a variety of concentrations and temperatures. In which of the following cases was the rate of reaction: (i) the highest and (ii) the lowest?
   A. 2.0 g of zinc reacted in 10 minutes.
   B. 0.20 g of zinc reacted in 1 minute.
   C. 0.50 g of zinc reacted in 1 minute.

5. Consider the diagram given below illustrating the effect of temperature on the rates of reaction.



i. Which reaction (A or B) was subjected to a higher temperature?
ii. Briefly justify your answer to (i) above using the collision theory.

6. A student wished to study the rate of reactions in the laboratory. In one flask (**A**), he placed big pieces of calcium carbonate ($CaCO_3$). In a similar sized flask (**B**), he placed powdered calcium carbonate [*of the same mass as placed in flask (A)*]. Following that, he added hydrochloric acid to both the flask and immediately, a balloon was placed on top of the flasks. The set – up was left for 2 minutes and it was seen that the balloons had inflated. Briefly explain the observation from the two flasks.



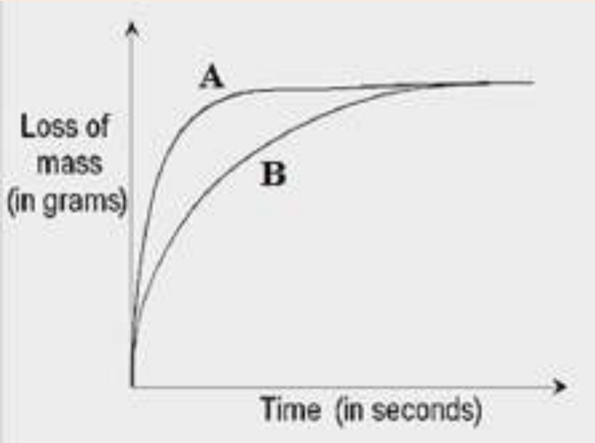
Hint: $CaCO_{3(s)} + 2HCl_{(l)} \rightarrow CaCl_{2(s)} + H_2O_{(aq)} + CO_{2\,(g)}$

7. Explain the following observations:
   a) When 6.0 g of powdered marble is put into 100 mL of 1.0 mol L^{-1} hydrochloric acid solution, the reaction is rapid and there is very little decrease in rate as all of the marble reacts. When 6.0 g lump of marble is put into a similar sample of hydrochloric acid solution, two observations are made: (i) The reaction is slower than the reaction of powdered marble and (ii) the reaction becomes slower and slower during the course of the reaction and is extremely slow by the time that the last traces of marble are left.
   b) If magnesium is placed in cold water, hardly any evidence of reaction can be seen. However, if the water is boiled, the reaction becomes quite vigorous.

8. In aqueous solution, ethanol and ethanoic acid reacts to from ethyl ethanoate.
   (i) According to the collision theory, does every collision between an ethanol molecule and an ethanoic acid molecule lead to the formation of an ethyl acetate molecule? Explain your answer in terms of the collision theory.
   (ii) What happens to the rate of reaction if the temperature is raised? Explain your answer in terms of the collision theory.

9. (a) Describe three ways to slow down the rate of gas production when zinc granules are allowed to react with dilute hydrochloric acid.
   (b) Use collision theory to explain each of your responses to (a).

10. A student reacted some marble chips with HCl in two separate beakers of the same size. Assume that the time of reaction, mass of the marble chips, concentration of HCl and the temperature at which the reaction took place were the same for both the beakers. The student followed the reaction rate and plotted a graph of the loss of mass of the marble chips against time. Determine which graph shows that smaller pieces of marble chips were used. Explain your answer.



3.3.2 Endothermic and Exothermic reactions

- A chemical reaction is always accompanied by a change in energy (enthalpy change).
- Energy change occurs in the form of absorbing heat from the surrounding or releasing heat to the surrounding.
- The general rule is:
  - **Breaking** chemical bonds **needs** energy.
  - **Forming** chemical bonds **releases** energy.
- If more energy is required to break existing bonds than the energy that is released when new bonds are formed, heat is absorbed. This is called an **endothermic reaction**. Endothermic reactions are recognised by the temperature of the surrounding decreasing. Examples of endothermic reactions include thermal decomposition, base dissolving in water, melting, ammonium chloride and potassium nitrate dissolving in water.
- If in the formation of new bonds more energy is released than is required to break the existing bonds, heat is given out. This is called an **exothermic reaction**. Exothermic reactions are recognized by the temperature of the surrounding increasing. Examples of exothermic reactions include combustion, respiration, neutralisation, dissolving acids in water, freezing and corrosion of metals.
- The symbol *H* is used to represent the **enthalpy** (heat content) of a chemical.

- The enthalpy of a chemical cannot be measured by itself; however, the **change in enthalpy** (ΔH) during a reaction can be measured.

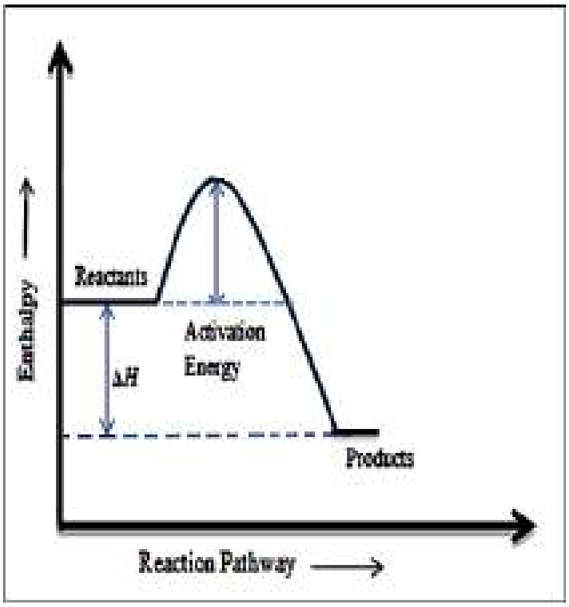
Change in enthalpy = Energy of products – Energy of reactants

$$\Delta H = H_p - H_r$$

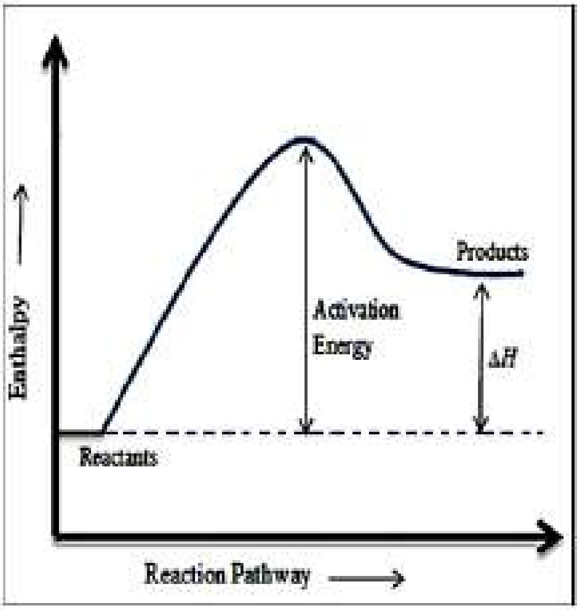
Note: If the ΔH is **negative,** the reaction is **exothermic** since heat energy is released.

If the ΔH is **positive**, the reaction is **endothermic** since heat energy is absorbed.

The amount of heat energy being absorbed or released depends on the strength of the bonds being broken or formed respectively.



Energy profile diagram of an exothermic reaction



Energy profile diagram of an endothermic reaction

- In this diagram, the reactants have more energy than the products, so ΔH will be a **negative** value, thus the reaction is exothermic.
- The activation energy (E_a) is measured from the reactant energy level to the peak of the curve.

- In this diagram, the reactants have **less** energy than the products so ΔH will be a **positive** value, thus the reaction is endothermic.
- The activation energy (E_a) is measured from the reactant energy level to the peak of the curve.

Activation Energy (E_a) - is the minimum amount of energy required by the reactants to start up a chemical reaction.

Significance of activation energy (E_a)

A small E_a means a faster rate of reaction.
A large E_a means a slower rate of reaction.

Comparison of exothermic and endothermic reactions

Exothermic	Endothermic
Heat is released to the surrounding	Heat is absorbed from the surrounding
Temperature of surrounding increases	Temperature of surrounding decreases
Enthalpy of products is less than enthalpy of reactants	Enthalpy of products is more than enthalpy of reactants
Enthalpy change (ΔH) is negative since $\Delta H = Hp - Hr, \Delta H < 0$	Enthalpy change (ΔH) is positive since $\Delta H = Hp - Hr, \Delta H > 0$
Reaction container becomes warm	Reaction container becomes cool

3.3.3 Enthalpy Change

- The enthalpy change is the energy change occurring during a reaction.
- This unit of measurement is useful for calculating the amount of energy per mole either released or produced in a reaction.

Example

1. $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ ΔH = -390 kJ/mol

   ✓ The equation shows that 390 kJ of heat energy is given out when 1 mole of carbon dioxide is formed from its elements.

   ✓ This reaction is exothermic since the ΔH is negative.

2. $\frac{1}{2}$ $H_{2\,(g)} + \frac{1}{2} I_{2\,(g)} \rightarrow HI_{\,(l)}$ ΔH = + 25.9 kJ/mol

   ✓ The equation shows that 25.9 kJ of heat energy is absorbed from the surrounding when 1 mole of HI is formed from its elements.

   ✓ This reaction is endothermic since the ΔH *is* positive.

Note:

- The overall energy change during a reaction is known as the enthalpy change (ΔH) of the reaction.
- The unit for enthalpy change is kJ/mol.
- The unit for energy is kJ.

Worked example for calculating enthalpy changes

Using the given reaction equation, calculate the amount of heat released when 6 g of carbon undergoes combustion.

$C_{(s)} + O_{2\,(g)} \rightarrow CO_{2\,(g)}$ ΔH = -393.5 kJ/mol

Solution

The given equation shows that when 12 g (1 mole) of carbon undergoes combustion, 393.5 kJ of heat energy is released.

Thus, 12 g : - 393.5 kJ

6 g : X

$$X = \frac{6\ \text{g} \times -393.5\ \text{kJ}}{12\ \text{g}}$$

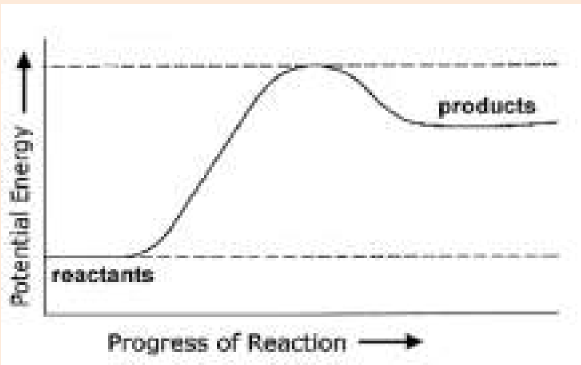
= - 196.75 kJ

Therefore the ΔH when 6 g of carbon undergoes combustion is **-196.75 kJ.**

*For your practical, see the experiment on **Energy changes**; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Exercise

1. What is activation energy?
2. Give some differences between endothermic and exothermic reactions.
3. Consider the diagram given below and answer the questions that follow.



i. Is the reaction endothermic or exothermic? Justify your answer.
ii. Label the Activation energy (E_a) and the enthalpy change (ΔH).

4. Consider the reaction given below and answer the questions that follow.

$$S_{(s)} + O_{2\ (g)} \rightarrow SO_{2\ (g)} \qquad \Delta H = -\ 296.8\ \text{kJ/mol}$$

   i. Is the reaction above endothermic or exothermic? Give a reason for your answer.
   ii. Calculate the heat energy released when 65 g of sulphur is burnt.

5. Using the reaction equation given below, calculate the amount of heat energy released when 8 g of methane undergoes combustion.

$$CH_{4\ (g)} + 2O_{2\ (g)} \rightarrow CO_{2\ (g)} + 2H_2O_{(g)} \qquad \Delta H = -\ 895\ \text{kJ/mol}$$

6. Consider the following equation which represents the burning of carbon monoxide:

$$2CO_{(g)} + O_{2\ (g)} \rightarrow 2O_{2\ (g)} \qquad \Delta H = -\ 566\ \text{kJ/mol}$$

   Calculate the enthalpy change when one mole of carbon monoxide burns.

3.3.4 Chemical Equilibrium

- Reactions can proceed in one direction or both directions. Reactions which proceed in one direction go to completion while reactions which proceed in both directions (forward and backward) are often reversible.
- When a system is in dynamic equilibrium, the rate of the forward reaction is equal to the rate of the backward reaction.
- However, reversible reactions can be made to go to completion by changing the position of the equilibrium.
- This change can be brought about by the **Le Chatelier's principle**, which states that if a change is applied to a system at dynamic equilibrium**,** the position of the equilibrium shifts to counteract the change and reestablish equilibrium.
- Change in a system can be brought about by changes in concentration, temperature and pressure.

Factors that change the position of the equilibrium

1. Concentration

Consider the reaction: $A + B \rightleftharpoons C + D$

- If one or both the reactants (A or B) are added at equilibrium, the forward reaction will be favored. Therefore, the rate of forward reaction will increase. The equilibrium will shift from left to right.

- If the concentration of the reactants is decreased, the concentration of the products (C or D) decreases. This favors the backward reaction. Equilibrium shifts from right to left.
- Increase in concentration of the products (C or D), increases the reactants, and thus favors the backward reaction. Equilibrium shifts from right to left.
- Decrease in concentration of products (C or D) decreases the reactants, thus favors the forward reaction. Equilibrium shifts from left to right.

2. <u>Temperature</u>

- An increase in temperature favors endothermic reaction. This is because the system counteracts the change made by absorbing the extra heat.
- A decrease in temperature favors the exothermic reaction. This is because the system counteracts the change made by producing more heat.

3. <u>Pressure</u>

- Pressure only affects gaseous system (where the reactants and products are in gaseous form).
- Gas particles in a container collide with the walls of the container, exerting a force called the pressure of the gas.
- The more the collision between the gas particles and the walls of the container, the greater the gas pressure.

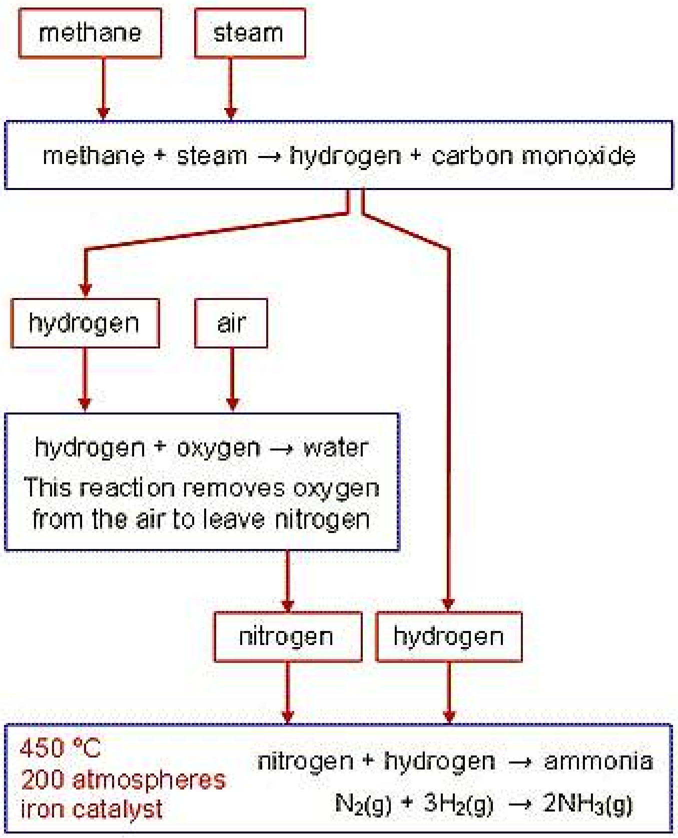
- Therefore, gas pressure can be increased by two ways:
    1. Adding more gas particles to a closed system
    2. Reducing the volume of a closed system.

- An increase in pressure favors the side which has lesser number of moles.
  This is because as the pressure increases, the system acts to oppose the change by reducing the pressure and this is done by decreasing the amount of gas particles.
- A decrease in pressure favors the side which has more number of moles.
- In reactions where the total number of moles of reactants and products is the same, pressure changes have no effect on the equilibrium.

Example: The Haber process

- Ammonia is manufactured by combining nitrogen and hydrogen in an important industrial process called the Haber process.
- The raw materials for this process are hydrogen and nitrogen.
- Nitrogen and hydrogen react together under these conditions:
  - a high temperature - about 450 ºC
  - a high pressure - about 200 atmospheres
  - an iron catalyst
- The reaction is reversible, thus some nitrogen and hydrogen remain mixed with the ammonia.
- The reaction mixture is cooled so that the ammonia liquefies and can be removed.
- The remaining nitrogen and hydrogen are recycled.
- The reaction equation is: $N_{2\,(g)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)}$ ΔH= -92 kJ/mol

The flow chart below shows the main stages in the Haber process



Source: http://www.bbc.co.uk

Effect of pressure on this reaction

$$N_{2\,(g)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)}$$

1 mole 3 moles 2 moles

An increase in pressure will favor the product side (i.e. NH_3) because it has only 2 moles of ammonia. Therefore, equilibrium shifts from left to right, forming more ammonia since the forward reaction is favored. This will reduce the amount of gas particles present in the mixture.

A decrease in pressure favors the reactants as there are altogether 4 moles on the reactant side. Therefore, equilibrium shifts from right to left, forming more H_2 and N_2, since the backward reaction is favored. This will increase the amount of gas particles present in the mixture.

Note: Adding a catalyst increases the rate of both the forward and backward reactions and this causes equilibrium to be established more rapidly. Therefore, a catalyst does not alter the equilibrium position since there is no overall change in the relative amounts of reactants and products present.

Exercise

1. What does the **Le Chatelier's principle** state?
2. The Haber Process for the manufacture of ammonia from nitrogen and hydrogen involves this reversible reaction:

$$N_{2\,(g)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)} \qquad \Delta H = -92 \text{ kJ/mol}$$

i. Describe the shift in equilibrium for the above system when:
   - More NH_3 is added.
   - N_2 is removed.
   - H_2 is added.
   - Pressure is decreased.

ii. What would be the effect on the position of equilibrium if you increased the pressure? Explain your answer using Le Chatelier's Principle

iii. In order to get the maximum possible percentage of ammonia in the equilibrium mixture, would you choose to use a high or a low temperature? Explain your answer using Le Chatelier's Principle.

3. Ethanoic acid and ethanol react reversibly to form ethyl ethanoate and water.
In a closed system, a dynamic equilibrium is set up.

Ethanoic acid + ethanol ⇌ ethyl ethanoate

i. This reaction is fairly slow, and is usually done in the presence of a small amount of concentrated sulphuric acid as a catalyst to speed it up. What effect would that have on the percentage of ethanoic acid converted into ethyl ethanoate? Explain your answer.

ii. If a lot of ethyl ethanoate was needed, why would it be a bad idea to use dilute sulphuric acid as the catalyst?

4. The container where the following reaction is occurring (at equilibrium) has its volume suddenly reduced by half. Which way will the equilibrium shift to compensate this change?

$$N_{2\,(g)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)} \qquad \Delta H = -92 \text{ kJ/mol}$$

5. The closed container holding the following reaction at equilibrium has its volume suddenly increased. Which way will the equilibrium shift to compensate for this change?

$$H_{2\,(g)} + Cl_{2\,(g)} \rightleftharpoons 2\,HCl_{\,(aq)}$$

6. What happens to the position of the equilibrium (does it shift right, left, or no change) when a catalyst is added to a system at equilibrium.

7. The conversion of natural gas to synthesis gas during the production of methanol is shown by the following equation :

$$CH_{4\,(g)} + H_2O_{\,(g)} \rightleftharpoons CO_{\,(g)} + 3H_{2\,(g)} \qquad \Delta H = +206 \text{ kJ/mol}$$

(i) What does the value ΔH = +206 kJ/mol indicate about the reaction?

(ii) What would you predict about the equilibrium amount of carbon monoxide (CO), if the system underwent an increase in pressure by decreasing the volume while the temperature remained constant?

(iii) How would the equilibrium amount of carbon monoxide alter, if the temperature was increased while the pressure remained constant?

(iv) How would the addition of a nickel catalyst affect the equilibrium amount of CO if the pressure and temperature remained constant ?

3.3.5 Equilibrium Constant, K_c

- Every reversible reaction in equilibrium can be represented by an equilibrium constant, K_c.
- For a general reaction of the form: aA + bB $\rightleftharpoons$ cC + dD ; $K_c = \frac{[C]^c\,[D]^d}{[A]^a[B]^b}$

Example

For the reaction: $N_{2\,(g)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)}$, $K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$

> **Note:** Pure solids and liquids are not included in the equilibrium constant expression. This is because they do not affect the reactant amount at equilibrium in the reaction, so they are disregarded and kept at 1. Therefore, even though reactants and products of solid and liquid states appear in the reaction equation, they are not included in the equilibrium constant expression.

Water Dissociation and K_w

- When water dissociates, it produces hydronium ions (H_3O^+) and hydroxide ions (OH^-).
- Pure water is a weak conductor of electricity because it contains very low concentrations of H^+ and OH^- ions resulting from the dissociation of water molecules.
- This results in an equilibrium system, where the dissociation equation of water is:

$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

OR

$$2H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

- The equilibrium constant expression (K_c), called $\boldsymbol{K_w}$ in the case of water is:

$$K_w = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

> **Note**: **[ ]** refers to **concentration**.

- The K_w value is very small and this indicates that most water does not dissociate.
- Thus $[H_2O]$ is **very large** compared to the $[H_3O^+]$ and $[OH^-]$, and it is effectively constant. This allows the K_w expression to be simplified to:

$$K_w = [H_3O^+]\,[OH^-]$$

> This is called the **ionic product** of water.

- The value of K_w at 25°C is = 1×10^{-14} mol^2 L^{-2}
- In a neutral solution, there is equal concentration of H_3O^+ and OH^- ions.

$$[H_3O^+] = [OH^-] = 1 \times 10^{-7} \text{ mol L}^{-1}$$

- Therefore, the product of the concentrations of H_3O^+ and OH^- has a constant value of 1×10^{-14} mol^2 L^{-2} at 25 °C. i.e. $[H_3O^+]\,[OH^-] = 1 \times 10^{-14}$ mol^2 L^{-2}

  or $(1 \times 10^{-7} \text{ mol L}^{-1}) \times (1 \times 10^{-7} \text{ mol L}^{-1}) = 1 \times 10^{-14} \text{ mol}^2 \text{ L}^{-2}$

- If the $[H_3O^+]$ and $[OH^-]$ changes in a solution, the ionic product can be used to calculate the concentration of the unknown.

Example

1. What is the $[H_3O^+]$ in a 0.1 mol L^{-1} KOH solution?

$$K_w = [H_3O^+][OH^-] = 1 \times 10^{-14}\ mol^2\ L^{-2}$$

$$[H_3O^+] \times (0.1\ mol\ L^{-1}) = 1 \times 10^{-14}\ mol^2\ L^{-2}$$

$$[H_3O^+] = \frac{1 \times 10^{-14}\ mol^2 L^{-2}}{0.1\ mol\ L^{-1}}$$

$$[H_3O^+] = \mathbf{1 \times 10^{-13}\ mol\ L^{-1}}$$

$KOH \rightarrow K^+ + OH^-$

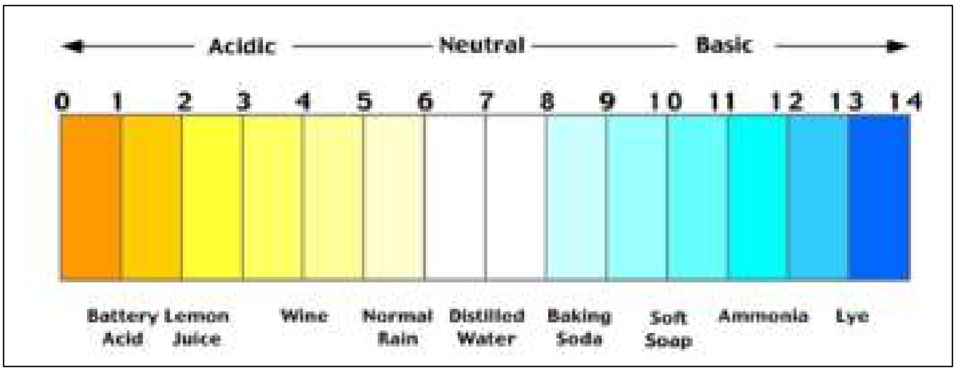
KOH : OH^-
1 : 1
0.1 mol L^{-1} : 0.1 mol L^{-1}

3.3.6 pH

- pH is a measure of how acidic or basic a substance is.
- The pH scale ranges from 0 - 14.

pH < 7 indicates an acidic solution.
pH > 7 indicates a basic solution.
pH = 7 indicates the solution is neutral.

Example



Source: https://sciencebasedpharmacy.wordpress.com

- pH can be calculated using the relationship : $pH = -\log [H_3O^+]$

Example 1

What is the pH of a 0.01 mol L^{-1} HCl solution?

Solution

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

HCl : H_3O^+
1 : 1
0.01 : 0.01

Therefore,
$pH = -\log [H_3O^+]$
$= -\log (0.01)$
$= \mathbf{2}$

Example 2

Determine the pH of a 0.5 mol L^{-1} sodium hydroxide solution.

Solution

i. Write the balanced dissociation equation.

$NaOH \rightarrow Na^{+} + OH^{-}$

ii. Compare the mole ratio.

$NaOH : OH^{-}$

1 : 1

0.5 mol L^{-1} : 0.5 mol L^{-1}

iii. Determine $[H_3O^+]$

$[H_3O^+][OH^-] = 10^{-14}$

$$[H_3O^+] = \frac{10^{-14}}{[OH^-]}$$

$$= \frac{10^{-14}}{0.5}$$

$$= 2 \times 10^{-14}$$

iv. Find pH.

$pH = -\log [H_3O^+]$

$= -\log (2 \times 10^{-14})$

= **13.6**

Exercise

1. Find the hydronium ion concentration in
   i. 0.01 mol L^{-1} NaOH solution.
   ii. 0.01 mol L^{-1} KOH solution.
   iii. 0.01 mol L^{-1} HNO_3 solution.

2. Determine the pH of
   i. 0.02 mol L^{-1} NaOH solution.
   ii. 0.01 mol L^{-1} HCl solution.
   iii. 0.1 mol L^{-1} HNO_3 solution.

3. During the course of an experiment, a student prepared a solution of sodium hydroxide by dissolving 2 g of the salt in 100 mL of water. Calculate the pH of the solution prepared.

4. 0.04 mol L^{-1} of ethanoic acid was dissolved in water to make a dilute solution of ethanoic acid.
   i. Write the dissociation equation of ethanoic acid in water.
   ii. Find the pH of the 0.04 mol L^{-1} ethanoic acid.

5. A solution of hydrochloric acid has a concentration of 3.70 g L^{-1}.
   i. Calculate the hydronium ion concentration in mol L^{-1}.
   ii. Calculate the pH of the solution.

3.3.6 Acids and Bases

According to the Bronsted-Lowry theory:

- ✓ An acid is a proton donor.
- ✓ A base is a proton acceptor.

Note: The proton here is referred to H^+.

Example

$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NH_{3\,(aq)}$

} HCl is an acid; NH_4^+ is an acid

$NH_{3\,(aq)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$

$CO_3^{2-}{}_{(aq)} + H_2O_{(l)} \rightarrow HCO_3^-{}_{(aq)} + OH^-_{(aq)}$

} NH_3 is a base; CO_3^{2-} is a base

Conjugate acids and bases

For every acid there is a conjugate base and for every base there is a conjugate acid.

Example

+ H⁺

$H_2O_{(l)} + CO_3^{2-}{}_{(aq)} \rightleftharpoons OH^-_{(aq)} + HCO_3^-{}_{(aq)}$

acid — base — Conjugate base — Conjugate acid

- H⁺

- * **The acid gives the conjugate base.**
- * **The base gives the conjugate acid.**

Note: For other examples below, the reactants and products are as follows:

Acid + Base ⇌ Conjugate base + Conjugate acid

This is the general equation for an acid-base reaction.

1. $HCl_{(aq)} + H_2O_{(aq)} \rightleftharpoons Cl^-_{(aq)} + H_3O^+_{(aq)}$
2. $H_2O_{(aq)} + NH_{3(aq)} \rightleftharpoons OH^-_{(aq)} + NH_4^+{}_{(aq)}$

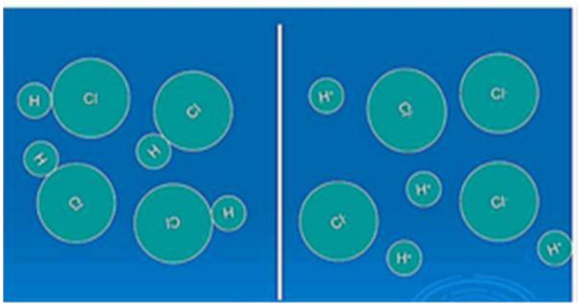
Some properties of acids and bases

Acid	Base
✓ Tastes sour. ✓ pH less than 7. ✓ Turns blue litmus red. ✓ Increases hydronium ion concentration in solution. ✓ Neutralizes a base producing a salt and water. ✓ Reacts with active metals producing hydrogen gas. ✓ React with carbonates to produce a salt, water and CO_2. ✓ Examples include: HNO_3, HCl, H_2SO_4, H_3PO_4,CH_3COOH	✓ Tastes bitter. ✓ pH greater than 7. ✓ Turns red litmus blue. ✓ Increases hydroxide ion concentration in solution. ✓ Neutralises acids producing salt and water. ✓ Feels soapy and slippery. ✓ Examples include: LiOH, NaOH, KOH, $Mg(OH)_2$, $Ca(OH)_2$, NH_4OH

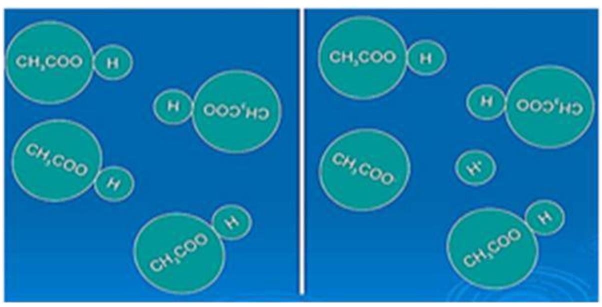
Strengths of Acids and Bases

Acids	
Strong acid	**Weak acid**
Ionises or dissociates completely in water. Has a strong tendency to donate its protons. **Example** **1. HCl** $H_2O_{(l)} + HCl_{(aq)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$ **2. HNO_3** $H_2O_{(l)} + HNO_{3(l)} \rightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$ **3. H_2SO_4** $H_2O_{(l)} + 2H_2SO_{4(l)} \rightarrow 2H_3O^+_{(aq)} + 2SO_4^{2-}{}_{(aq)}$	Only partially ionises or dissociates in water. Has low tendency to donate its protons. **Example** **1. CH_3COOH** $H_2O_{(l)} + CH_3COOH \rightleftharpoons H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$ **2. H_2CO_3** $H_2O_{(l)} + H_2CO_{3(l)} \rightleftharpoons H_3O^+_{(aq)} + HCO_3^-{}_{(aq)}$ **3. HCOOH** $H_2O_{(l)} + HCOOH \rightleftharpoons H_3O^+_{(aq)} + HCOO^-_{(aq)}$

Example of complete dissociation of HCl.



Example of partial dissociation of CH_3COOH.



Base	
Strong base	**Weak base**
- Is one which completely ionizes or dissociates in water. - It has a strong tendency to accept protons. **<u>Example</u>** 1. $NaOH_{(l)}$ $NaOH_{(l)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$ 2. $KOH_{(l)}$ $KOH_{(l)} \rightarrow K^{+}_{(aq)} + OH^{-}_{(aq)}$	- Is one which only partially ionizes or dissociates in water. - It has a weak tendency to accept protons. **<u>Example</u>** 1. $NH_4OH_{(l)}$ $NH_4OH_{(l)} \rightleftharpoons NH_4^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$ 2. $NH_{3(g)}$ $NH_3 \rightleftharpoons NH_4^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$

<u>Types of acids</u>

Monoprotic acids	**Polyprotic acids**	**Amphiprotic**
- Have only one ionisable proton or hydrogen ion.	- Have more than one ionisable protons or hydrogen ions.	- Are substances which are able to react as both an acid and a base. - Depending on the reaction type, such substance can accept proton as well as donate protons.
<u>Examples:</u> HCl, NHO_3, CH_3COOH E.g. $HCl \rightarrow H^{+} + Cl^{-}$ One H^{+}	**<u>Examples</u>**: H_2SO_4, H_2CO_3 **Note:** Acids with two ionisable protons are specifically called *diprotic acids*. E.g. $H_2SO_4 \rightarrow 2H^{+} + SO_4^{2-}$ Two H^{+}	**<u>Example:</u>** H_2O <u>Amphiprotic nature of H_2O</u> ✓ Acting as a base (accepts proton): $HCl + H_2O \rightarrow H_3O^{+} + Cl^{-}$ ✓ Acting as an acid (donates proton) $H_2O + NH_3 \rightarrow OH^{-} + NH_4^{+}$

*For your practical, see the experiment on **Strengths of acids**; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Acidic solutions

- Acidic solutions are an important part of our everyday lives.
- Examples of some acidic solutions are soft drinks, lemon juice, and stomach acid (HCl).
- Acidic solutions are commonly made through one of two ways.
  i. Dissolving a compound in solid form (such as citric acid) into water.
  ii. Bubble gases, like carbon dioxide or HCl, through water.

Reactions of acidic solution

1. Acids react with metals

- When acids are reacted with active metals, they evolve hydrogen gas (H_2).

  Note: *Some metals, like gold, silver or platinum, are very unreactive and it takes very extreme conditions to get these metals to react.*
- Metals which react very easily with acids include alkali metals, alkaline earth metals as well as zinc and aluminum.

  General equation:

Acid + Metal → Salt + Hydrogen

  Example: $2HCl_{(aq)} + Zn_{(s)} \rightarrow ZnCl_{2\,(aq)} + H_{2\,(g)}$

2. Acids react with carbonates

General equation:

Acid + Carbonate → Salt + Carbon dioxide + Water

Example: $HCl_{(aq)} + CaCO_{3(s)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

3. Acids react with metal oxides

General equation:

Acid + Metal oxide → Salt + Water

Example: $2HCl_{(aq)} + CuO_{(s)} \rightarrow CuCl_{2\,(aq)} + H_2O_{(l)}$

Alkaline solution and basic substances.

- An alkaline solution consists of base solids dissolved in water. Substances are classified as basic because of one of these:

  i. They dissolve in water forming hydroxide ions (OH^-).

  Example

  $NaOH_{(s)} + H_2O_{(l)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

  ii. They dissolve in acidic solutions.

  Example

  $HCl_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

Exercise

1. Identify the conjugate acids and bases in the following equations
   i. $HCl_{(aq)} + H_2O_{(l)} \rightleftharpoons Cl^-_{(aq)} + H_3O^+_{(aq)}$
   ii. $H_2O_{(l)} + NH_{3(aq)} \rightleftharpoons OH^-_{(aq)} + NH_4^+_{(aq)}$

2. Complete and balance the following equations.
   i. $HCl_{(aq)} + Na_2CO_{3(s)} \rightarrow$
   ii. $HCl_{(aq)} + ZnO_{(s)} \rightarrow$

3. Use suitable equations to explain whether water acts as an acid or as a base when it reacts with:
   i. Hydrochloric acid solution.
   ii. Ammonia solution.

4. What special term is given to substances such as water, which can act as an acid and a base?

5. Sulphuric acid (H_2SO_4) is an example of a polyprotic acid. Explain the term polyprotic acid and write the dissociation equation of this acid.

6. Using equations, explain why ethanoic acid is a weak acid while hydrochloric acid is a strong acid.

7. Write down equations showing the dissociation of the following substances in water.
   i. CH_3COOH
   ii. H_2SO_4
   iii. HNO_3

8. Sodium hydroxide is regarded as a basic substance. Use equations to show if you agree or disagree with the above statement.

STRAND 4

MATERIALS

Strand Outcome

Investigate the physical and chemical properties of materials and how these influence the way they are used.

Sub-strands

4.1 Inorganic Chemistry
4.2 Organic Chemistry

4.1 Inorganic Chemistry

Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Describe the properties of the oxides and chlorides of the Period 3 elements.
- ✓ Describe the laboratory and industrial preparation of chlorine gas.
- ✓ Describe and explain the properties and uses of chlorine.
- ✓ Describe the reactions of chlorine.
- ✓ Describe confirmatory tests for ionic species.
- ✓ Write complete and net ionic equations for precipitation reactions.

4.1.1 The Periodic Table

- The Periodic Table is a tabular arrangement of the chemical elements, organised on the basis of their atomic number, electron configurations, and recurring chemical properties.
- The elements can be classified as metals, non-metals and metalloids.

Oxides of Period 3 elements

- Oxides are chemical compounds with one or more oxygen atoms combined with another element.
- Examples include Na_2O, MgO, Al_2O_3, SiO_2 and SO_3.These compounds are termed as oxides because here, oxygen is in combination with only one element.
- Oxides are classified as acidic or basic based on their acid-base characteristics.
- An oxide that combines with water to give an acid is termed as an acidic oxide.
- An oxide that combines with water to give a base is known as a basic oxide.
- Some oxides can react directly with water to form an acidic or basic solution.

> **Note:**
> An **amphoteric oxide** is a substance that can chemically react as both, an acid and a base.

1. Basic Oxides

- Basic oxides are the oxides of metals.
- Generally, Group I and Group II elements form basic oxides.
- If soluble in water, they react with water to produce hydroxides (alkalis).

Examples of some basic oxides with their physical and chemical properties are:

i. Sodium oxide (Na_2O)

- White solid at 20 °C.
- Melting point is 1275 °C.
- Has ionic bonding with giant ionic structure.
- Conductor of electricity in solution and molten state, but not in solid state.
- Reacts with water to form sodium hydroxide (basic solution).

  $Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2NaOH_{(aq)}$

ii. Magnesium oxide (MgO)

- White solid at 20 °C.
- Melting point is 2825 °C.
- Ionic bonding with giant ionic structure.
- A conductor of electricity in solution and molten state but not in solid state.
- Only slightly soluble in water and reacts to forms a very small amount of magnesium hydroxide (basic solution).

  $MgO + H_2O \rightarrow Mg(OH)_2$

2. Amphoteric oxides

- Amphoteric oxides are metallic oxides, which reacts with both, an acid and a base.
- Generally, amphoteric oxides form with metalloids.
- An example of an amphoteric oxide is aluminium oxide.

Aluminium oxide (Al_2O_3)

- Also known as alumina.
- White solid at 20 °C.
- Melting point is 2072 °C.
- Has ionic bonding with giant ionic structure.
- A conductor of electricity in molten state.
- Does not react with water (insoluble in water).
- Reacts with acids and bases.

Reaction with an acid	Reaction with a base
Generally: $Al_2O_{3(s)} + 6H^+_{(aq)} \rightarrow 2Al^{3+}_{(aq)} + 3H_2O_{(l)}$ **Example:** $Al_2O_{3(s)} + 6HCl_{(l)} \rightarrow 2AlCl_{3(s)} + 3H_2O_{(l)}$	**Generally:** $Al_2O_{3(s)} + 2OH^-_{(aq)} \rightarrow 2AlO_2^-{}_{(aq)} + H_2O_{(l)}$ **Example:** $Al_2O_{3(s)} + 2NaOH_{(aq)} \rightarrow 2NaAlO_{2(aq)} + H_2O_{(l)}$

3. Acidic oxides

- Acidic oxides are the oxides of non-metals. These oxides form acids with water.
- Examples include:

i. Silicon dioxide (SiO_2)

- White solid at 20 °C.
- Melting point is 1610 °C.
- Has covalent bonding with giant molecular structure.
- A non-conductor of heat and electricity.
- Does not react with water.

ii. Phosphorous pentoxide (P_4O_{10})

- White solid at 20 °C.
- Melting point is 340 °C.
- Has covalent bonding with simple molecular structure.
- A non-conductor of heat and electricity.
- Reacts with water to form phosphoric acid.

  $P_4O_{10(s)} + 6H_2O_{(l)} \rightarrow 4H_3PO_{4(aq)}$

iii. Sulphur trioxide (SO_3)

- Gas at 20 °C.
- Melting point for SO_3 is 17 °C.
- Covalent bonding with simple molecular structure.
- They are non-conductors of heat and electricity.

  Reaction with water: SO_3 reacts to form sulphuric acid.

  $SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$

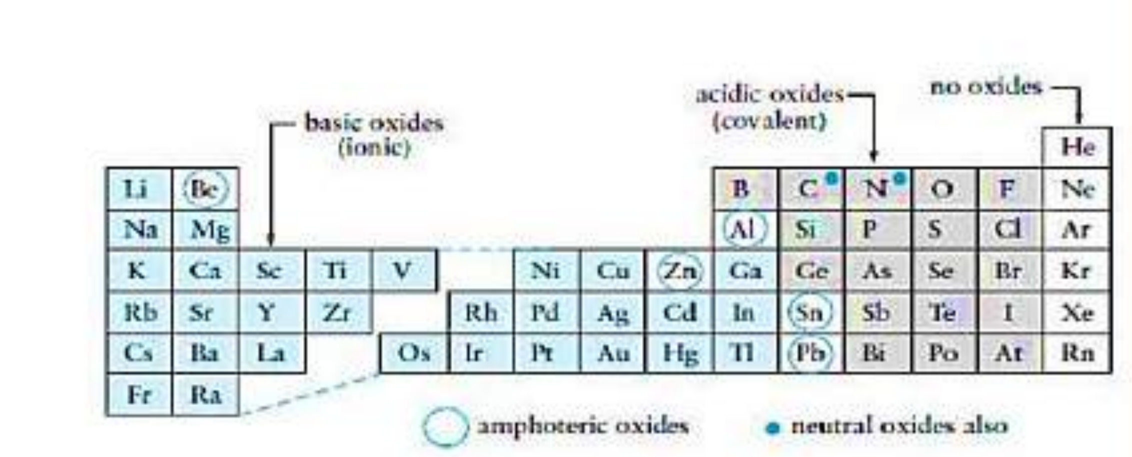
iv. Chlorine monoxide (Cl_2O)

- Gas at 20 °C.
- Melting point is -20 °C.
- Covalent bonding with simple molecular structure.
- They are non-conductors of heat and electricity.
- Reacts with water to form hypochlorous acid.

  Reaction with water:

  $Cl_2O_{(g)} + H_2O_{(l)} \rightarrow 2HClO_{(aq)}$

Summary:



Source: http://www.slideshare.net/bsvab/acidic-and-basic-oxides-16541388

General trends in the oxides of Period 3 elements across left to right of the period are:

- Bonding changes from ionic to covalent.
- Structure changes from giant ionic to macromolecular to simple (discrete).
- The melting point decreases.
- The nature of the oxides changes from basic to amphoteric to acidic.

Chlorides of Period 3 elements

i. Sodium chloride (NaCl)

- White solid at 20 °C.
- Melting point is 801 °C.
- Has ionic bonds and a giant array of sodium and chloride ions (ionic lattice).
- A conductor of electricity in solution and molten form but not in solid state.
- Dissolves readily in water to form sodium and chloride ions (neutral solution).

  $NaCl_{(aq)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$

ii. Magnesium chloride ($MgCl_2$)

- White solid at 20 °C.
- Melting point is 712 °C.
- Has ionic bonds and a giant array of magnesium and chloride ions (ionic lattice).
- A conductor of electricity in solution and molten form but not in solid state.
- Dissolves readily in water to form magnesium and chloride ions. The solution is slightly acidic.

  $MgCl_{2(aq)} \rightarrow Mg^{2+}_{(aq)} + 2Cl^{-}_{(aq)}$

iii. Aluminum chloride ($AlCl_3$)

- White solid at 20 °C.
- Melting point is 180 °C.
- Has an ionic lattice but with a lot of covalent character at room temperature.
- A very poor conductor of electricity in molten form.
- Aluminum chloride reacts dramatically with small amounts of water producing heat and hydrogen chloride fumes.

$$AlCl_{3(s)} + H_2O_{(l)} \rightarrow AlCl_2OH_{(s)} + HCl_{(g)}$$

iv. Silicon tetrachloride ($SiCl_4$)

- Colorless liquid at 20 °C.
- Melting point is 68 °C.
- Has covalent bonds and simple molecular structure.
- It does not conduct electricity.
- Reacts violently with small amounts of water producing heat and hydrogen chloride fumes.

$$SiCl_{4(l)} + 4H_2O_{(l)} \rightarrow Si(OH)_{4(s)} + 4HCl_{(g)}$$

v. Phosphorous trichloride (PCl_3)

- A colorless fuming liquid at 20 °C.
- Melting point is -91 °C.
- Has covalent bonds with simple molecular structure.
- It does not conduct electricity.
- It reacts violently with water to produce phosphorous acid (acidic solution) and fumes of hydrogen chloride.

$$PCl_{3(l)} + 3H_2O_{(l)} \rightarrow H_3PO_{3\,(aq)} + 3HCl_{(g)}$$

vi. Sulphur dichloride (SCl_2)

- Cherry red liquid at 20 °C.
- Melting point is -80 °C.
- Has covalent bonds with simple molecular structure.
- Is a non-conductor of heat and electricity.
- It reacts violently with water to form HCl (acidic solution) and sulphur dioxide.

$$SCl_{2\,(l)} + H_2O_{\,(l)} \rightarrow SO_{2\,(g)} + HCl_{\,(aq)}$$

vii. Chlorine (Cl_2)

- Greenish yellow gas at 20 °C.
- Melting point is -101 °C.
- Has covalent bonds with simple molecular structure.
- Is a non-conductor of heat and electricity.
- Reacts with water to form hypochlorous and hydrochloric acids.

$$Cl_{2(g)} + H_2O_{(l)} \rightarrow HClO_{(aq)} + HCl_{(aq)}$$

Summary

General trends in the chlorides of Period 3 elements:

1. Bonding and structure

   Ionic (lattice) → covalent (simple molecular)

2. Melting Point

   High → low (decreases across the period)

3. Nature

   Neutral → acidic

4. Conductivity

   Conductor to non-conductor

5. State

   Solid → liquid → gas

Exercise

1. Most of the oxides of non-metallic elements are
   A. ionic and basic.
   B. ionic and acidic.
   C. covalent and basic.
   D. covalent and acidic.

2. Which of the following chlorides gives a neutral solution when added to water?

   I NaCl II $AlCl_3$ III PCl_3

   A. I only. B. I and II only. C. II and III only. D. I, II and III.

3. Which of the following statement(s) is true for magnesium, phosphorus and sulphur?

A. Their chlorides are all ionic solids.
B. Their oxides all have high melting points.
C. They are in the same period of the Periodic Table.
D. They all form predominantly covalently bonded compounds.

4. Which of the following chlorides has the lowest melting point?

A. NaCl
B. PCl_3
C. $SiCl_4$
D. $MgCl_2$

5. Which of the following would react with cold water vigorously and form a strong basic solution?

A. Sodium
B. Sulphur
C. Chlorine
D. Magnesium

6. The solution that would not change the colour of litmus paper is

A. NaCl.
B. NaOH.
C. HCl.
D. NH_4OH.

7. The compounds Na_2O, Al_2O_3 and SO_2, respectively, are

A. acidic, amphoteric and basic.
B. amphoteric, basic and acidic.
C. basic, acidic and amphoteric.
D. basic, amphoteric and acidic.

8. When sodium oxide and sulphur dioxide are added to separate test tubes containing water the solution will be __________ and ____________, respectively.

A. acidic, acidic
B. acidic, basic
C. basic, acidic
D. basic, basic

9. Complete the following equations:

i. $Na_2O_{(s)} + H_2O_{(l)} \rightarrow$ ----------
ii. $Al_2O_{3(s)}$ + ------- $\rightarrow 2AlCl_{3(s)} + 3H_2O_{(l)}$
iii. $P_4O_{10(s)} + 6H_2O_{(l)} \rightarrow$ ---------
iv. $SO_{3\,(g)}$ + -------- $\rightarrow H_2SO_{4(aq)}$
v. $SiCl_{4(l)}$ + -------- $\rightarrow Si(OH)_{4(s)} + 4HCl_{(g)}$
vi. $PCl_{3(l)} + 3H_2O_{(l)} \rightarrow H_3PO_{3(aq)}$ + ---------
vii. $Cl_{2(g)} + H_2O_{(l)} \rightarrow$ ------- $+ HCl_{(aq)}$

10. From the oxides given below, answer the questions that follow.
Na_2O MgO Al_2O_3 P_4O_{10} SO_2

   i. Identify the oxide that would dissolve in water to give the most basic solution.
   ii. Identify the oxide that is solid at room temperature and would react with both sodium hydroxide and hydrochloric acid.
   iii. Identify an oxide that exists as gas at room temperature.

11.. Justify the large difference in the melting points of MgO (1275 °C) and SO_3 (17 °C).

12. What are amphoteric oxides? Give an example with reaction equations to support your answer.

13. Describe the trend in bond type and the nature of the oxides of Period 3 elements.

14. Describe the trend in bond type and the nature of the chlorides of Period 3 elements.

15. State the trend in the melting points of Period 3 oxides. Give explanations for the suggested trend.

16. State the trend in the melting points of Period 3 chlorides. Give explanations for the suggested trend.

17. Complete the following table by filling in for (i) – (viii).

Oxide	State at 25°C	Bonding	Structure
Na_2O	(i)	(ii)	(iii)
SiO_2	solid	(iv)	(v)
P_4O_{10}	solid	covalent	simple molecular where molecules are held together by weak intermolecular forces
SO_3	(vi)	(vii)	(viii)

18. Complete the following table by filling in for (i) – (viii).

Chloride	State at 25°C	Bonding	Structure
NaCl	(i)	(ii)	ionic lattice
$SiCl_4$	solid	covalent	(iii)
PCl_3	(iv)	(v)	(vi)
SCl_2	(vii)	(viii)	simple molecular where molecules are held together by weak intermolecular forces

19. The properties of two compounds of an element Q from Period 3 are given in the table below.

	State at 25°C	Structure	Effect of adding water
Oxide of Q	Solid	Ionic lattice	Dissolves in water; forms alkaline solution.
Chloride of Q	Solid	Ionic lattice	Dissolves in water; forms a neutral solution.

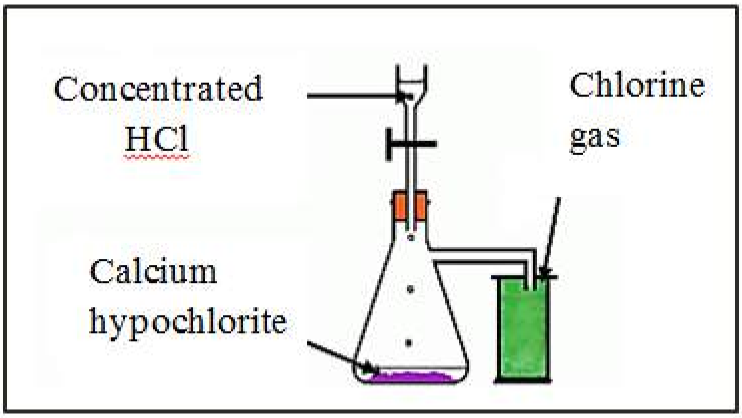
   i. What is the identity of element Q?
   ii. Justify your answer in i above with suitable explanation and equations.

4.1.2 Chlorine Gas

- Chlorine is a chemical element with symbol Cl and atomic number 17.
- It is in the halogen group (Group VII) and is the second lightest halogen following fluorine.
- Exists as diatomic molecules (Cl_2).
- A yellow-green gas at 20 ^{o}C.
- Bleaches moist blue litmus paper.
- It is poisonous and has a suffocating or choking or irritating smell.

Laboratory Preparation of Chlorine gas

Chlorine can be prepared in the laboratory by reacting acids and calcium hypochlorite (bleaching powder).



Source: http://watertreatmentchemical.org

Method of Collection:

Chlorine gas is collected by upward displacement of air because it is denser than air.

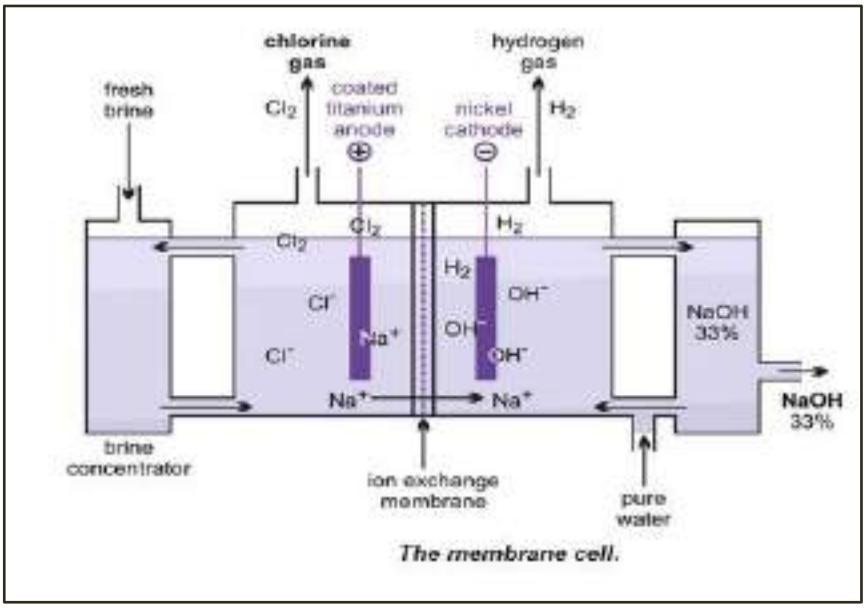
Note: HNO_3 can also be used in place of concentrated HCl.

Reaction equation: $Ca(OCl)_{2(s)} + 4HCl_{(l)} \rightarrow CaCl_{2(s)} + 2H_2O_{(l)} + 2Cl_{2(g)}$

*For your practical, see the experiment on **Chlorine demonstration**; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Industrial Preparation of Chlorine gas

- Chlorine gas can be prepared in the industries by the electrolysis of sodium chloride solution (brine).



Source: http://www.essentialchemicalindustry.org

The reactions occurring during electrolysis are:

Anode (Oxidation)	Cathode (Reduction)
Chlorine is oxidised to chlorine gas. $2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$	Water is reduced to hydrogen gas. $2H_2O_{(l)} + 2e^- \rightarrow 2OH^-_{(aq)} + H_{2(g)}$
The OH^- from water reacts with Na^+ in the electrolyte to form NaOH $Na^+_{(aq)} + OH^-_{(aq)} \rightarrow NaOH_{(aq)}$ Thus the overall equation for the reaction is: $2NaCl_{(aq)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + Cl_{2(g)} + H_{2(g)}$	

<u>Reactions of chlorine gas (Cl_2)</u>

1. Reaction with iron (Fe)

- When chlorine gas is passed over hot iron in a combustion tank, the iron burns to form iron(III) chloride.

- Iron(III) chloride forms black crystals.
  $2Fe_{(s)} + 3Cl_{2(g)} \rightarrow 2FeCl_{3(s)}$

2. **Reaction with calcium hydroxide solution**

- Calcium hydroxide reacts with chlorine gas to produce the bleaching agent, calcium hypochlorite.

$$Ca(OH)_{2(s)} + Cl_{2(g)} \rightarrow Ca(OCl)_{2(aq)} + H_2O_{(l)}$$

3. **Reaction with sodium hydroxide solution**

- Chlorine reacts with warm concentrated NaOH solution to give sodium chloride and sodium hypochlorate ($NaClO_3$).

- The reaction between chlorine and warm concentrated sodium hydroxide solution is:

$$6NaOH_{(aq)} + 3Cl_{2(g)} \rightarrow 5NaCl_{(aq)} + NaClO_{3(aq)} + 3H_2O_{(l)}$$

- The reaction between chlorine and cold dilute sodium hydroxide solution produces sodium hypochlorite (NaClO).

$$2NaOH_{(aq)} + Cl_{2(g)} \rightarrow NaCl_{(aq)} + NaClO_{(aq)} + H_2O_{(l)}$$

4. **Reaction with water**

- Chlorine is only slightly soluble in water. It usually forms a mixture of two acids; hypochlorous acid (HOCl) and hydrochloric acid (HCl).

$$Cl_{2(g)} + H_2O_{(l)} \rightarrow HOCl_{(aq)} + HCl_{(l)}$$

5. **Reaction with moist litmus paper (Test for chlorine)**

- Chlorine gas (Cl_2) turns moist blue litmus paper red and then bleaches it white.
- The litmus paper turns red because of formation of HCl and it turns white because of HClO.

$$Cl_{2(g)} + H_2O_{(l)} \rightarrow HOCl_{(aq)} + HCl_{(l)}$$

6. **Reaction with damp starch iodide paper (Test for chlorine)**

- Chlorine makes damp starch-iodide paper turn bluish-black.
- This is because the chlorine releases iodine from the potassium iodide and then iodine reacts with starch to produce a bluish-black color.

$$2KI_{(aq)} + Cl_{2(g)} \rightarrow 2KCl_{(aq)} + I_{2(g)}$$

$$2I^-_{(aq)} + Cl_{2(g)} \rightarrow 2Cl^-_{(aq)} + I_{2(g)}$$

<u>Uses of chlorinating agent (sodium hypochlorite)</u>

1. **Bleaching**

- Sodium hypochlorite (NaClO) is the main ingredient in laundry bleach.
- It is used extensively as a bleaching agent in the textile, detergents, and paper and pulp industries.

2. Anti-bacterial

- Large quantities of sodium hypochlorite are also used as a disinfectant in water and waste water treatment and sanitary equipment.
- For example, in food processing, sodium hypochlorite is used to sanitize food preparation equipment. It is also used in fruits and vegetable processing, mushroom production, hog and poultry production, maple syrup production, and fish processing.

3. Oxidising Properties

- Sodium hypochlorite is also used as an oxidizing agent for organic products.

Example: Paper and Pulp Industry

The paper and pulp industry converts wood or recycled fibre into pulp and primary forms of paper.

- The uses of chlorine and NaClO in paper and pulp industry are as follows:

Step 1: Chlorination

Chlorination of the unbleached pulp is the first stage that converts the lignins into soluble chlorolignins.
It also oxidises other coloured materials found in the wood pulp. The purpose of the bleaching process is to whiten the pulp.

$$\text{Lignin} + \text{Chlorine} \rightarrow \text{Chlorolignin} + \text{HCl}$$

Step 2 Caustic soda extraction stage (alkaline extraction)

The soluble chlorolignin formed from Step 1 is removed by treating it with sodium hydroxide(NaOH). After alkaline treatment, the pulp is washed to remove the soluble material.

Step 3 Hypochlorite stage

Lignin and other colored compounds remaining after chlorination and alkali extraction are removed by the oxidation process. Sodium hypochlorite (NaClO) is used as an oxidant and bleaching agent in this step.

Other Uses of Chlorine (Cl_2)

- Chlorine is used to make consumer products such as paper, paints and other textiles and insecticides.
- About 20% of chlorine produced is used to make PVC.
- Another major use for chlorine is in organic chemistry. It is used as an oxidising agent.

Exercise

1. Chlorine can be prepared in the laboratory using concentrated HCl and calcium hypochlorite using upward displacement of air.

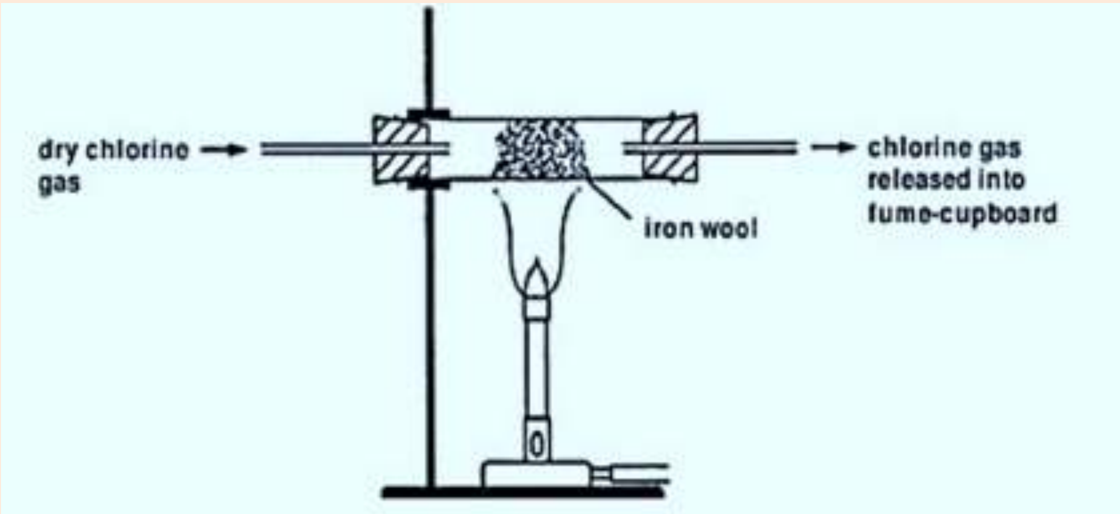
   i. Write the balanced chemical equation for the preparation of chlorine gas as described above.
   ii. Why is it possible to collect the gas by upward displacement of air?
   iii. After preparing chlorine gas in the laboratory, a student wished to test for the gas prepared. He placed damp blue litmus paper at the mouth of the test-tube containing chlorine gas. Briefly explain what the student would observe and why.
   iv. State and describe another method the student can use to test for the chlorine gas prepared.

2. Chlorine can also be prepared in the industries by the electrolysis of sodium chloride solution. Write the balanced chemical equation which occurs at the anode and cathode during electrolysis of sodium chloride solution.

3. Chlorine and chlorinating agents play an important role in the paper and pulp industry.

   i. Write a balanced equation for the reaction of chlorine and sodium hydroxide to produce sodium hypochlorite.
   ii. What purpose does chlorine and sodium hypochlorite serve in the paper and pulp industry?

4. Study the diagram given below and answer the question that follows.



Source: http://www.nuffieldfoundation.org

Write the balanced chemical equation showing iron wool reacting with chlorine gas.

4.1.3 Precipitation Reactions

- Precipitation reactions occur when cations and anions in aqueous solution combine to form an insoluble ionic solid called a precipitate (an insoluble solid that forms from mixing two solutions).
- A common example is the mixing of clear solutions of silver nitrate ($AgNO_3$) and sodium chloride (NaCl).
  The reaction is: $AgNO_{3\,(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3\,(aq)}$
  White Precipitate
- The solubility rules for common ionic solids can be used to determine whether or not such a reaction occurs.
- Since not all aqueous reactions form precipitates, one must consult the solubility rules before determining the state of the products and writing a *net ionic equation*.
- The ability to predict these reactions allow scientists to determine which ions are present in a solution, and also allow industries to form chemicals by extracting components from these reactions.

Solubility Rules

Substance	**Soluble**	**Insoluble**
Nitrates	All	None
Sulphates	Most	Lead sulphate, barium sulphate, calcium sulphate (slightly)
Chlorides	Most	Silver chloride and lead chloride
Carbonates	Sodium carbonate, potassium carbonate and ammonium carbonate	Most other carbonates
Hydroxides	Sodium hydroxide. potassium hydroxide and ammonium carbonate	Most other hydroxides

Note:

- Nitrates and most chlorides are soluble. This is why many of the chemicals used in the laboratory are nitrates or chlorides.
- In order to make an insoluble salt, we can react together two soluble salts in a precipitation reaction.

1. **Carbonate ion (CO_3^{2-})**

- Add any dilute strong acid to the substance suspected to be a solid carbonate. If a fizzing, colourless gas is given off, test with limewater.

$$CO_3^{2-}{}_{(aq)} + 2H^+{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$$

- If the limewater turns milky and cloudy fine white precipitates are formed, this indicates that the substance is a carbonate.

$$CO_{2(g)} + Ca(OH)_{2(aq)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$$

Lime water

2. **Chloride ions (Cl^-)**

- To a solution of the salt suspected to have chloride ions, add silver nitrate solution. A white precipitate will form which will be insoluble in dilute mineral acids and when exposed to sunlight it will turn dark (violet) indicating that the salt is silver chloride.

$$Ag^+{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow AgCl_{(aq)}$$

- Add aqueous ammonia solution to the precipitate to confirm the test.
- The precipitate should dissolve indicating the presence of silver chloride.

3. **Sulphate ions (SO_4^{2-})**

- Add barium chloride to the solution of the suspected salt. A white precipitate will form.

$$Ba^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$$

- The precipitate will be insoluble in all acids, no color change will occur when exposed to sunlight and it will not dissolve in aqueous ammonia solution.

Confirmatory tests for the presence of cations

1. **Iron(II) ion (Fe^{2+})**

❖ Upon addition of sodium hydroxide solution to the sample suspected to have Fe^{2+}, a green precipitate of $Fe(OH)_2$ forms.

$$Fe^{2+}{}_{(aq)} + 2OH^-{}_{(aq)} \rightarrow Fe(OH)_{2(s)}$$

2. **Iron(III) ion (Fe^{3+})**

i. Upon addition of sodium hydroxide solution to the sample suspected to have Fe^{3+}, an orange precipitate of Fe $(OH)_3$ forms.

$$Fe^{3+}_{(aq)} + 3OH^-_{(aq)} \rightarrow Fe(OH)_{3(s)}$$

ii. Another test for Fe^{3+} is to add few drops of potassium thiocyanate (KSCN) or ammonium thiocyanate (NH_4SCN) solution to the sample suspected to have Fe^{3+}. Formation of a dark red solution indicates presence of Fe^{3+}.

3. **Copper ions (Cu^{2+})**

❖ Upon addition of sodium hydroxide to the sample suspected to have Cu^{2+}, a blue precipitate of $Cu(OH)_2$ will form.

$$Cu^{2+}_{(aq)} + OH_{(aq)} \rightarrow Cu(OH)_{2(s)}$$

❖ When concentrated ammonia solution is added to a new sample, a powdery light blue precipitate of copper(II) hydroxide forms which dissolves to form deep blue solution upon addition of excess ammonia solution.

4. **Silver ions (Ag^+)**

❖ Upon addition of sodium hydroxide to the sample suspected to have Ag^+, a brown precipitate of Ag_2O forms.

$$2Ag^+_{(aq)} + 2OH^-_{(aq)} \rightarrow Ag_2O_{(s)} + H_2O_{(l)}$$

❖ When ammonia solution is added to a new sample, a brown precipitate forms which dissolves in excess ammonia solution.

5. **Magnesium ions (Mg^{2+})**

❖ Upon addition of sodium hydroxide to the sample suspected to have Mg^{2+}, a white precipitate of Mg $(OH)_2$ forms which is insoluble in excess of NaOH.

$$Mg^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Mg(OH)_{2(s)}$$

6. **Aluminium ions (Al^{3+})**

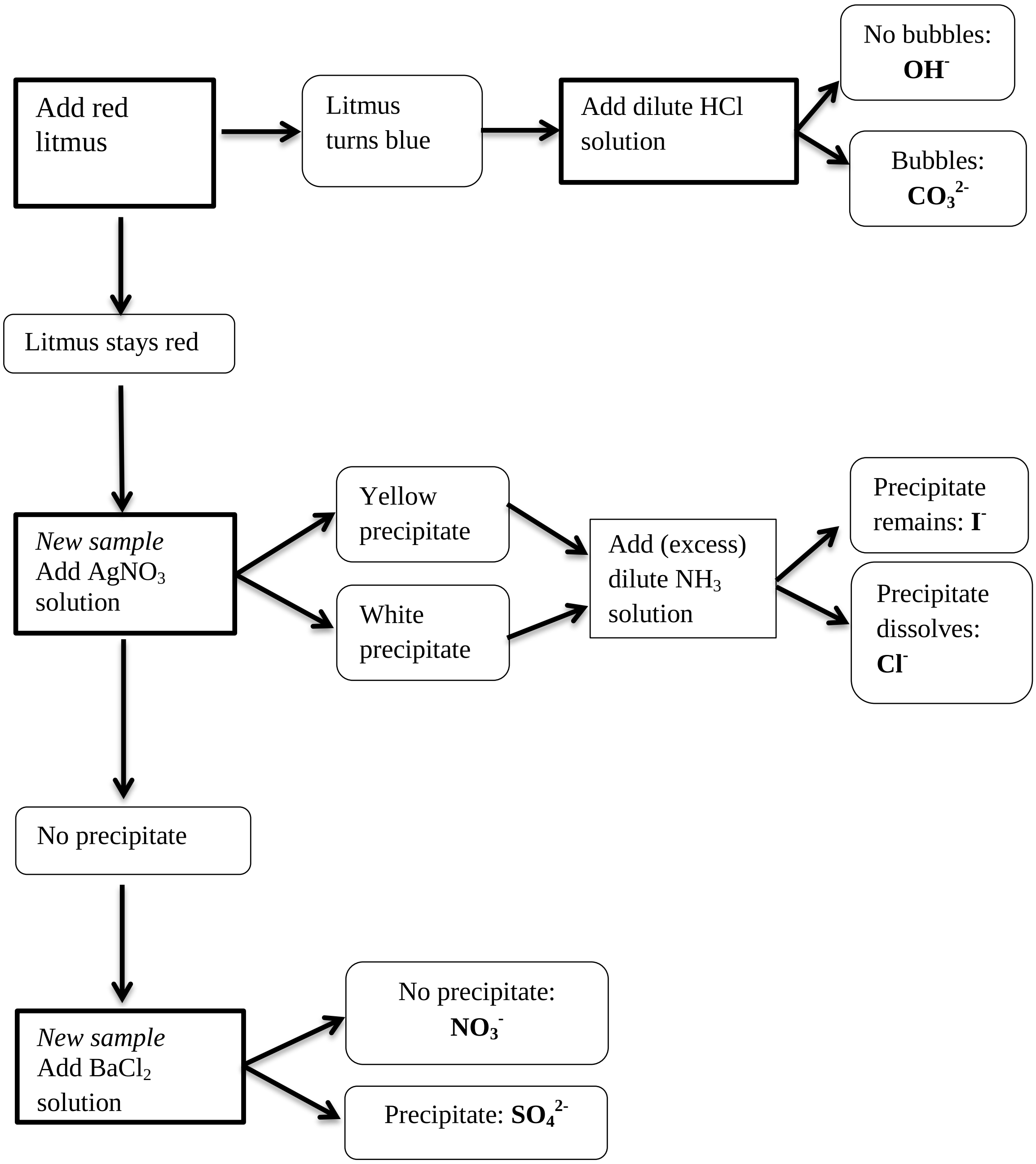
❖ Upon addition of sodium hydroxide to the sample suspected to have Al^{3+}, a white precipitate of Al $(OH)_3$ forms which dissolves in excess of NaOH.

$$Al^{3+}_{(aq)} + 3OH^-_{(aq)} \rightarrow Al(OH)_{3(s)}$$

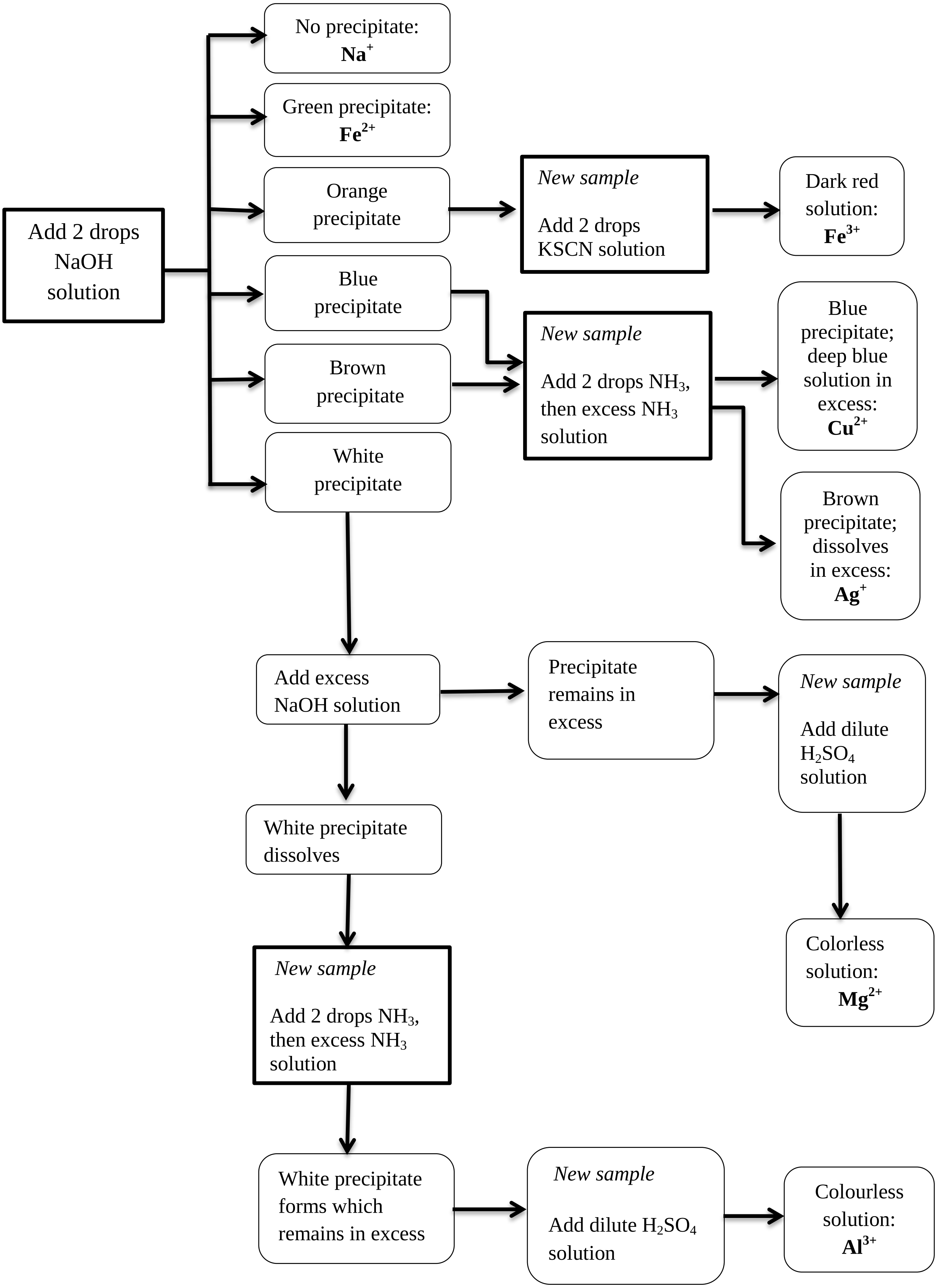
❖ When ammonia solution is added to a new sample, a white precipitate forms which does not dissolve even in excess ammonia solution.

*For your practical, see the experiment on **Test for ions**; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Summary: Testing for anions flow chart: Cl^-, CO_3^{2-}, I^-, NO^{3-}, OH^-, SO_4^{2-}



Summary: Testing for cations flow chart: Ag^+, Al^{3+}, Cu^{2+}, Fe^{2+}, Fe^{3+}, Mg^{2+}



Exercise

1. Classify the following compounds into two groups: soluble and insoluble.

   Lead chloride, silver nitrate, calcium carbonate, calcium hydroxide, sodium hydroxide, barium sulphate, silver iodide, ammonium nitrate, potassium chloride, lead sulphate, sodium carbonate and magnesium chloride.

2. Copy and complete the table below which summarizes the observations of the reactions that would occur upon mixing two pairs of solutions.

Reaction solutions	Sodium carbonate	Sodium chloride	Sodium sulphate
Silver nitrate	**(a)**	**(b)**	A white precipitate of silver sulphate
Barium chloride	A white precipitate of barium carbonate	No reaction	**(c)**
Copper chloride	**(d)**	**(e)**	

3. Copy and complete the table below which summarizes the observations of the reactions that would occur upon mixing two pairs of solutions.

Reaction solutions	Solution containing $Cu^{2+}_{(aq)}$	Solution containing $Al^{3+}_{(aq)}$	Solution containing $Ag^{+}_{(aq)}$
Sodium hydroxide	**(a)**	A white precipitate forms which dissolves in excess of NaOH.	**(c)**
Ammonia	A powdery light blue precipitate forms which dissolves to form deep blue solution upon addition of excess ammonia solution.	**(b)**	**(d)**

4. Describe a simple laboratory test to distinguish between the Fe^{2+} and Fe^{3+} ions.

4.2 Organic Chemistry

Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Use IUPAC rules to name hydrocarbons.
- ✓ Name hydrocarbons using IUPAC rules.
- ✓ Write and describe the different formula representations of hydrocarbons.
- ✓ Describe the production of ethene and ethyne with equations.
- ✓ Describe the reactions of hydrocarbons with balanced chemical equations.
- ✓ Describe isomerism in hydrocarbons.
- ✓ Describe the properties, preparation and reactions of alcohols carboxylic acid.
- ✓ Describe and explain the chemistry of the breathalyzer test.
- ✓ Describe the properties and preparation of esters.
- ✓ Describe the production of some organic substances with relevant equations.

4.2.1 Alkanes, Alkenes and Alkynes

Alkanes, alkenes and alkynes are hydrocarbons. Hydrocarbons are organic compounds that contain only carbon and hydrogen atoms joined together by covalent bonds.

Functional group: A functional group is an atom or group of atoms within a molecule that shows a characteristic set of physical and chemical behaviors. Functional groups are important in organic chemistry because:

1. They take part in chemical reactions.
2. They determine the physical properties of a molecule.
3. They help in classifying and naming organic compounds.

Homologous series: is a series of compounds with the same general formula, usually varying by a single parameter such as the length of a carbon chain.

Hydrocarbons are divided into two groups: Saturated and unsaturated hydrocarbons.

Saturated hydrocarbon	Unsaturated hydrocarbon
✓ Hydrocarbon that contains only carbon-carbon single bonds (*There are no carbon-carbon double bonds or triple bonds*). ✓ Alkanes are saturated hydrocarbons.	✓ A hydrocarbon that contains one or more carbon-carbon double bonds and triple bonds. ✓ Alkenes and alkynes are unsaturated hydrocarbons.

Alkanes

- The molecular formula of this group of hydrocarbons is C_nH_{2n+2}.
  (n is the number of carbon atoms, 2n + 2 is the number of hydrogen atoms).
- Alkanes form homologous series because each successive member differs from the next by a CH_2 group.

Nomenclature of Alkanes

IUPAC Rules for Alkane Nomenclature

Chemists have adopted a set of rules established by the International Union of Pure and Applied Chemistry (IUPAC) for naming hydrocarbons.

The general steps for naming alkanes are:

1. Find and name the longest continuous carbon chain.

   Note: The chain does NOT have to be only straight. This means it can be from any direction.

2. Identify and name the substituents attached to this chain.
3. Number the chain consecutively, starting at the end nearest a substituent group.
4. Designate the location of each substituent group by an appropriate number and name.
5. Assemble the name, listing groups in alphabetical order using the full name.

   Note: Always use a dash between a number and a word and always use a comma between numbers.
6. If the alkyl group is present more than once, prefixes are used in front of the alkyl group.

   (Example: The prefixes di for 2, tri for 3 and tetra for 4 are used to designate several groups of the same kind. The prefixes are ***not*** *considered when alphabetizing).*
7. If two possible longest chains of equal length exist, the chain with the greatest number of substituents is selected as the parent chain.
8. When more than one substituents are present at an equal distance from either end of the parent chain, the parent chain should be numbered in such a way that gives the lowest position to the substituents.
9. When two or more substituents are present at an equal position from either side of the longest chain, alphabetical order of the substituent's names must be used in assigning a lower number.

The parent name of the longest continuous carbon chain is as follows:

Number of carbon atoms in the longest chain	Parent name
1	Methane
2	Ethane
3	Propane
4	Butane
5	Pentane
6	Hexane

Substituent groups

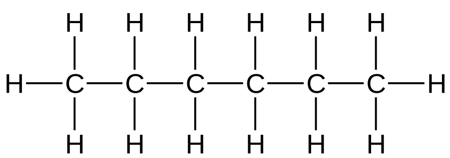
Substituents groups are the side branches in organic compounds. A substituent group derived from an alkane by removal of a hydrogen atom is called an **alkyl group (R-)**. Alkyl groups are named by dropping the "-ane" from the name of the parent alkane and adding the suffix "**-yl**".

Example

Number of carbon atoms in the substituent	Name of substituent
1 (CH_3)	methyl
2 (C_2H_5)	ethyl
3 (C_3H_7)	propyl
4 (C_4H_9)	butyl
Some substituents derived from other elements or other species than alkanes	
F	Fluoro
Cl	Chloro
Br	Bromo

Example 1

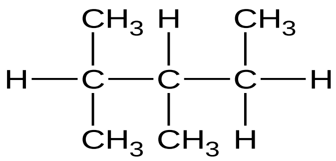
Name the organic compound shown below.



Solution

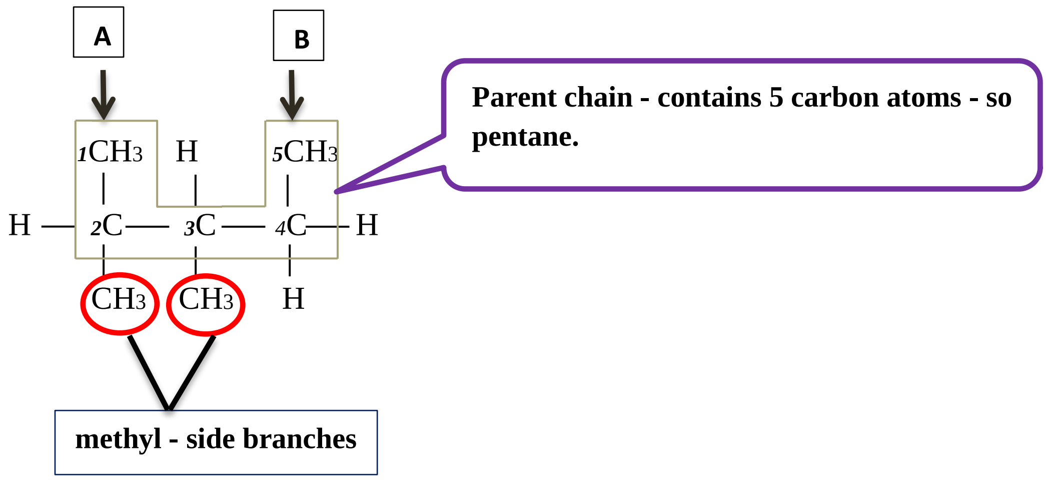
This is an unbranched alkane. The longest chain has 6 carbon atoms. Thus the name is **hexane**.

Example 2
Name the organic compound shown below



Solution

The compound given above is a branched alkane because it has substituents (side branches) sticking out of the parent chain (longest chain).



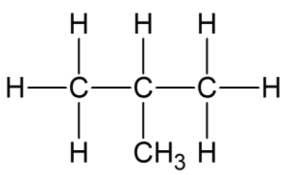
The alkane has **2 side branches**, both are **methy**l groups. The first one is located on the **2nd** carbon atom and the second one is on the **3rd** carbon atom. Thus, location and name of the substituents is: 2,3-dimethyl

(While counting from side A, the first substituent falls on Carbon 2, while counting from side B, the first substituent falls on Carbon 3, that is why we are counting from side A).

Therefore, the name of the above alkane is **2, 3-dimethlypentane**.

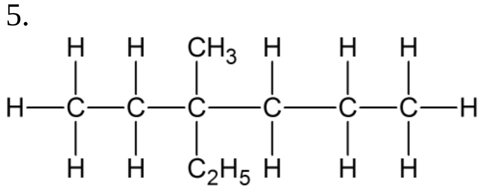
Other examples

3.



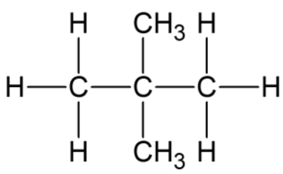
2-methylpropane

5.



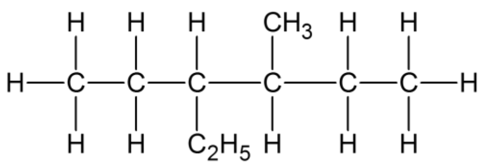
3-ethyl-3-methylhexane

4.



2,2-dimethylpropane

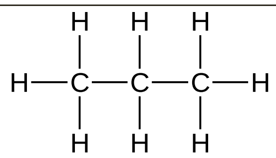
6.



3-ethyl-4-methylhexane

Representation of organic compounds

- The formula of an organic compound can be represented by the molecular formula or by the structural formula.
- Structural formulas can be represented in three ways: **Expanded (complete) structural formula**, **Condensed structural formula**, and **Line-angle formula**.

Structural formula	Description	Propane as an example
Expanded (Complete) Structural formula	- The carbon atoms are shown attached to all hydrogen atoms. - In other words, all connections are shown.	
Condensed structural formula	- The hydrogen atoms are grouped with their carbon atom. - The number of hydrogen atoms is written as subscript. *(Hydrogen atoms are shown right next to the carbon atoms to which they are attached).*	$CH_3CH_2CH_3$ or CH_3-CH_2-CH_3
Line-angle formula.	- Carbon atoms are implied at the corners and ends of lines, and each carbon atom is understood to be attached to enough hydrogen atoms to give each carbon atom four bonds.	 or 

<u>**Physical properties of alkanes**</u>

- They are nonpolar compounds.
- They are generally insoluble in water (because water is polar) and they are soluble in nonpolar solvents.
- They have a lower density than water.
- Nearly all alkanes have densities less than 1.0 g mL^{-1} and are therefore less dense than water.
- They can be gases (with 1 to 4 carbon atoms), liquids (with 5 to 17 carbon atoms), or solids (with 18 or more carbon atoms).
- Melting and boiling points of alkane's increases with increasing number of carbon atoms.

Chemical properties of alkanes

- In general, alkanes have low reactivity.
- Their most important reactions are combustion (reaction with oxygen) and substitution reactions.

1. Combustion Reactions

- Alkanes react with oxygen to produce carbon dioxide (CO_2), water (H_2O), and heat energy depending on whether the reaction is complete or incomplete.

Complete combustion	Incomplete combustion
▪ Occurs in excess oxygen and gives blue non smoky flame. **Example:** $CH_{4\,(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$	▪ Occurs in limited amount of oxygen and produces carbon (sooty flame) or carbon monoxide, depending on the amount of oxygen available. **Example:** $2CH_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{(g)} + 4H_2O_{(l)}$ or $CH_{4(g)} + O_{2(g)} \rightarrow C_{(s)} + 2H_2O_{(l)}$

2. Substitution Reactions

- Alkanes are saturated hydrocarbons therefore they undergo substitution reactions.
- In this reaction, a single hydrogen atom is successively replaced by a halogen such as chlorine or bromine.
- These reactions are catalyzed in the presence of sunlight or UV light.

Example

The reaction of methane with chlorine (Cl_2):

$$CH_4 + Cl_2 \xrightarrow{\text{light}} \underset{\text{Chloromethane}}{CH_3Cl} + HCl$$

If chloromethane is allowed to react with more chlorine:

$$CH_3Cl + Cl_2 \xrightarrow{\text{light}} \underset{\text{Dichloromethane}}{CH_2Cl_2} + HCl$$

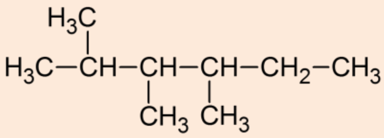
$$CH_2Cl_2 + Cl_2 \xrightarrow{\text{light}} \underset{\text{Trichloromethane}}{CHCl_3} + HCl$$

$$CHCl_3 + Cl_2 \xrightarrow{\text{light}} \underset{\text{Tetrachloromethane}}{CCl_4} + HCl$$

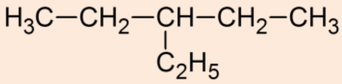
Exercise

1. Name the alkanes shown below.

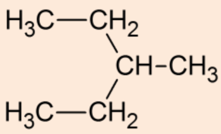
A.



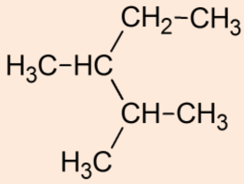
B.



C.



D.



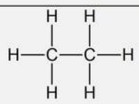
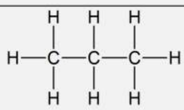
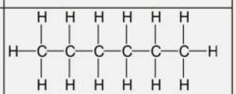
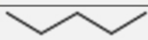
2. Draw the expanded structure of the following alkanes.

i. 2-methlybutane ii. 3-methlyhexane iii. 2,3-dimethylhexane

3. Name the following alkanes:

i. $CH_3CH(CH_3)CH_3$ ii. $CH_3CH(CH_3)CH_2(CH_3)$ iii. $(CH_3)_3CCH_2(CH_3)$

3. Complete the table below.

Alkanes						
General Formula	C_nH_{2n+2}					
Alkane	**Methane**	**Ethane**	**Propane**	**Butane**	**Pentane**	**Hexane**
Molecular Formula	CH_4	C_2H_6	C_3H_8	C_4H_{10}	C_5H_{12}	C_6H_{14}
Expanded structural formula						
Condensed structural formula	CH_4	CH_3CH_3 or CH_3-CH_3	$CH_3CH_2CH_3$ or CH_3-CH_2-CH_3	$CH_3CH_2CH_2CH_3$ or $CH_3(CH_2)_2CH_3$ or CH_3-CH_2-CH_2-CH_3		
Line-angle formula (or Skeletal formula)			 or 		 or 	

Alkenes

- Alkenes are unsaturated hydrocarbons that contain one or more carbon-carbon double bonds.
- General formula is C_nH_{2n}.

IUPAC Rules for Alkene Nomenclature

1. The general rules for naming alkanes apply to alkenes also, with some exceptions.
2. The **"ene"** suffix indicates an alkene. Wherever necessary, a number is written in front of "ene" to indicate the position of the carbon to carbon double bond.
3. The longest chain chosen for the parent name must include both carbon atoms of the double bond.
4. The parent chain must be numbered from the end nearest a double bond carbon atom.
5. If the double bond is in the center of the chain, the nearest substituent rule is used to determine the end where numbering starts. The smaller of the two numbers designating the carbon atoms of the double bond is used as the double bond locator.

The parent name of the longest continuous carbon chain is as follows:

Number of carbon atoms in the longest chain	Parent name
2	Ethene
3	Propene
4	Butene
5	Pentene
6	Hexene

Example 1

Name the alkene shown below.

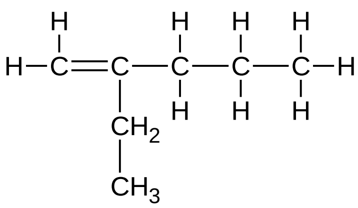
$$H_2C{=}CH_2$$

Solution

The longest chain contains 2 carbon atoms with no substituents so the name is **ethene.**

Example 2

Name the organic compound shown below.



Solution

- The compound is an alkene since it contains a carbon-carbon double bond.
- The longest chain with the double bond has 5 carbon atoms; therefore the parent name is pentene.
- The double bond falls on the 1^{st} carbon atom. Thus the parent name will be pent-1-ene.
- The substituent has 2 carbon atoms, thus will be named ethyl.
- The substituent falls on carbon number 2 (*counting from the side which gives the carbon with a double bond a lower number).*
- Therefore, the name of the above compound is **2-ethylpent-1-ene**.

Preparation of Ethene

Ethene gas can be produced in two ways:

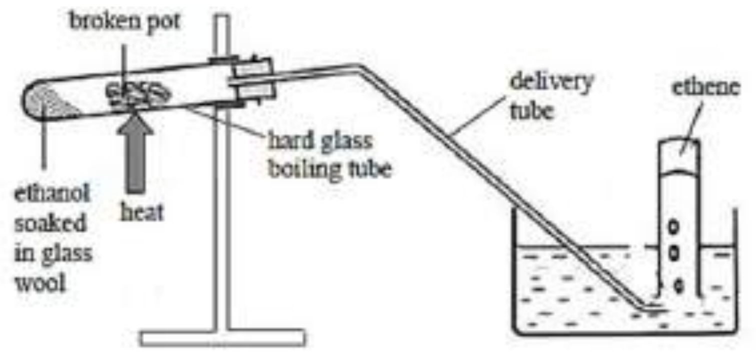
1. **Steam cracking of alkanes-** Steam cracking is a process in which saturated hydrocarbons (long chain alkane molecules) are broken down into smaller, often unsaturated, hydrocarbons.

   Example:

$$C_{11}H_{24} \xrightarrow[900\ ^oC]{steam} C_9H_{20} + C_2H_4$$

2. **Dehydration of ethanol** – Dehydration is the process of removing water molecules from a substance. Ethanol can be dehydrated to produce ethene and water when heated with concentrated sulfuric acid (H_2SO_4).

   Example



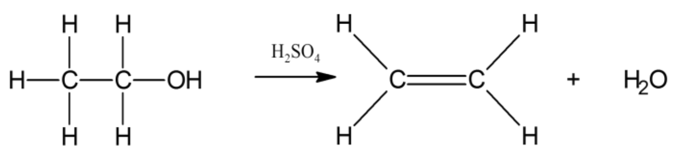
Dehydration of Ethanol

Source: http://www.nuffieldfoundation.org

Reaction equation

$$CH_3CH_2OH \xrightarrow{H_2SO_4} C_2H_4 + H_2O$$

OR



Ethene

Alkynes

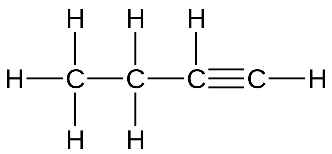
- Alkynes are unsaturated hydrocarbons that contain one or more carbon-carbon triple bond.
- The general formula of this group of organic compounds is C_nH_{2n-2}.

IUPAC Rules for Alkyne Nomenclature

1. The rules for naming alkynes are the same as that for naming alkenes, with some exceptions.
2. The "**yne**" suffix indicates an alkyne. Wherever necessary, a number is written in front of "yne" to indicate the position of the carbon to carbon triple bond.
3. The longest chain chosen for the parent name must include both carbon atoms of the triple bond.
4. The parent chain must be numbered from the end nearest a triple bond carbon atom.
5. If the triple bond is in the center of the chain, the nearest substituent rule is used to determine the end where numbering starts.
6. The smaller of the two numbers designating the carbon atoms of the triple bond is used as the triple bond locator.

Example 1

Name the alkyne shown below.

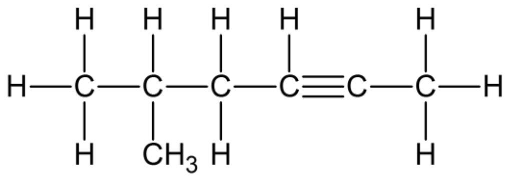


Solution

- The longest chain contains 4 carbon atoms. Thus parent name is butyne.
- There is no side branch in this alkyne.
- The triple bond falls on carbon number 1. Therefore the name of the alkyne is **but-1-yne** or **1-butyne.**

Example 2

Name the organic compound shown below.



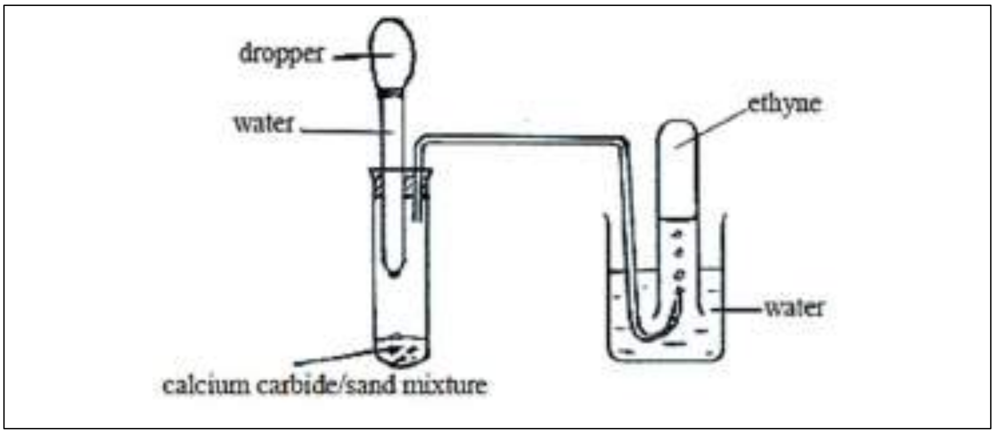
Solution

The compound is an alkyne since it contains a carbon to carbon triple bond.

- The longest chain with the triple bond has 6 carbon atoms. Thus the parent name is hexyne.
- The triple bond falls on the second carbon atom. Thus, the parent name will be: hex-2-yne.
- The side branch has 1 carbon atom, thus will be named, methyl.
- The side branch falls on carbon number 5 (*counting from the side which gives the carbon with a triple bond a lower number).*Thus 5-methyl.
- Therefore the name of the above compound is: **5-methylhex-2-yne**.

Laboratory preparation of ethyne gas

Ethyne gas can be prepared by the reaction of water with calcium carbide (CaC_2).
Ethyne is a colorless gas.



Source: Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.

Reaction equation: $CaC_{2(s)} + 2H_2O_{(l)} \rightarrow C_2H_{2(g)} + Ca(OH)_{2(aq)}$

Chemical reactions of alkenes and alkynes

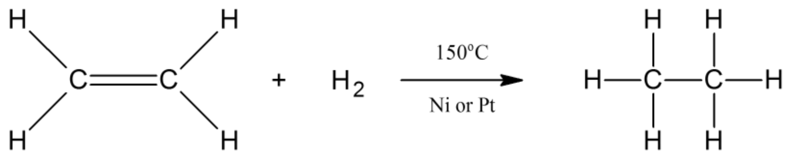
- These groups of organic compounds are more reactive than alkanes.
- The most characteristic reaction of alkenes and alkynes is **addition** to the carbon-carbon double bond or carbon to carbon triple bond.
- The double bond is broken and in its place single bonds form to two new atoms or groups of atoms.
- The triple bond is broken and in its place single bonds form to four new atoms or groups of atoms.

- There are four important addition reactions:
  1. Addition of hydrogen (**Hydrogenation or Reduction**).
  2. Addition of hydrogen halides (**Hydrohalogenation**).
  3. Addition of water (**Hydration**).
  4. Addition of bromine and chlorine (**Halogenation**).

1. Hydrogenation

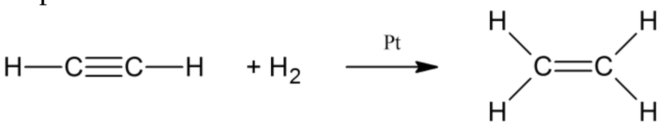
- Atoms of hydrogen add to the carbons in a double or triple bond to form alkanes.
- A catalyst such as platinum (Pt) or nickel (Ni) is added to catalyze the reaction.

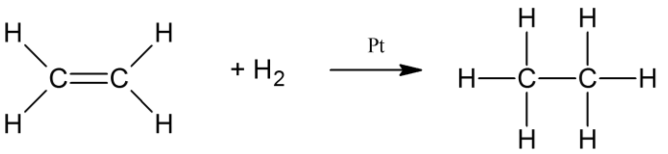
Example 1: Hydrogenation of alkenes



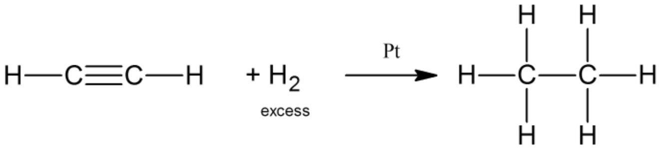
Example 2: Hydrogenation of alkynes

During the catalytic hydrogenation of ethyne, ethene is formed first which in the next step is further reduced to ethane.





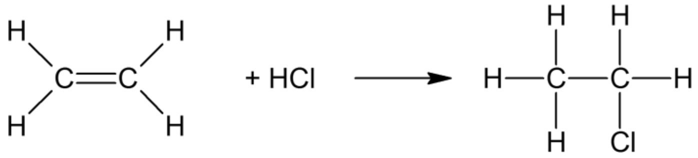
However if excess of H_2 is used, then ethyne is converted to ethane in a single step as shown below.



Hydrohalogenation (addition of hydrogen halides)

- Molecules such as HF, HCl, HBr and HI are added to alkenes.

Example:

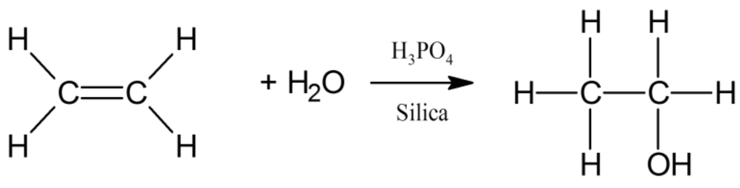


Chloroethane

2. Hydration

- Hydration is the addition of water molecules to alkenes to form alcohols.

Example: Ethene reacts with steam at 300°C and a pressure of 70 atm to form ethanol.

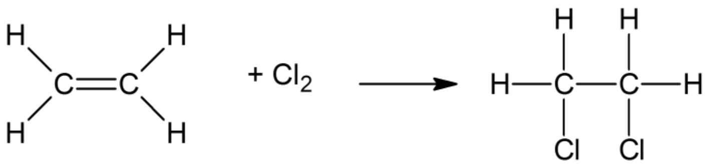


Ethanol

3. Halogenation

- Halogenation is the addition of halogens such as chlorine and bromine to alkenes and alkynes.
- Chlorine and bromine react with alkenes and alkynes at room temperature with addition of halogen atoms to the carbon atoms of the double bond or triple bond.

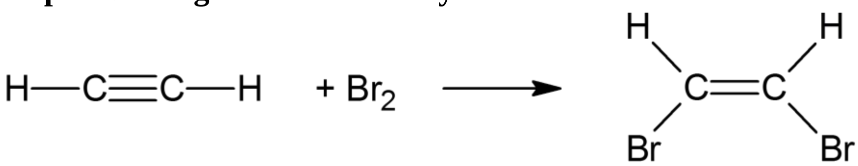
Example 1: Halogenation of an alkene



1,2-dichloroethane

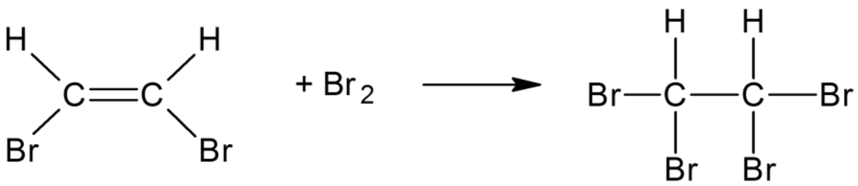
Example 2: Halogenation of an alkyne

i



1,2-dibromoethane

ii.



1,1,2,2-tetrabromoethane

Note:

Addition of bromine is a useful qualitative test for unsaturation (presence of an alkene or an alkyne). When an alkene or alkyne is bubbled through red brown bromine water, the bromine water gets decolorized. The reaction is faster and more vigorous for alkynes then alkenes. The disappearance of the red color as bromine adds to the double bond or triple bond indicates that alkene or alkyne is present.

Isomerism

Isomerism occurs where two or more compounds have the same molecular formula but different structural formulas.

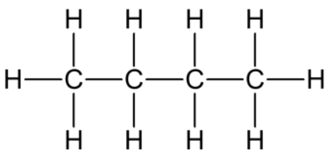
Two common types of isomers are:

1. **Structural (also known as constitutional isomers)**

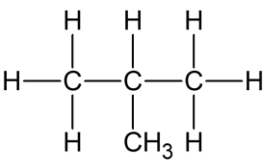
- Structural isomerism occurs when the compounds have the same number and types of atoms but are arranged in different ways.
- The number of possible isomers increases rapidly as the length of the chain increases.

Example

The two structural isomers of the hydrocarbon with molecular formula C_4H_{10} are:



and



Butane

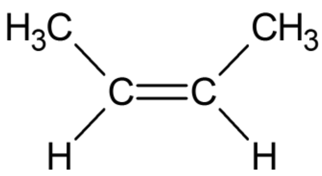
2-methlypropane

2. **Geometrical Isomers**

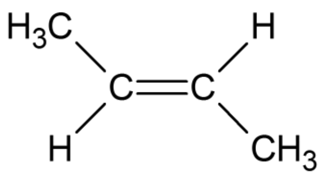
- The double and triple bonds found in alkenes and alkynes are fixed and you will not be able to rotate it.
- This allows the atoms or the groups of atoms bonded to the carbons of the double bonds to have different arrangements. These different arrangements are called geometrical isomers.
- When the groups are on the same side of the double or triple bond, they area called **cis** isomers.

When the groups are on opposite sides of the double or triple bond, they are called **trans** isomers.

Example



cis-but-2-ene



trans-but-2-ene

Exercise

1. The reaction of bromine with ethyne is an example of what type of reaction?
   A. Substitution
   B. Halogenation
   C. Hydrohalogenation
   D. Hydration

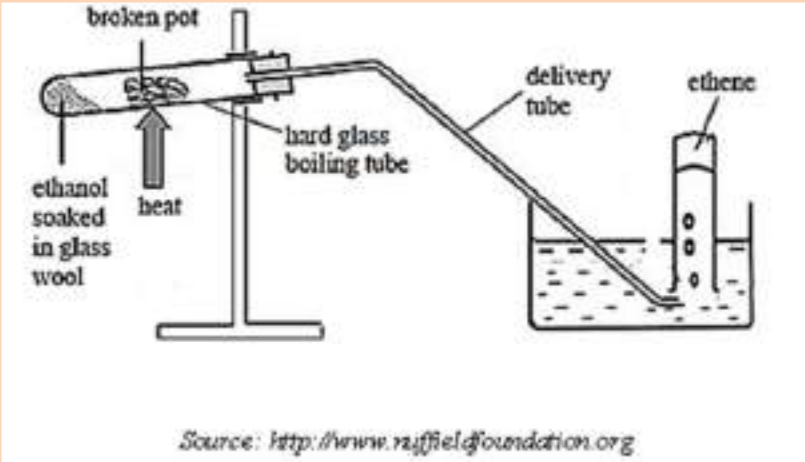
2. When the colorless gas ethene, C_2H_4 is bubbled through bromine water, the solution will turn
   A. clear.
   B. dark brown.
   C. blue.
   D. pink.

3. The representation C_2H_4 for ethene is
   A. the empirical formula.
   B. the molecular formula.
   C. the structural formula.
   D. the line-angle formula.

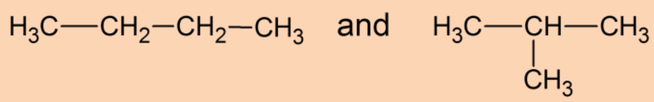
4. A student found two unlabeled gas cylinders in the laboratory. Through discussion with some teachers, he found out that one cylinder contained propane and the other contained ethyne gas. In order to identify them, the student decided to carry out two tests with each of the gases.
   Test 1
   A sample of each gas was bubbled through separate solutions of bromine.
   Test 2
   A lightened match stick was placed at the mouth of a jar of each of the gases.

Describe what the student would observe in each of the tests and how these observations will help the student identify the gases. *(Use balanced equations where applicable to explain your answer).*

5. Write equations to show the complete and incomplete combustion reactions of ethane gas.
6. Explain with a suitable example what you understand by the phrase "alkanes undergo substitution reactions".
7. Give the product of the addition reaction of ethene with:
    a. Hydrogen (H_2)
    b. Hydrogen chloride (HCl)
    c. Water (H_2O)
8. Name a catalyst which is used to convert ethene to ethane.
9. The diagram below shows the experimental set-up for the preparation of ethene gas.



    i. State the method of ethene gas production shown above.
    ii. What function does the broken pot perform in the above preparation.
    iii. Write a balanced equation for the reaction.
    iv. A lighted match stick was introduced to the mouth of the test tube full of ethene gas. Describe an observation that would be made.
    v. State two precautionary measures that should be taken when carrying out this experiment.
10. Give one relationship between the following compounds.



11. The molecular formula for butene is C_4H_8. Name and draw the structural formula of the two geometrical isomers of butene.

4.2.2 Alcohols

- Alcohols are recognised by the presence of the hydroxyl functional group (–OH) bonded to a carbon atom of an alkyl or substituted alkyl group.
- The hydroxyl functional group strongly contributes to the physical properties of alcohols.
- The general formula of alcohol is $C_nH_{2n+1}OH$.
- Alcohols are represented as: R-OH, where R is the alkyl group and OH is the hydroxyl group.

Properties of alcohols

- Alcohols are colorless.
- Alcohols have much higher boiling points than other compounds of similar molecular weight. The boiling point of alcohols increases with the increasing number of carbon atoms.
- Small alcohols are miscible in water. However, once a certain number of carbons in the alkyl chain of the alcohol is reached, the alcohol is no longer soluble.
- The longer the carbon chains in an alcohol, the lower the solubility in polar solvents and the higher the solubility in nonpolar solvents.
- Alcohols with less than seven carbon atoms are liquid at room temperature.
- The smaller alcohols are more volatile than the larger ones.

IUPAC Naming of alcohols

- Naming alcohols is similar to alkenes and alkynes with some exceptions.
- Alcohols are named by replacing the terminal “e” of the corresponding alkane by “ol”. Wherever necessary, a number is written in front of “ol” to indicate the position of the -OH functional group.
- While numbering the carbon atoms, priority should be given to the hydroxy group over the alkyl group.

Steps

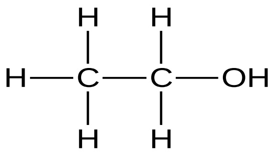
1. Select the longest carbon chain containing the functional group -OH.
2. Number the chain with the -OH group getting the lowest possible number.
3. Replace the “e” at the end of the suffix of the alkyl chain with -ol.
4. If necessary, add a prefix number to indicate which carbon the -OH group is bonded to.

The first six members of the alcohol series

Number of carbon atoms in the longest chain	Parent name
1	Methanol
2	Ethanol
3	Propanol
4	Butanol
5	Pentanol
6	Hexanol

Example

1. Name the alcohol given below.



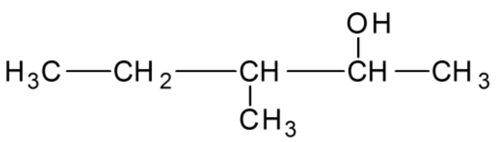
Solution

i. The longest chain with the –OH group contain two carbon atoms, thus the parent name is ethanol.

(Note: The alkane containing two carbon atoms is ethan**e**, thus it has now changed to ethan**ol**.)

ii. Since there are no other substituents, the name of the alcohol is **ethanol.**

2. Name the alcohol given below.



Solution

i. The longest chain with the –OH group contains five carbon atoms, thus parent name will be pentanol.

ii. While numbering the carbon atoms from the side which would give the –OH group a lower number (from right to left), it can be seen that the –OH group falls on carbon number two.

iii. Therefore, the parent name will be pentan-2-ol.

iv. There is one side branch on carbon number 3.

v. The side branch is an alkyl group containing one carbon atom, thus will be named methyl. The position and name of the side branch would be: 3-methyl.

vi. The name of the alcohol is therefore: **3-methylpentan-2-ol.**

Exercise

1. Name the following alcohols.

A. $CH_3-CH_2-CH(CH_3)-CH_2-OH$

```
A.  CH3-CH2-CH-CH2-OH
             |
             CH3
```

B. $CH_3-CH_2-CH(OH)-CH_2-CH_2-CH_3$

```
B.  CH3-CH2-CH-CH2-CH2-CH3
             |
             OH
```

C. $CH_3-CH_2-C(CH_3)_2-CH_2-OH$

```
             CH3
              |
C.  CH3-CH2-C-CH2-OH
              |
             CH3
```

2. Draw the structure of the following alcohols.

i. 2-methylpentan-3-ol ii. 2,3-dimethylbutan-1-ol iii. 2-ethylhexan-3-ol

Chemical reactions of alcohols

1. Combustion reaction

- Alcohols burn with blue flame releasing heat and light energy and producing carbon dioxide and water.

Complete Combustion	Incomplete Combustion
$C_2H_5OH_{(l)} + 6O_{2(g)} \rightarrow 2CO_{2\,(g)} + 3H_2O_{(l)}$	$C_2H_5OH_{(l)} + 2O_{2(g)} \rightarrow 2CO_{(g)} + 3H_2O_{(l)}$

2. Dehydration

- Dehydration is done by heating the alcohol with concentrated sulfuric acid (which acts as a dehydrating agent).
- In this reaction, alcohols are used to produce the corresponding alkenes and water.

Example

$$C_2H_5OH \xrightarrow[\Delta]{Con.H_2SO_4} C_2H_4 + H_2O$$

3. Oxidation of alcohol

- Alcohols can be oxidized to carboxylic acid by oxidizing agents such as acidified potassium dichromate ($Cr_2O_7^{2-}/H^+$).

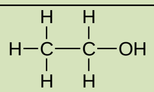
Example

$$CH_3CH_2OH \xrightarrow[Cr_2O_7^{2-}]{H^+} CH_3COOH$$

- Oxidation of alcohol is used commonly in the breathalyzer test. The breathalyzer test determines the amount of alcohol in a person's breath. Acidified potassium dichromate is placed in glass tubing which is attached to a plastic bag. The person is allowed to blow air into the glass tubing. The alcohol in the breath changes the orange dichromate ions ($Cr_2O_7^{2-}$) to green chromium ions (Cr^{3+}). The time in which the orange crystals turn green indicates the concentration of alcohol in the person's breath.

> *For your practical, see the experiment on **Oxidation of alcohol**; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Quick Exercise: Copy and complete the table below.

Alcohols						
General Formula	**$C_nH_{2n+1}OH$**					
Alcohol	**Methanol**	**Ethanol**	**Propanol**	**Butanol**	**Pentanol**	**Hexanol**
Molecular Formula	CH_4O	C_2H_6O	C_3H_8O	C_4H_9OH	$C_5H_{11}OH$	$C_6H_{13}OH$
Expanded structural formula						
Condensed structural formula	CH_3OH	CH_3CH_2OH or CH_3-CH_2-OH	$CH_3CH_2CH_2OH$ or CH_3-CH_2-CH_2-OH	$CH_3CH_2CH_2CH_2OH$ or $CH_3(CH_2)_2CH_2OH$ or CH_3-CH_2-CH_2-CH_2-OH		

4.2.3 Carboxylic Acids

- Carboxylic acids are recognised by the presence of a carboxyl (-COOH) functional group bonded to a carbon atom of an alkyl or substituted alkyl group.
- The general formula for carboxylic acids is $C_nH_{2n}O_2$.
- Carboxylic acids are represented as: R-COOH, where R is the alkyl group and COOH is the carboxyl group.

Properties of carboxylic acids

- Carboxylic acids are weak acids which react in the same way as dilute mineral acids.
- The solubility of acids decreases as the number of carbon atoms increases.
- Small carboxylic acids are readily soluble in water.
- They only partially dissociate into H^+ and $R\text{-}COO^-$ in aqueous solutions.
- They have a higher melting and boiling points compared to the relative hydrocarbons and alcohols.
- Most often have strong odors of its derivative.

IUPAC naming of carboxylic acids

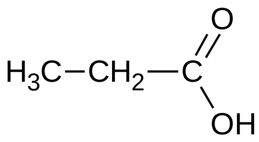
- Carboxylic acids are named like alcohols, with some exceptions.
- The parent chain is the one with the greatest number of carbon atoms and includes the functional group (-COOH).
- Carboxylic acids are named by replacing the terminal "e" of the corresponding alkane with "**oic acid**".
- While numbering the carbon atoms, priority should be given to the carboxyl group over the alkyl group and ensure that a lower number is given to the carbon which contains the -COOH group.

Number of carbon atoms in longest chain	Parent name
1	Methanoic acid
2	Ethanoic acid
3	Propanoic acid
4	Butanoic acid
5	Pentanoic acid
6	Hexanoic acid

The first six members of the carboxylic acid series

Example

1. Name the carboxylic acid shown below.

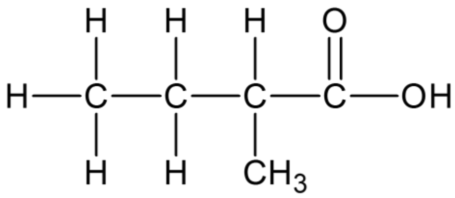


Solution

The longest continuous carbon chain which includes the -COOH functional group has three carbon atoms. It has no side branches. Therefore the name is **propanoic acid**.

(Note: The alkane containing three carbon atoms is propan**e**, thus it has now changed to propan**oic acid**).

2. Name the carboxylic acid shown below.



Solution

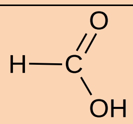
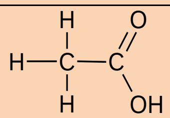
i. The longest chain with the -COOH group contains four carbon atoms, thus parent name will be butanoic acid.

ii. There is one side branch on carbon number 2.

iii. The side branch is an alkyl group containing one carbon atom, thus will be named methyl. The position and name of the side branch will be 2-methyl.

iv. Thus, the name of the carboxylic acid would be **2-methylbutanoic acid**.

Quick Exercise: Copy and complete the table below

Carboxylic acids						
General Formula	$C_nH_{2n}O_2$					
Alcohol	Methanoic acid	Ethanoic acid	Propanoic acid	Butanoic acid	Pentanoic acid	Hexanoic acid
Molecular Formula	CH_2O_2	$C_2H_4O_2$	$C_3H_6O_2$	$C_4H_8O_2$	$C_5H_{10}O_2$	$C_6H_{12}O_2$
Empirical Formula			$C_3H_6O_2$			
Expanded structural formula						
Condensed structural formula	HCOOH			$CH_3CH_2CH_2COOH$ or $CH_3(CH_2)_2COOH$ or CH_3-CH_2-CH_2-COOH		

Reactions of Carboxylic acids

1. Reaction with water

- The carboxylic acids are acidic because of the hydrogen in the -COOH group.
- If mixed with water, a hydrogen ion is transferred from the -COOH group to a water molecule.
- For example, when ethanoic acid is mixed with water, an ethanoate ion is formed together with a hydronium ion.

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

2. Reaction with base

- These are simple neutralisation reactions and are just the same as any other reaction in which hydrogen ions from an acid react with hydroxide ions.
- They are most quickly and easily represented by the equation:

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

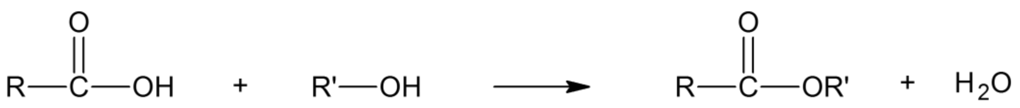
- For example, if dilute ethanoic acid is mixed with sodium hydroxide solution, a colorless solution containing sodium ethanoate is formed.
- This is an exothermic reaction thus temperature of the reaction mixture will increase.

$$CH_3COOH_{(aq)} \;+\; NaOH_{(aq)} \;\rightarrow\; CH_3COONa_{\,(aq)} \;+\; H_2O_{(l)}$$

Sodium ethanoate

3. Reaction with alcohol

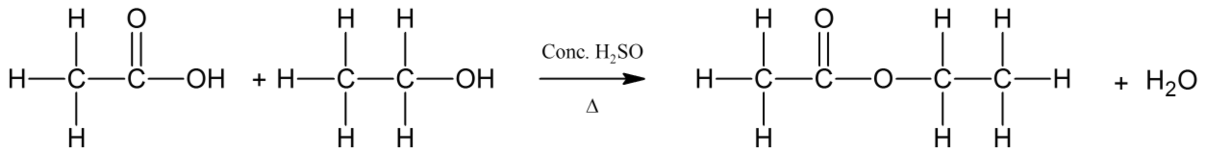
- The condensation reaction between an alcohol and a carboxylic acid produces the organic compound called **esters**.
- Esters have characteristic smells and are insoluble in water.
- Esterification is the process where an alcohol reacts with a carboxylic acid under acidic conditions and high temperature.
- This process is known as refluxing. Sulfuric acid is added to the mixture while it is being refluxed since it acts as a catalyst to allow more ester to be formed during the reaction. It also removes the water formed and acts to create the acidic conditions necessary to produce the ester.
- Sodium carbonate is added to the mixture to neutralize any unchanged acid. The ester formed is immiscible with water and can be separated by decantation.
- The general equation for the production of esters from carboxylic acids and alcohol is:



IUPAC Naming Esters

1. Change the name of the parent alcohol to end in –yl.
2. Change the name of the parent acid to end in –oate.
3. Combine the alcohol and the acid name (*The alcohol name is written first followed by the acid name).*

Example: When ethanol is refluxed with ethanoic acid in the presence of sulphuric acid catalyst, the ester **ethyl ethanoate** is formed.



Ethyl ethanoate

4.2.4 Production of some organic substances

1. Polyvinylchloride (PVC)

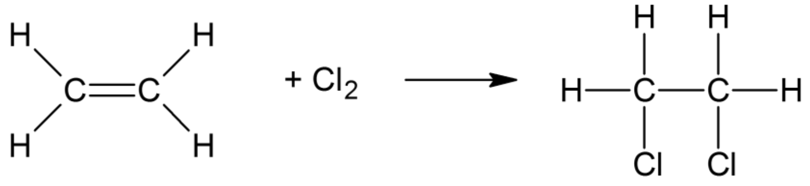
- Polyvinylchloride (PVC) is a major plastic material which has many uses in building, transport, packaging, electrical/electronic and healthcare applications.
- They are used for making water pipes, garden hose pipes and electrical insulators.
- Some properties of PVC are: It is long lasting, cheap and light, does not corrode and water resistant.

Production of PVC

PVC is made from vinyl chloride (chloroethane) molecules.

Step 1

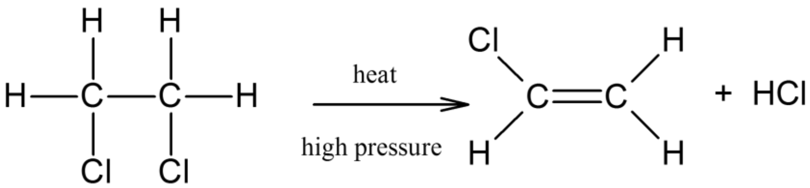
- Chlorine is added to ethene to form 1, 2-dichloroethane.



1, 2-dichloroethane

Step 2

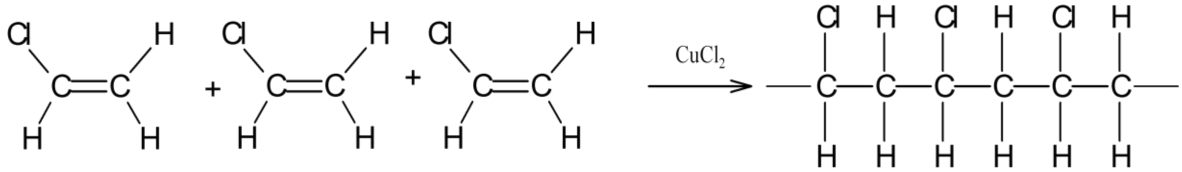
- 1,2-dichloroethane is then heated under high pressure to form vinyl chloride (chloroethene) and hydrogen chloride.



Chloroethene

Step 3

- The hydrogen chloride in step 2 is removed by dissolving in water and the vinyl chloride is heated in the presence of a catalyst, normally copper chloride ($CuCl_2$) to produce polyvinyl chloride. This process is known as polymerization.



(Part of PVC)

2. **Methanol**

Methanol can be made from natural gas.

Step 1

- Sulphur and sulphur compounds are removed from natural gas.

Step 2

- The alkane (methane) is decomposed using steam and nickel catalyst and converted to synthesis gas using nickel catalyst at around 800-850°C.
- This process is known as steam reforming and the product obtained are carbon monoxide and hydrogen gas.

$$CH_{4(g)} + H_2O_{(l)} \xrightarrow[800\,^{\circ}C]{Ni} CO_{(g)} + 3H_{2(g)}$$

Step 3

- The synthesis gas is then cooled, compressed and reheated in a methanol converter using copper oxide catalyst.
- The products from the converter undergo fractional distillation to produce methanol.

$$CO_{(g)} + H_{2(g)} \xrightarrow{CuO} CH_3OH_{(g)}$$

Methanol

3. **Ethanol**

Ethanol can be produced by fermenting sugar under the enzymes present in the yeast.

<u>Step 1</u>

- Cane sugar (sucrose) is hydrolyzed to glucose and fructose using the enzyme invertase.

$$C_{12}H_{22}O_{11} + H_2O \xrightarrow{Invertase} C_6H_{12}O_6 + C_6H_{12}O_6$$

Sucrose Glucose Fructose

<u>Step 2</u>

- Glucose and fructose produced in step 1are decomposed to ethanol and carbon dioxide using the enzyme zymase.

$$C_6H_{12}O_6 \xrightarrow{Zymase} 2C_2H_5OH + 2CO_2$$

<u>Step 3</u>

- Ethanol is collected by fractional distillation which yields a mixture containing about 96% ethanol.

*For your practical, see the experiment on **Preparation of ethanol**; Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.*

Exercise

1. Draw the structure of the following carboxylic acids.
   i. 2-methlypentanoic acid
   ii. 3,3-dimethlybutanoic acid
   iii. 2,2-dimethylpropanoic acid

2. Ethanol is warmed (refluxed) with ethanoic acid to which a few drops of concentrated sulphuric acid has been added.
   i. Write balanced equation for the reaction that occurs between ethanol and ethanoic acid.
   ii. What is the purpose of sulphuric acid?
   iii. To what class of organic compound does the product belong to?
   iv. State one distinctive property of this product.

3. Copy and complete the table below.

Alcohol used to form ester	**Carboxylic acid used to form ester**	**Name of ester formed**	**Expanded structural formula of ester**
Methanol	Methanoic acid		
Methanol	Ethanoic acid		
Ethanol	Methanoic acid		
Methanol	Propanoic acid		

4. Briefly explain how the breathalyzer tests works.

5. Explain what is meant by the term functional group.

6. Draw the expanded structural formula of butanol and circle its functional group.

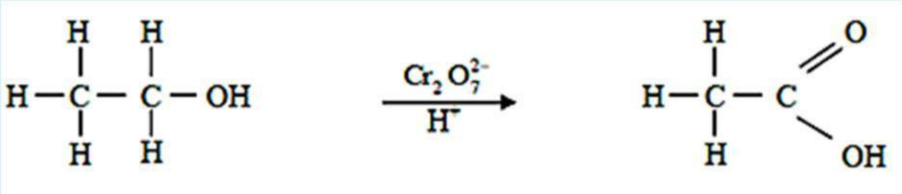
7. Name the two organic compounds which can be used to prepare the ester, $CH_3\ (CH_2)_2COOCH_2CH_3$ which is responsible for the odor of pineapples.

8. The ester, propyl methanoate, can be prepared in the laboratory by reacting propanol and methanoic acid.

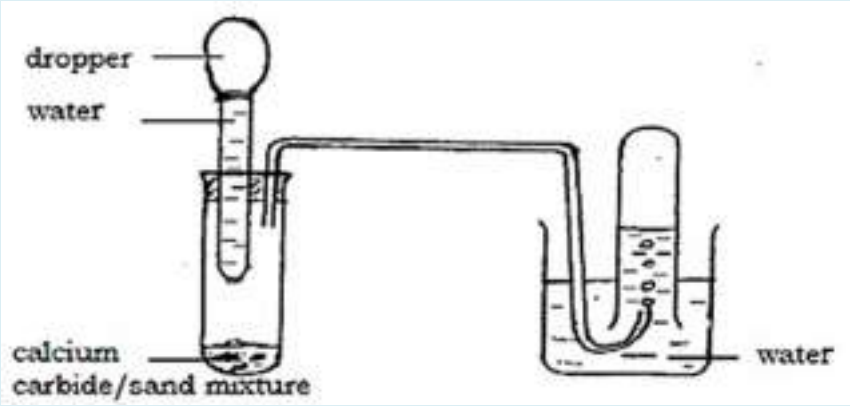
   i. Write the expanded structural formula of propyl methanoate.
   ii. Once the ester propyl methanoate is prepared, aqueous sodium carbonate is added. Explain why aqueous sodium carbonate is added.
   iii. Give one other condition that is necessary for the preparation of propyl methanoate.
   iv. State two uses of esters.

9. Name a compound that can react with sodium hydroxide to give a salt and water.

10. Identify the type of reaction shown below.



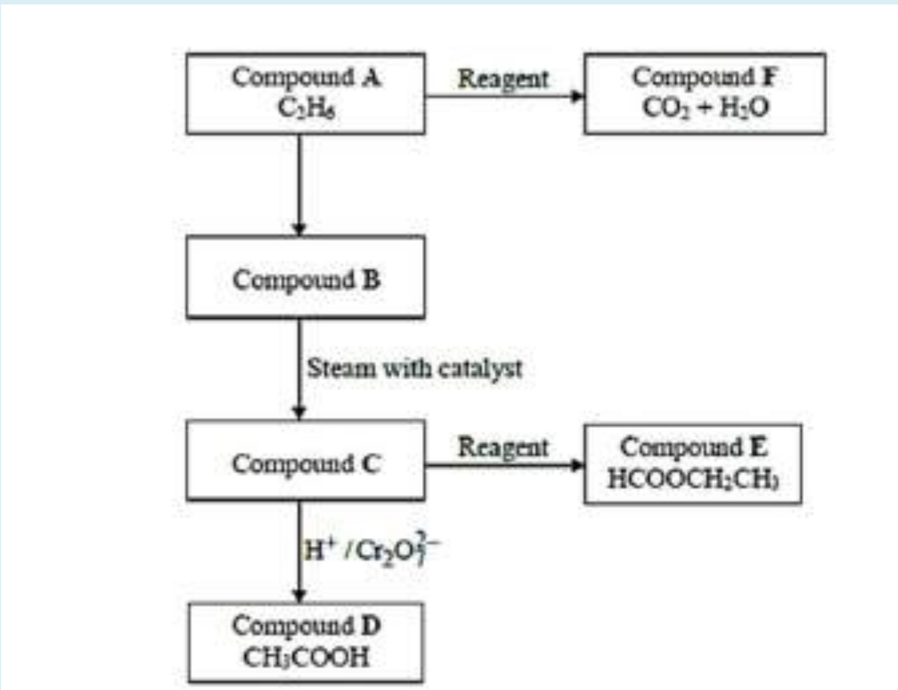
11. The experimental set-up given below shown how ethyne gas can be prepared in the school laboratory.



Source: Experiments in Sixth Form Chemistry, Students Laboratory Manual, Ministry of Education, Fiji.

i. Write the balanced chemical reaction for the production of ethyne as shown above.
ii. Why is calcium carbide mixed with sand?
iii. State one observation made in the reaction of bromine with ethyne.
iv. How does the observation in (iii) above compare with the observation of the reaction of bromine with ethene?

12. Consider the flow diagram shown below and answer the questions that follow.



i. Name the Compounds B and C.

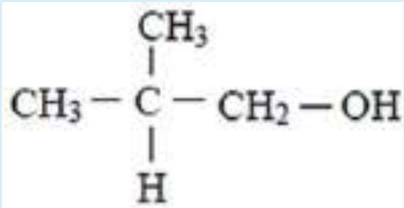
ii. Name the reagent needed to convert:

  I. Compound A to Compound F.

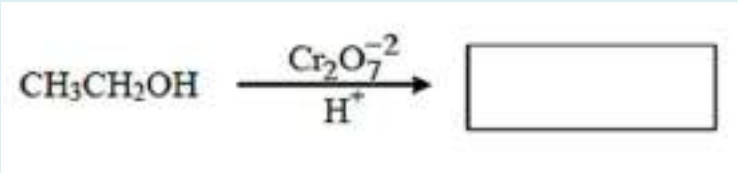
  II. Compound C to Compound E.

iii. Compound B decolorizes bromine water when bubbled through it. Draw the expanded structural formula of the product and give its name.

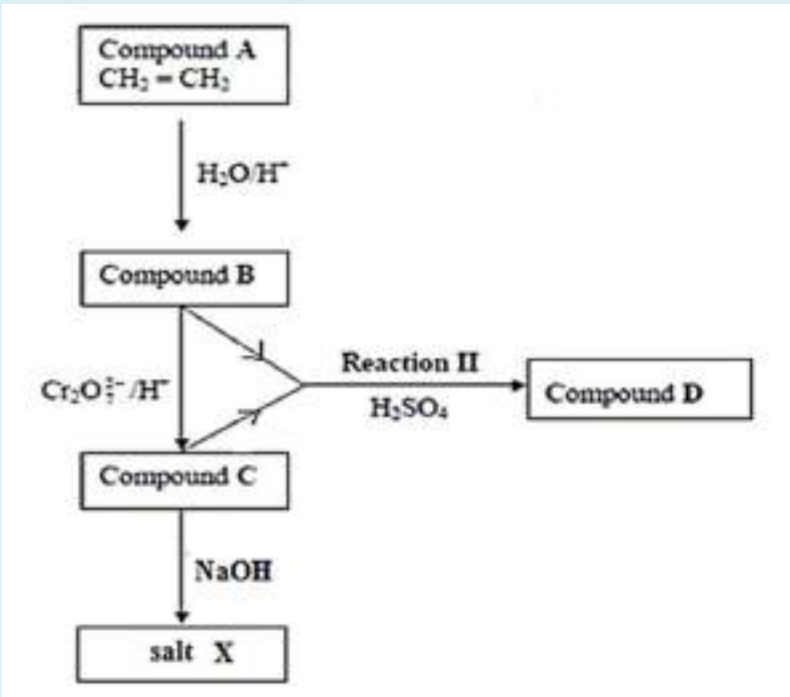
13. Name the following compound.



14. Complete the equation shown below.

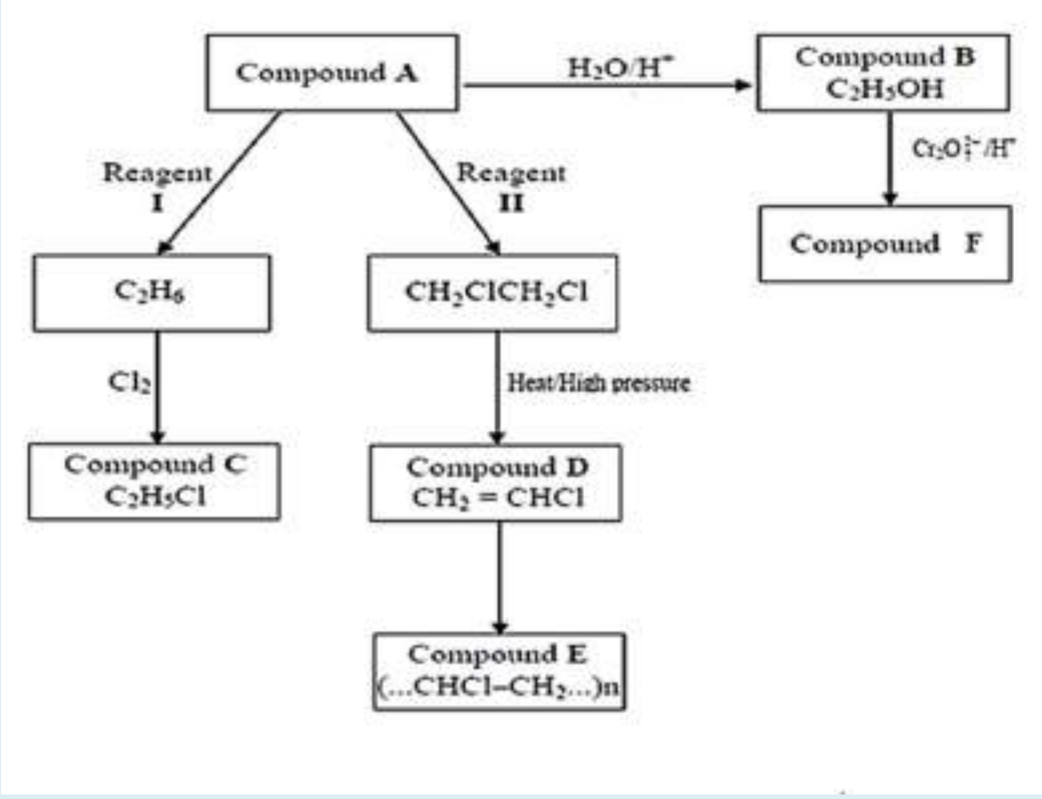


15. Use the reaction sequence in the following diagram to answer the questions that follow.



i. Name the compounds B, C and D.
ii. Name the reactions I and II.
iii. Draw the expanded structural formula of Compound D.
iv. Give the name of salt X.
v. Give one use of compound D.

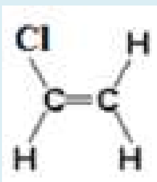
16. Consider the following reaction sequence to answer the questions that follow.



i. Name Compounds A, B, C, D.
ii. Draw the expanded structural formula of Compound F.
iii. Give the name and formula of the reagents I and II.
iv. Name the process that converts Compound D to Compound E.
v. What type of reaction does Compound A undergo to be converted to C_2H_6.
vi. Name the process that converts Compound B back to Compound A.
vii. Give the name of a reagent and the process that will convert Compound B back to Compound A.

17. Name and draw the structural formula of the main organic product in each of the following reactions given below.
   i. Ethyne reacting with excess chlorine
   ii. Propene reacting with hydrogen
   iii. Fermentation of cane sugar

18. The structural formula of vinyl chloride, a monomer of polyvinyl chloride (PVC) is given below.



i. Using the structural formula, rewrite the following equation.
   Vinyl chloride + vinyl chloride → Part of PVC molecule.
ii. Give two properties of PVC which make it useful for making water pipes.

Strand Outcome

Demonstrate an understanding of the chemicals that are encountered in everyday life.

Sub-strands

5.1 Food Chemistry
5.2 Chemistry of Medicines and Drugs
5.3 Household Chemistry

5.1 Food Chemistry

Achievement Indicators

Upon completion of this sub-strand, students will be able to:

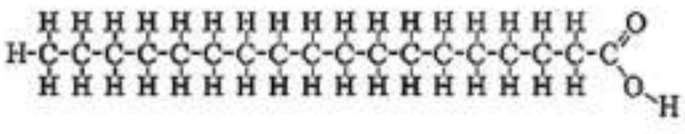
- ✓ Describe the sources, properties, uses and reactions of fats and oils.
- ✓ Describe and distinguish between the different types of fatty acids.
- ✓ Describe the processing of fats and oils.
- ✓ Describe the different types of carbohydrates.
- ✓ Describe the chemical nature, properties, uses and reactions of carbohydrates.
- ✓ Describe the tests for glucose.
- ✓ Describe the chemical nature, properties and reactions of amino acids.
- ✓ Describe the purpose of common added agents to processed foods.

5.1.1 <u>Fats and Oils</u>

- Fats are esters of fatty acids and glycerol.
- Fats provide a very efficient way for the body to store energy and it also forms part of cell membranes.
- Lipids are any classes of organic compounds that are fatty acids or their derivatives and are insoluble in water but soluble in organic solvents.
- Fatty acids have a long hydrocarbon chain with a carboxyl (acid) group. The chains usually contain 16 to 18 carbons.
- Fatty acids do not usually exist as "free" fatty acids, but are often in sets of two or three.
- Triglycerides are sets of three fatty acids bound together. Triglycerides are the form in which fat is carried through the blood to the body tissues.
- Most often, triglyceride is called a fat if it is a solid at 25°C and it is called an oil if it is a liquid at 25°C. Therefore, an oil is any triglyceride which is liquid at 25^{o}C.
- The differences in melting points of fats and oils reflect differences in the degree of unsaturation and the number of carbon atoms in the constituent fatty acids.
- Triglycerides obtained from animal sources are usually solids, while those of plant origin are generally oils.

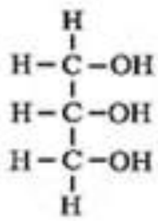
<u>Example</u>

<u>A fatty acid</u>



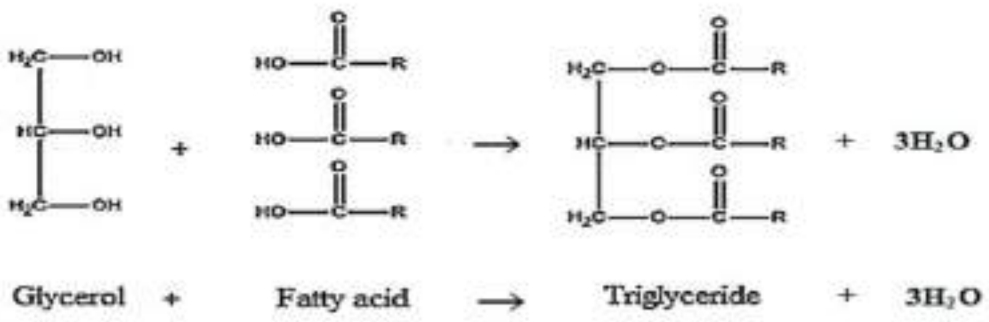
Molecular formula - $C_{18}H_{36}O_2$

<u>Glycerol</u>



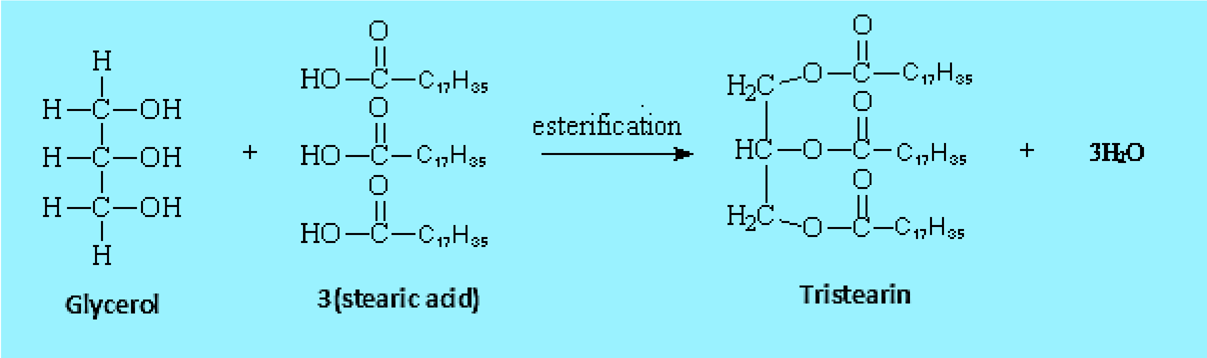
Molecular formula - $C_3H_8O_3$

- Glycerol reacts with 3 fatty acids to form a triglyceride (an ester).



Note: R- represents an alkyl group in the fatty acid.

The example below shows the formation of a triglyceride from stearic acid ($C_{18}H_{36}O_2$).



Sources and uses of fats and oils

- Animal fats and oils are derived from both terrestrial and marine animals. Marine fats include liver oils, blubber oils and fish oils.
- Vegetable fats and oils are found in greatest abundance in fruits and seeds.
- Fats and oils are the primary constituents of margarines, butter, shortenings and oils used in salads and cooking.
- Fats and oils are also found in high quantities in many bakery goods, infant formulas, dairy products and some sweets.
- Oils, butter or margarine are sometimes used directly on food.

Fats and Oil Processing

- The fats and oils derived from plant and animal sources are subject to several commercial refining processes before the final products reach the consumer market.
- The basic steps for extraction and processing of fats and oil are:

1. Extraction
2. Refining
3. Bleaching
4. Deodorization

1. Extraction

- Vegetable oils are extracted by heating the seeds of the relevant plant with organic solvents (such as hexane) and/or with pressure and then extracting the oil.
- Animal fats are extracted when oil-bearing tissues are chopped into small pieces and boiled in water. The oil floats to the surface of the water and it is then skimmed.
- Plant oils can also be extracted through steam distillation.

2. Refining

- Free fatty acids, pigments and waxes exist in the extracted oil which can lead to undesirable properties of the final products.
- Many of these impurities are removed by treating fats at 40 ˚C to 85 ˚C with caustic soda solution (sodium hydroxide) or sodium carbonate.
- The free fatty acids react with the alkali forming soap.
- The impurities settle at the bottom and are drawn off.
- The soap dissolves in water and the triglyceride (oils) float on top.
- These are then separated by centrifuge and the oils are then washed to remove any residual soap.
- Further purification of the oils is obtained by steam distillation where high pressure steam is injected into the fat at temperatures between 180 – 250 ˚C and a vacuum of about 1 kPa. The volatile products are removed in the steam leaving the fat almost tasteless.

3. Bleaching

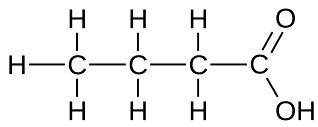
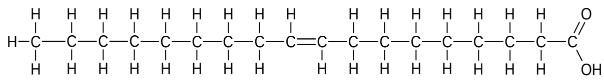
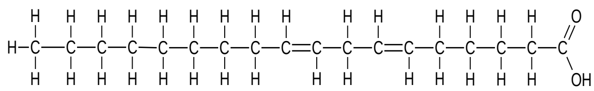
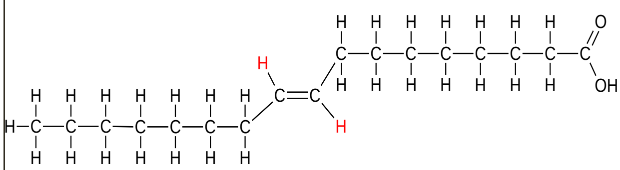
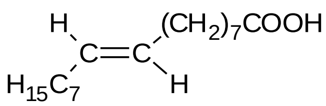
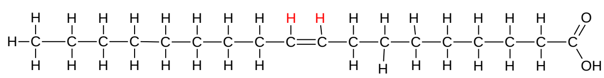
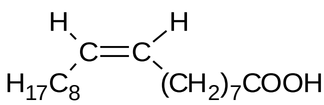
- The major purpose of bleaching is the removal of undesired colored materials in the oil.
- Heated oil (~85 ^oC) may be treated with various bleaching agents such as fuller's earth, activated carbon or activated clays.
- Many impurities, including chlorophyll and pigments, are adsorbed onto such agents and removed by filtration.

4. Deodorization

- Deodorization is the final step in the refining of oils.
- Deodorization involves the use of steam distillation under reduced pressure.
- Volatile compounds with undesirable odors and tastes can be driven off, resulting in an odorless product.
- The oil produced is referred to as "refined oil" and is ready to be consumed or for the manufacture of other products.

Groups of fats

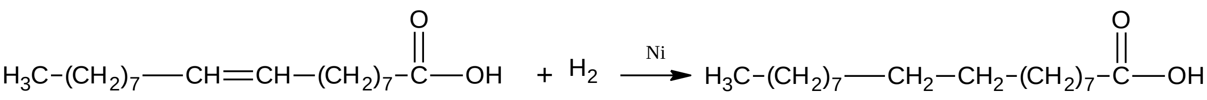
Dietary fat can be classified into several groups. These are summarized in the table below:

Type of fat	Description	Example Structure	Major food source
Saturated fats	These fats contain only single carbon-carbon double bonds in their structure.	 Butyric acid	High fat cuts of meat (beef, lamb, pork), chicken, whole-fat dairy products (milk, cream), Butter, cheese, ice cream, lard
Unsaturated fats	Contains at least one carbon-carbon double bonds in its structure.		
i. Monounsaturated	These fats contain only one carbon-carbon double bond in their structure.	 Oleic Acid	Oils (Olive, Canola, Sesame, Peanut Oil, Sunflower oil), Avocados, Olive, Nuts (almonds, peanuts, cashews)
ii. Polyunsaturated fats	These fats contain more than one carbon–carbon double bond in their structure.	 Linoleic Acid	Oils (Soybean, Corn), Walnuts, Soy milk, Fatty fish (Salmon, tuna, mackerel, sardines)
Trans fat	In *trans* configuration, the carbon chain extends from opposite sides of the double bond.	 or 	Commercially baked pastries, cookies, cakes, pizza margarine, vegetable shortening, fried foods(French fries, fried chicken, chicken nuggets, breaded fish), candy bars.
Cis fat	In *cis* configuration, the carbon chain extends from the same side of the double bond.	 or 	Naturally occurring fats and oils.

Hydrogenation

- Unsaturated fatty acids can be converted to saturated fatty acids by the process of hydrogenation using a nickel catalyst at temperatures of about 160 - 220 ˚C and pressure of 2 - 10 atm.
- Hydrogenation is the process of adding hydrogen to the double bonds of the fatty acids. This increases its saturation and leads to hardening of the fat which increases the melting point of fatty acids.

Example: Hydrogenation of oleic acid



- The hydrogenation process is used to convert oils to fats. This is because:
  1. It is easier to handle fats then oils.
  2. Oils tend to be less stable because of their unsaturation. Thus hydrogenation prevents rancidity of oils.

Comparing Butter and Margarine

Butter	Margarine
✓ Butter is made from milk fat and water and contains short chains of saturated fatty acids. ✓ The fats in milk have high degree of saturation and hence are less readily auto-oxidized and are kept better then vegetable oils.	✓ Table margarines are made from vegetable oils and contain polyunsaturated fatty acids. ✓ However, trans fatty acids are created during processing when vegetable oils are hydrogenated or hardened into a solid form which does not melt at room temperature. ✓ Hydrogenation makes the margarine hard. ✓ In general, the more trans-fat a margarine has, the more solid it is.

- Solid fats contain more saturated fatty acids than oils.
- Oils contain more unsaturated and polyunsaturated fatty acids

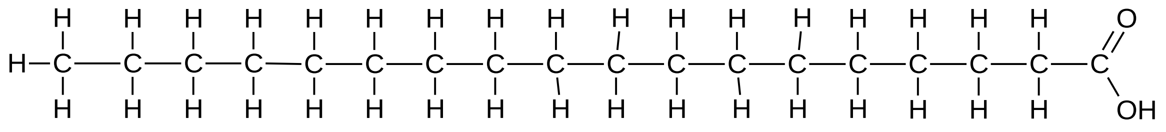
Physical properties of fats and oils

Difference between fats and oils	
Fat	**Oil**
✓ Remains solid at room temperature	✓ Remains liquid at room temperature
✓ Relatively more saturated	✓ Relatively more unsaturated
✓ Relatively high melting point	✓ Relatively low melting point
✓ More stable	✓ Less stable

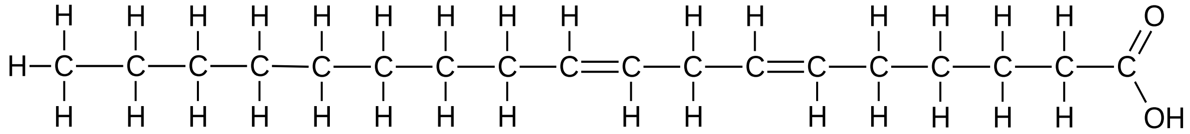
1. **Melting Point**

- ❖ A fat or oil, which is a mixture of several triglycerides do not have a sharp melting point, but melt over a range of temperature depending on the composition of the fat.
- ❖ The melting points of individual fatty acids depend on the chain length of carbon atoms and the degree of unsaturation in the chain.
- ❖ **The longer the chain of carbon atoms, the higher the melting point.**
- ❖ **The more saturated the chain of carbon atoms, the higher the melting point.**

Example: Compare the melting points of the fatty acids given below.



Stearic acid (m.pt. = 69.6 ˚C)



Linoleic acid (m.pt. = -5 ˚C)

Both fatty acids contain the same number of carbon atoms and similar molar masses. However, the stearic acid has no carbon-carbon double bonds, thus a higher melting point. This is because the tetrahedral arrangement of saturated fatty acids such as stearic acids allows them to pack very closely together. Therefore, the van der Waals forces holding the molecules together are stronger, requiring more energy to be broken, leading to higher melting point. On the other hand, unsaturated fatty acids are unable to pack so closely and the van der Waals forces between the molecules are weaker, resulting in a lower melting point.

2. Solubility

- As the chain length of the fatty acid increases, their solubility decreases.
- Fats are generally soluble in organic solvents such as alcohols, chloroform and petrol. They are generally insoluble in water.

3. Smoke Point

- The **smoke point** is the temperature at which the oil is decomposed or broken down.
- The smoke point of various fats is important to note because a fat is no longer good for consumption after it has exceeded its smoke point and has begun to break down.
- The fats that have gone past their smoke points contain large quantities of chemical substances which contribute to risk of cancer and many other diseases.

The table below shows the smoke point of some cooking oils.

Oil	Smoke point	Oil	Smoke point
Canola oil	205 °C	Palm oil	232 ˚C
Soybean oil	230 °C	Corn Oil (Unrefined)	160 ˚C
Coconut oil	177 °C	Refined Corn oil	232 ˚C
Virgin olive oil	216 °C	Canola oil	200 ˚C
Extra virgin olive oil	207 ^{o}C	Sunflower oil	227 ˚C
Mustard Oils	254°C	Sesame oil	210 ˚C
Ghee	252°C	Butter	177 ˚C

Source: http://www.shareyouressays.com

Source: http://www.herbs-info.com

Exercise

1. Briefly explain the following terms:
   i. Saturated fat
   ii. Unsaturated fat
   iii. Hydrogenation
2. What is the purpose of hydrogenation of fats?
3. Describe the extraction and refining processes of fats and oils in the food industry.
4. During the refining process of fats and oils, explain the purpose of washing with strong caustic soda and steam distillation.
5. Briefly describe the relationship between saturation and solubility of fats.
6. Butter and margarine are two common consumer items. Compare these two with respect to their:
   i. melting points.
   ii. degree of unsaturation of fatty acids.
   iii. chain length.
7. Lipids are esters obtained from glycerol and long chain fatty acids.
   (i) Draw the structure of a glycerol molecule.
   (ii) The structure of oleic acid is given below.

$$CH_3-(CH_2)_7CH{=}CH(CH_2)_7-\overset{\overset{\displaystyle O}{||}}{C}-OH$$

   Draw the structure of the triglyceride formed when glycerol reacts with three oleic acid molecules.

   (iii) What name is given to the type of reaction in (ii) above?

8. Shown below are the formulas of two long chain fatty acids X and Y.

<u>Fatty Acid X</u>

$$CH_3-CH_2-CH_2-CH_2-CH=CH-CH=CH-CH_2-CH_2-CH_2-COOH$$

<u>Fatty Acid Y</u>

$$CH_3-CH_2-CH_2-CH_2-CH=CH-CH=CH-CH=CH-CH_2-COOH$$

i. Write the molecular formula of the two fatty acids.

ii. Which of the two fatty acids would you expect to have a higher melting point? Give a reason for your answer.

iii. From the two functional groups present in the fatty acids, identify the functional group that would react with sodium hydroxide solution.

9. Study the table given below and answer the questions that follow.

Fatty Acid	No. of double bonds in structure	Melting Point (^oC)
Oleic acid	1 double bond	10.5
Linoleic acid	2 double bonds	-5.0
Linolenic acid	3 double bonds	-11.0

i. Give the relationship between saturation and the melting point of the fatty acids.

ii. Name one process that can convert linolenic acid to linoleic acid.

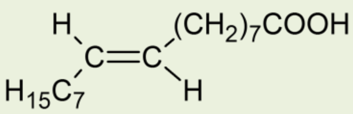
iii. Name the catalyst that can be used in the process mentioned in ii above.

10. Give some differences between fats and oils.

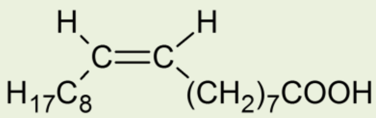
11. What is the significance of smoke points of oils?

12. Consider the structures of two fatty acids given below.

A



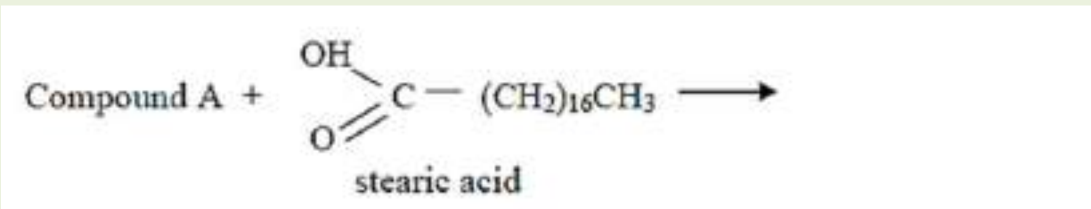
B



Which structure shows?

i. A *cis* fatty acid.
ii. A *trans* fatty acid.
iii. Explain the rationale behind your choices in i and ii above.

13. An organic saturated Compound A has the molecular formula $C_3H_8O_3$. It has three hydroxyl group attached to the carbon chain.
i. Draw the structural formula of compound A and give its systematic name.
ii. Complete the equation for the reaction of compound A with stearic acid in the formation of a lipid molecule.

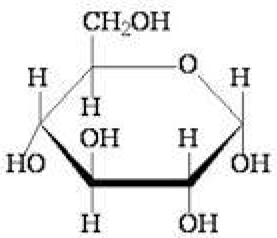
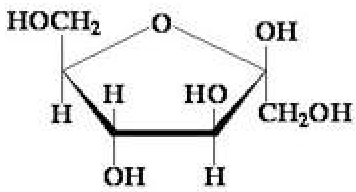


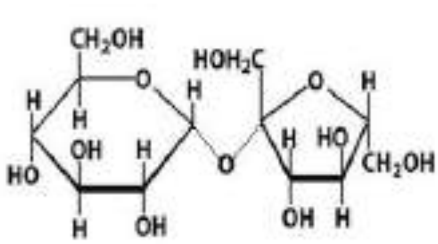
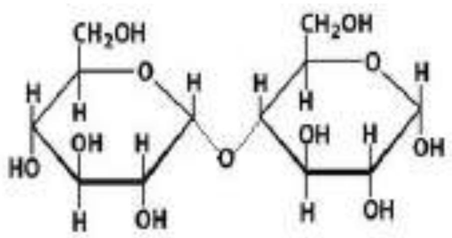
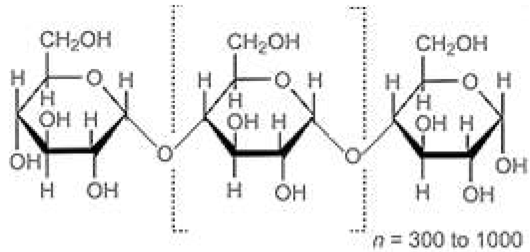
5.1.2 Carbohydrates

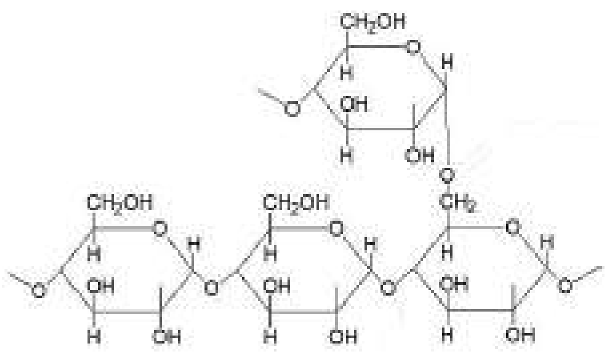
- Carbohydrates are one of the main types of nutrients. They are the most important source of energy for the body.
- The digestive system changes the carbohydrates into glucose (blood sugar) for energy for the cells, tissues and organs.
- Carbohydrates are produced by plants through the process of photosynthesis. The overall process can be represented as:

$$6CO_{2(g)} + 6H_2O_{(l)} \xrightarrow[\text{chlorophyll}]{\text{light}} C_6H_{12}O_{6(s)} + 6O_{2(g)}$$

- Glucose and more simple carbohydrates are then converted to more complicated carbohydrates such as starch or cellulose.
- Carbohydrates can be categorised into three main groups as summarized in the table below.

Molecule	Description
1. Monosaccharides	- Monosaccharides are the simplest form of carbohydrate and all of them have the empirical formula CH_2O. - Examples are glucose ($C_6H_{12}O_6$) and fructose ($C_6H_{12}O_6$). - **Glucose** **Fructose** - These molecules have the same molecular weight but the arrangement of the atoms in each of these molecules is slightly different (see structures above). - Therefore, glucose and fructose are isomers of each other. - Glucose is a monosaccharide that occurs in significant quantities in nature and fructose is a sugar found naturally in many fruits and vegetables, and added to various beverages such as soda and fruit-flavored drinks. - The primary function of monosaccharide is as a source of energy. - In the presence of oxygen, the monosaccharides can be broken down to carbon dioxide and water with the release of energy. This energy is used by the cell to do work. - Monosaccharides serve as "building blocks" for the formation of disaccharides and polysaccharides.

<table>
<tr>
<td>2. Disaccharides</td>
<td>
- Two joined monosaccharides is called a disaccharide.
- These are the simplest polysaccharides.
- Examples include sucrose and maltose.

Sucrose

Maltose
- Sucrose is a disaccharide formed by linking glucose and fructose molecules
- Maltose is a disaccharide formed by linking two glucose molecules.
- The primary function of disaccharides is as a nutritional source of monosaccharides. Many of the sugars found in foodstuffs are disaccharides.
</td>
</tr>
<tr>
<td>3. Polysaccharides</td>
<td>
- Polysaccharides are long chains of monosaccharide units bound together by glycosidic linkages.
- The two important polysaccharides are starch (made by plants) and glycogen (made by animals).
- The function of starch and glycogen is as a major storage form of glucose.
- Polysaccharides can also be used as structural components.


Starch

Glycogen
</td>
</tr>
</table>

<u>Test for glucose with Fehling's or Benedict's solution</u>

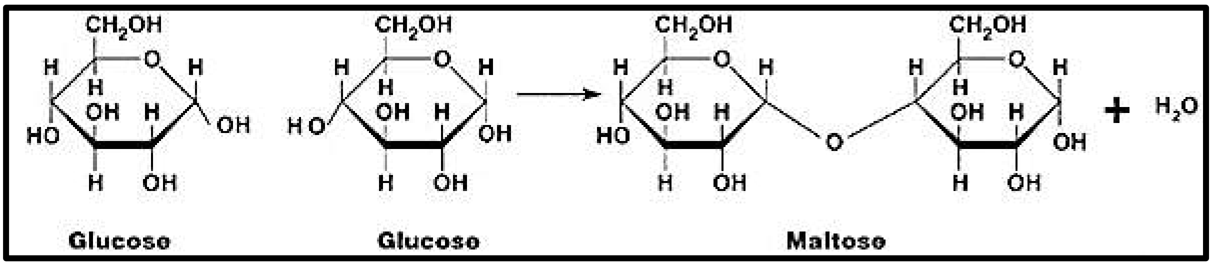
- Glucose reduces the blue copper sulphate solution in the Fehling's or Benedict's solution (containing Cu^{2+} ions) to insoluble reddish-brown copper(I) oxide, which is seen as a precipitate.
- Positive test is indicated by a green suspension and a red precipitate.

Condensation and Hydrolysis Reactions

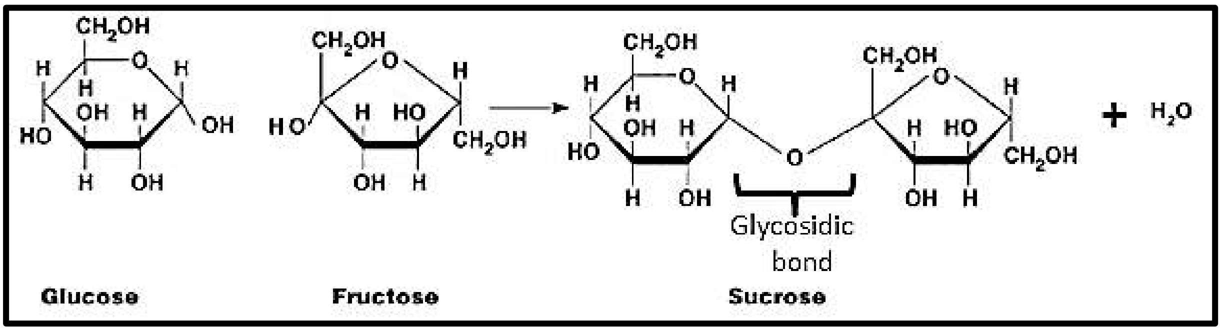
1. Condensation

- A chemical process by which two or more molecules are **joined** together to make a larger, more complex molecule with the **loss of a water molecule.**
- Examples include formation of maltose and sucrose.
- A **glycosidic bond** is a type of chemical linkage between the monosaccharide units of disaccharides and polysaccharides which is formed by the removal of a molecule of water.

Example 1: Formation of maltose



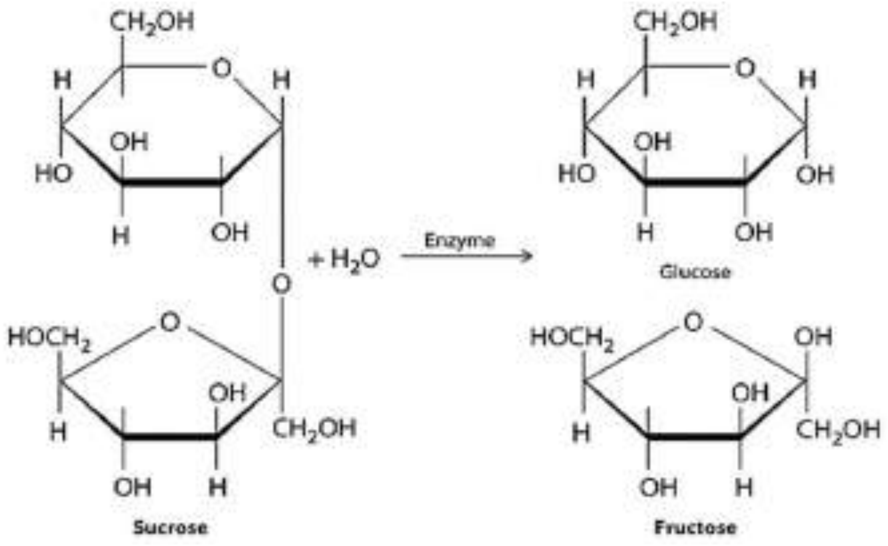
Example 2: Formation of sucrose



Hydrolysis

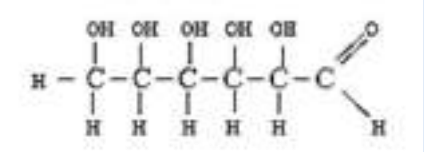
- A hydrolysis reaction uses enzymes or acids to separate a compound into its sub-units through the addition of a water molecule. For instance, disaccharides are hydrolyzed into two monosaccharides.

Example: Hydrolysis of sucrose




Exercise

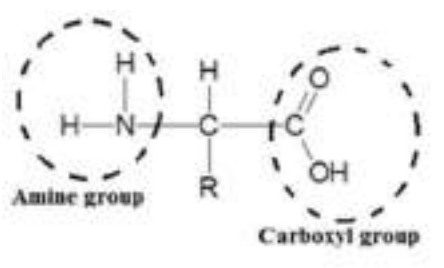
1. Explain the difference between monosaccharides, disaccharide and polysaccharides.
2. Explain the difference between condensation and hydrolysis reaction.
3. What is a glycosidic bond?
4. Polysaccharide and disaccharides are two forms of carbohydrates.
   - i. Give an example of a polysaccharide.
   - ii. Name the polysaccharide that is used as a food reserve in plants.
   - iii. Name the two reagents that initiate hydrolysis in polysaccharides and disaccharides.
5. In the school laboratory, a student found two amber bottles, without labels. Upon discussion with some teachers, the student found out that one of the bottles definitely contained glucose solution. Briefly describe how you would carry out a test to determine which of the two bottles contained glucose solution.
6. Starch can be converted to glucose by reacting it with water in the presence of dilute hydrochloric acid.
   - i. What is the name of the type of chemical reaction involved?
   - ii. What is the function of the hydrochloric acid?
   - iii. How could you show that the final reaction mixture contains glucose?
7. The structure of glucose, a simple sugar is shown below.



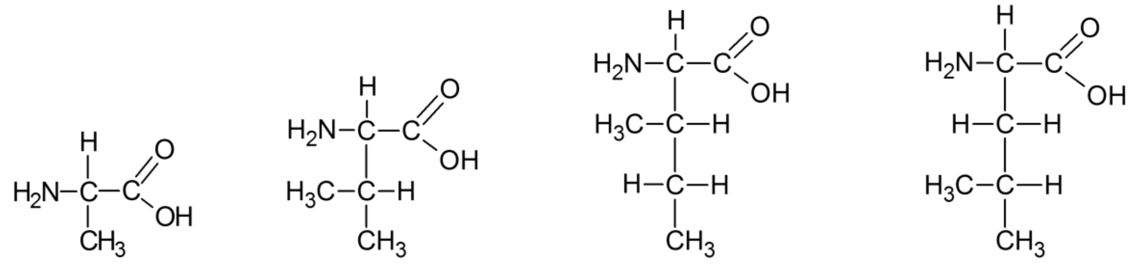
Write down the molecular formula of glucose and state one importance of glucose in the human body.

5.1.3 <u>Proteins</u>

- Proteins are polymers consisting of a large number of simple building units called amino acids.
- The amino acids are bifunctional molecules consisting of both an amine group (NH_2) and a carboxyl group (COOH). Due to its bi functionality, it is able to act as both an acid and as a base.
- Examples of amino acids include: alanine, valine, leucine, isoleucine, proline, phenylalanine, tryptophane, cysteine and methionine.



R = Side chain



Alanine Valine Isoleucine Leucine

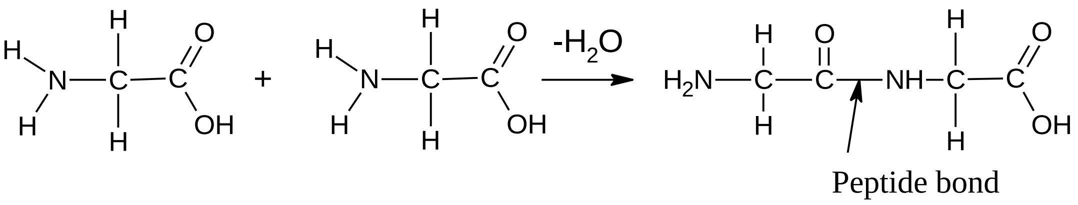
<u>Condensation and Hydrolysis reactions</u>

1. <u>Condensation</u>

- A condensation reaction occurs between the amino group ($-NH_2$) of one amino acid and the carboxyl group (-COOH) of another amino acid.
- This reaction forms a dipeptide that is held together by a peptide bond. During the process, there is loss of a water molecule.
- Multiple amino acids can be joined together by peptide bonds to form a polypeptide chain.

<u>Example</u>: Formation of a dipeptide

- A dipeptide is a molecule consisting of two amino acids joined by a single peptide bond.
- Dipeptides are produced from polypeptides by the hydrolysis process.
- Dietary proteins are digested to dipeptides and amino acids, and the dipeptides are absorbed more rapidly than the amino acids, because their uptake involves a separate mechanism.

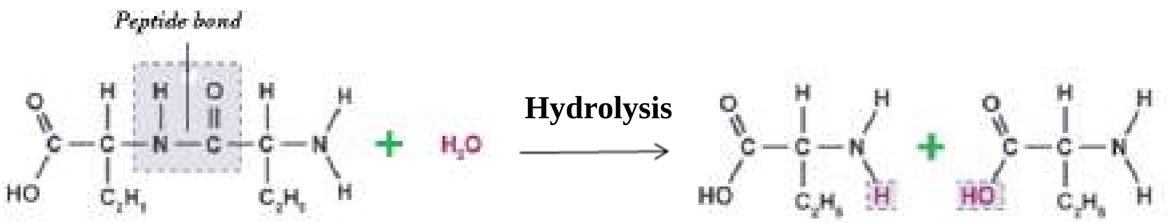


A dipeptide

2. **<u>Hydrolysis</u>**

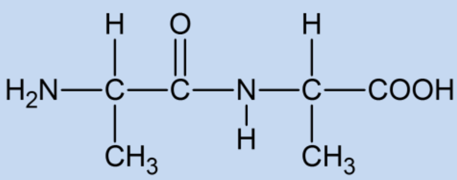
- In the presence of acids or enzymes, peptides can be broken down into individual amino acids via hydrolysis reactions.

<u>Example</u>



Exercise

1. Give some sources of proteins.
2. What is the general formula of amino acids?
3. What is a peptide bond? Give an illustration to support your answer.
4. Differentiate between a dipeptide and a polypeptide.
5. The product obtained when two molecules of the amino acid alanine are linked together is shown below.

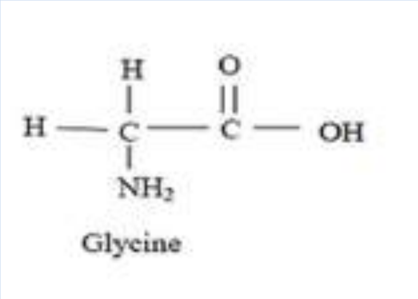


i. Name the other substance that is formed when the two alanine molecules combine.

ii. What structural feature does the above molecule have in common with a protein molecule?

iii. Draw the expanded structural formula of one molecule of alanine.

6. Consider the structure of the amino acid glycine drawn below.



(i) Draw the structure of the molecule that is formed when two glycine molecules combine.

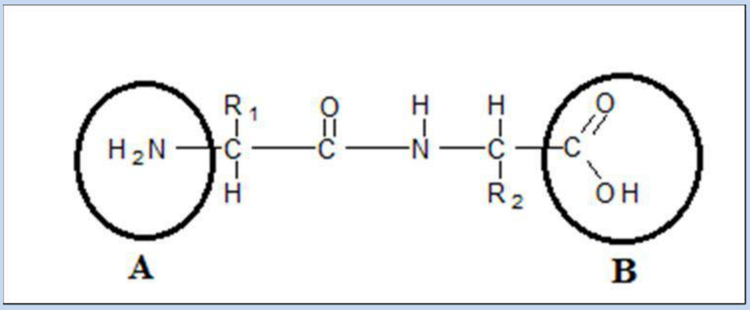
(ii) In the structure drawn in (i) above, circle and name the special bond that joins the two amino acids together.

(iii) What name is given to such reactions where two amino acids combine?

(iv) Name the other product that is formed during this reaction.

(v) What name is given to the process of breaking up the molecule that is formed from joining two glycine molecules?

7. The structural formula of a protein which is formed from two amino acids joined together is shown below.



i. Name the functional groups labeled A and B.

ii. Draw the structures of the two amino acids used in the above reaction.

8. Amino acids can act as both an acid and a base. Briefly account for this statement.

1.4 Added Chemicals in Processed Food

Processed Foods

- The term 'processed food' applies to any food that has been altered from its natural state in some way, either for safety reasons or convenience.
- Foods that have been processed may contain added salt, sugar, fat or other chemicals.

Some added chemicals in processed food are summarised in the table below.

<table>
<tr><th>Added chemical</th><th>Description/Purpose</th><th>Examples</th></tr>
<tr><td>Preservatives</td><td>

Are a group of chemical substances which are used in food production to slow down spoilage, discoloration, or contamination by bacteria and other disease forming organisms.
The main categories of artificial preservatives are antimicrobials, antioxidants and chelating agents.

<table>
<tr><th>Preservative</th><th>Description/Purpose</th></tr>
<tr><td>Antimicrobial</td><td>✓ Helps to prevent the overgrowth of bacteria and mold.</td></tr>
<tr><td>Antioxidants</td><td>✓ Helps prevent discoloration of food.</td></tr>
<tr><td>Chelating agents</td><td>✓ Helps to bind metals, usually copper and iron to prevent the metals from oxidizing and speeding up spoilage.</td></tr>
</table>
</td><td>
Antimicrobials
Benzoates (found in many beverages)

Sorbates (help to prevent mold, yeast and fungi growth in foods and beverages)

Proprionates (mold inhibitors used in baked goods)

Antioxidants

Sulphites, synthetic Vitamin E and C.

Chelating agents

Polyphosphates, citric acid.
</td></tr>
<tr><td>Sweeteners</td><td>
i. Added sugars - These are a synthetic form of natural sugars. Examples include corn syrup, fruit juice concentrate and molasses.
ii. Artificial sweeteners - These are a type of sugar substitute used instead of regular table sugar (sucrose). They are synthetic sugar substitutes, but may be derived from natural substances, including herbs or sugar itself. They are found in a variety of food and beverages marketed as "sugar-free" or "diet". Common food items containing artificial sweeteners include: soft drinks, chewing gum, jellies, baked goods, candy, fruit juice, and ice cream.
</td><td>
Examples of artificial sweeteners are:

Sulfame K, Aspartame, Neotame, Saccharin, Sucralose.
</td></tr>
</table>

Added Chemical	Description/Purpose	Examples
Food dyes	• Natural and artificial dyes and coloring give the food a more appealing look or makes the food tasty. • Mostly abbreviated as 'FD&C'.	Turmeric, saffron, chlorophyllin
Carbon dioxide	• A carbonated drink is a beverage that has dissolved carbon dioxide, most often to improve the taste and/or texture of the drink.	Cola, lemonade and other fizzy drinks
Monosodium glutamate (MSG)	• Monosodium glutamate is the sodium salt of glutamic acid, one of the most abundant naturally - occurring non-essential amino acids. • MSG added to foods acts as a flavor enhancer and provides a savory, broth-like or meaty taste.	Found in noodle flavours, savories and snacks

<u>Class Activity</u>

Look through the package of a cheesy snack given below and list some added chemicals in them. Determine the purpose of these chemicals in this particular snack.

Exercise

1. Which class of added chemicals does antimicrobials and antioxidants belong to? What is the specific purpose of adding such chemicals to food?
2. What gives the fizzy taste of cola drinks?
3. What does the food additive, MSG, stand for? Describe its chemical composition.
4. Explain the use of food dyes in processed foods.
5. Suggest the importance of descriptive labels for ingredients in food packaging.

5.2 Chemistry of Medicines and Drugs

Achievement Indicators

Upon completion of this sub-strand, students will be able to:

✓ Identify and describe some common medicines and drugs.

5.2.1 Chemicals in Drugs

- Drugs refer to any chemical substance that has an effect on the body- positive, negative, intentional, or side-effect.
- Medicines are legal drugs that are taken with the intention of healing or improving health conditions.
- Some types of legal drugs are described in the table below

Legal Drugs		
Type	**Description/Purpose**	**Example**
Analgesics	➢ An analgesic is a medicine that takes away physical pain. ➢ In simple terms, it is a pain reliever.	Paracetamol/ Acetaminophen and Aspirin
Antipyretics	➢ An antipyretic is a substance that reduces fever.	Aspirin and Ibuprofen
Antibiotics	➢ Antibiotics are medicines used to treat infections or diseases caused by bacteria.	Amoxicillin and Penicillin
Antacids	➢ An antacid is a type of inorganic medication that can control acid levels in the stomach.	Magnesium sulphates and bicarbonates

5.2.2 Chemicals in non-medicinal drugs

1. Cigarettes

- Cigarette and cigarette smoke contains a huge number and range of organic compounds.
- There are approximately 600 ingredients in cigarettes. When burnt, they create more than 7,000 chemicals.
- Chemicals in cigarette smoke enter the blood stream and affect the entire body and hence it leads to cancer, heart diseases and various lung diseases.
- Some chemicals in tobacco smoke include: Acetone, acetic acid, arsenic, ammonia, benzene, butane, cadmium, carbon monoxide, formaldehyde, hexamine, lead, naphthalene, methanol, nicotine, tar and toluene.

2. Alcoholic Beverages

- An alcoholic beverage is a drink which contains substantial amount of ethanol (alcohol).
- Since alcoholic beverages contain ethanol, the health effects of ethanol apply to alcohols as well.
- Examples of alcoholic beverages are summarized in the table below.

Name	Description	Percentage Ethanol
Beer	Produced by the saccharification (*hydrolysis of polysaccharides to soluble sugars*) of starch and fermentation of the resulting sugar.	2 – 12%
Wine	Wine is an alcoholic beverage made from fermented grapes or other fruits. Due to the natural chemical balance, grapes ferment without the addition of sugars, acids, enzymes, water, or other nutrients.	9 –16%
Spirits	A distilled beverage, spirit, liquor, or hard liquor is an alcoholic beverage produced by distillation of a mixture produced from alcoholic fermentation.	40 – 80%

3. **Kava**

- The tropical shrub *Piper methysticum* is widely cultivated in the South Pacific. The name "kava" refers to the plant or the beverage prepared from the plant.
- When dried, rootstock consists of approximately 43% starch, 20% fibers, 12 % water, 3.2 % sugars, 3.6 % proteins, 3.2 % minerals and 15 % kavalactones.
- Kavalactones are the active chemical ingredients of the kava root.
- The kavalactone component of kava can vary between 3% and 20 % of rootstock dry weight depending on the age of the plant and the cultivar.

Source: http://www.raw-devotion.com

4. **Inhalants**

- Non-medicinal inhalants are volatile substances that produce harmful chemical vapors that can be inhaled.
- Some common non-medicinal inhalants include: Aerosols, solvents (paint thinners, petrol, markers, paint removers, cleaning removers and cleaning fluids).

5.2.3 Some common illegal drugs

- Most illegal drugs are stimulants and/or depressants.
- Stimulants are drugs that induce temporary improvements in either mental or physical functions or both. Examples of these kinds of effects may include increased alertness and physical activity.
- Depressants are substances that slow down brain activity. Examples of short term effects of depressants include: poor concentration, confusion, fatigue, dizziness and depression.

- Examples of some illegal drugs are summarized below:

Name	Description
Marijuana	Marijuana refers to the dried leaves, flowers, stems, and seeds from the hemp plant, *Cannabis sativa*. The plant contains mind-altering chemicals. Marijuana acts as both stimulant and depressant, but it remains in body organs longer than alcohol.
Cocaine	Cocaine is a powerfully addictive stimulant, derived from coca or prepared synthetically and used as a stimulant.
Heroin	Heroin is an opioid drug that is synthesized from morphine, a naturally occurring substance extracted from the seed pod of the Asian opium poppy plant. Heroin is a depressant.
Amphetamines	Amphetamines are synthetic drugs that act as a stimulant.

Exercise

1. Give the purpose and examples of the following types of drugs.
   i. Antibiotic ii. Analgesic iii. Antipyretic iv. Antacid
2. Name some chemicals found in cigarettes.
3. What is the main component of any alcoholic beverage?
4. What are antacids?
5. Give the main components of dried rootstocks of kava.
6. Differentiate between a depressant and a stimulant.
7. Classify the following illegal drugs as a depressant, stimulant or both.

   i. Heroin ii. Marijuana iii. Cocaine iv. Amphetamines

5.3 HOUSEHOLD CHEMISTRY

Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Describe the purpose of some chemicals used in everyday household products.
- ✓ Describe the formation of soap and compare the cleaning action of soaps and detergents.

There are many different types of substances used in everyday life which has many different chemicals.

Some common substances are summarised in the table below.

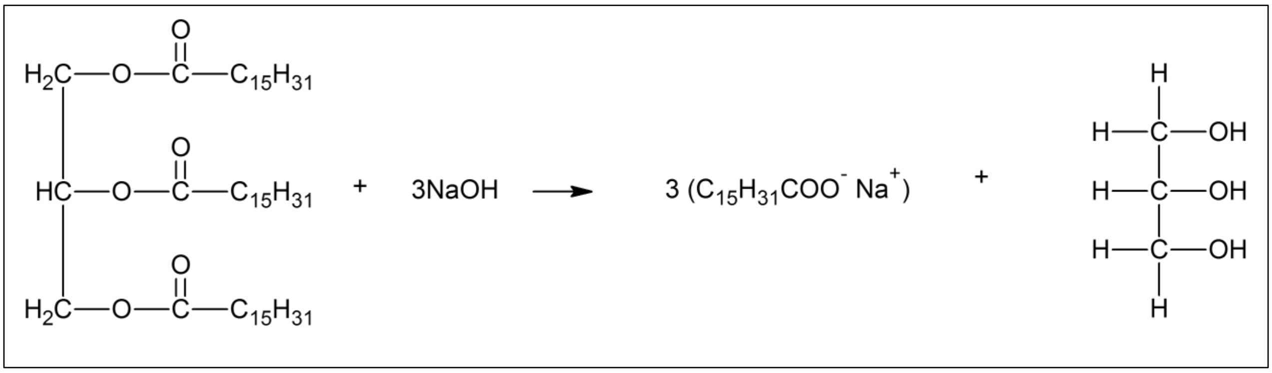
Name	Description	Example
Pesticides	Pesticide is a broad term, covering a range of harmful chemical based products that are used to control pests.	Some pesticides includes ant powder, insect killers (*insecticides*),mould and fungi killers (*fungicides*),weed killers (*herbicides*), slug pellets (*molluscicides*),plant growth regulators, bird and animal repellents, and rat and mouse killers (*rodenticides*). Common examples include paraquat and orthene.
Fertilizers	Fertilizers are substances used to add nutrients to the soil to promote soil fertility and increase plant growth. Natural fertilizers include: leaves, cow dung and bone meal compost which are used to make up the deficiency of nitrogen, phosphorous and potassium in soils. Artificial or chemical fertilizers are compounds that are manufactured to provide nitrogen, potassium and phosphorous to plants.	Urea, Ammonium sulphate, Potassium sulphate, NPK  *Source: http://uptorg.com*

Fuel	✚ Fuels are any materials that store potential energy in forms that can be practicably released and used for work or as heat energy. ✚ Fuels can be classified as distillate fuel or a residual fuel depending on the method of production. ✚ All fuel oils consist of complex mixtures of hydrocarbons.	✚ Benzene, Diesel, paraffin's, kerosene  *Source:http://www.slideshare.net*
Batteries	✚ A battery is a container consisting of one or more cells, in which chemical energy is converted into electricity and used as a source of power. ✚ Precaution when using batteries applies when touching damaged cells and when handling lead acid systems that have access to lead and sulfuric acid. ✚ Lead is a toxic metal that can enter the body by inhalation of lead dust or ingestion. ✚ If leaked onto the ground, the acid and lead particles contaminate the soil.	 *Source:https://learn.sparkfun.com*
Cleaners/ Cleaning Agents	✚ They are substances (usually liquids, powders, sprays or granules) which are used to remove dirt from surfaces. ✚ Cleaning agents can be acidic, alkaline or neutral depending on the use. ✚ Acidic cleaning agents- used for the removal of inorganic deposits like scaling. Cleaners containing HCl and vinegar are common examples. ✚ Alkaline cleaning agents can dissolve fats, oils and protein- based substances. Bleaches and ammonia are common examples.	Soap, detergent,solvent cleaners  *Source: http://www.harelmallac-export.com/*

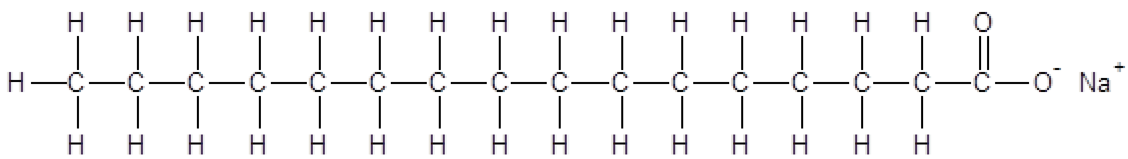
Formation of soap

- To understand what is needed to achieve effective cleaning, it is helpful to have a basic knowledge of soap and detergent chemistry.
- In the presence of alkali, fats can be hydrolyzed to form soap. This process is known as saponification.
- This process is the reverse of esterification.

Example



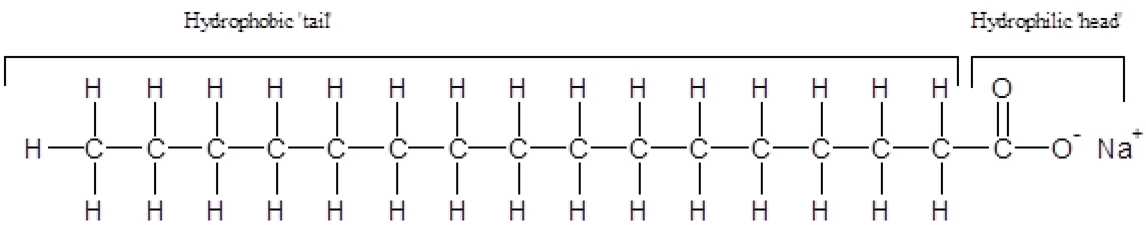
Note: The expanded structural formula of $C_{15}H_{31}COO^{-}\ Na^{+}$ is:



Cleaning action of soap and detergent

- Mostly, the dirt present on clothes are organic in nature and non-polar thus insoluble in water. Therefore, cleaning with water alone has little effect when stains consist of non-polar substances such as grease and oil.
- Soap consists of carboxylate ion (hydrophilic = soluble in water) and long hydrocarbon tail (hydrophobic = soluble in grease/oils).

Example



- Likewise detergent consists of sulphate ion (hydrophilic = soluble in water) and long hydrocarbon tail (hydrophobic = soluble in grease/oils).
- When soap or detergent is dissolved in water, the hydrophilic part dissolves in water and the hydrophobic ends attach themselves to the dirt (including grease and oils) and remove it from the cloth.
- Then the molecules of soap/detergent arrange themselves in micelle formation and trap the dirt at the center of the cluster.
- Scrubbing helps to pull the dirt and break them into small droplets.
- These droplets do not coagulate and redeposit on the surface of the cloth because of the repulsion between negative charges on the surface.
- The droplets are suspended in water .Foam is produced to help float the droplets of dirt.
- Rinsing helps to remove these droplets.

Summary: Soap functions because their long non-polar hydrophobic 'tail' dissolves is oil and grease to form a micelle and the micelle is surrounded by polar hydrophilic 'head 'of the soap molecules which make it soluble in water.

- Soap do not work well in hard water as the calcium ions (Ca^{2+}) and Magnesium ions (Mg^{2+}) causes the precipitation of the insoluble calcium and magnesium salts of the fatty acids known as 'scum'. This prevents formation of lather.

Example

$$2\,H_3C-(CH_2)_{16}-\overset{\overset{\displaystyle O}{\|}}{C}-O-Na \;+\; Mg^{2+} \longrightarrow (C_{17}H_{35}COO)_2Mg \;+\; 2Na^{+}$$

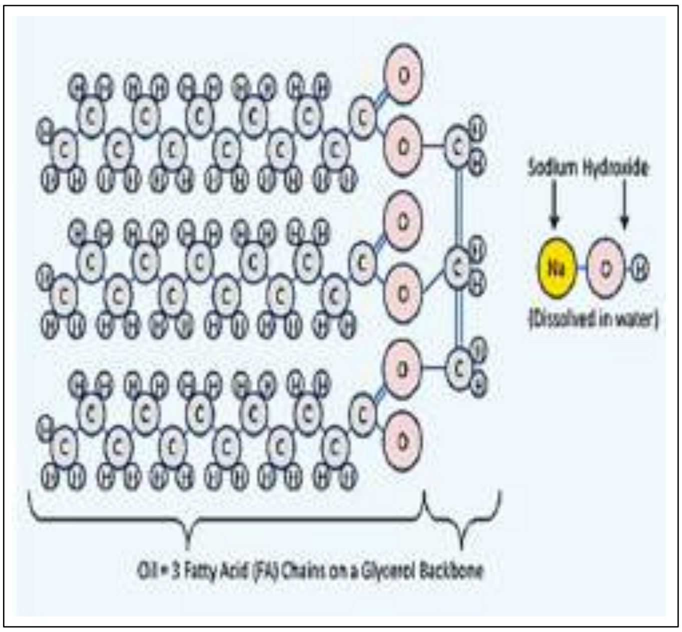
Soap White precipitate

- Detergents are not affected by hard water because the magnesium and calcium salts which forms from the reaction of detergents and magnesium and calcium ions are soluble so the cleaning action of detergents are not affected.

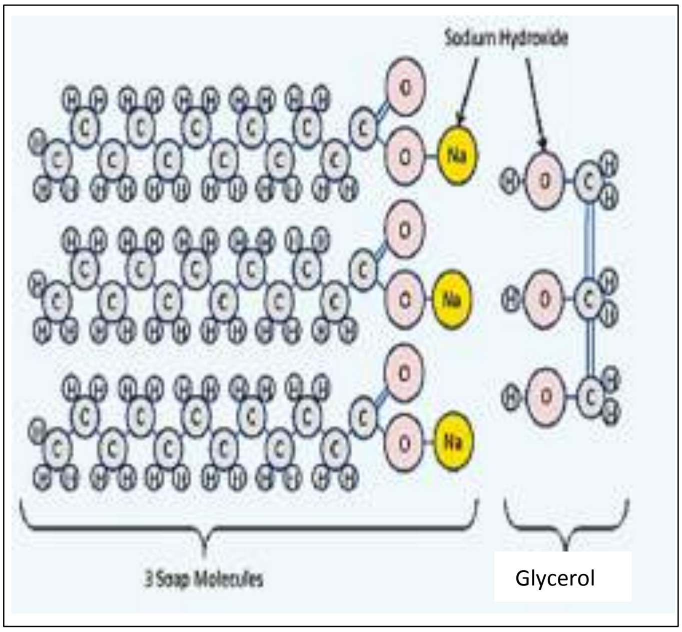
Important:

1. Never mix bleach or any bleach-containing product with any other cleaner containing ammonia**.** The gases created from this combination can lead to chronic breathing problems and even death.
2. Read all labels before buying and follow instructions when using cleaning products.
3. As a safer cleaning alternative, warm water and soap will often do the trick, especially at home. Baking soda is good for scrubbing. A mix of vinegar and water can clean glass.
4. When using cleaning or household products, keep the area well ventilated. Open windows and doors. Never use cleaning products in a small, enclosed space. Exposure can cause irritation in the eyes, mouth, lungs and on skin.

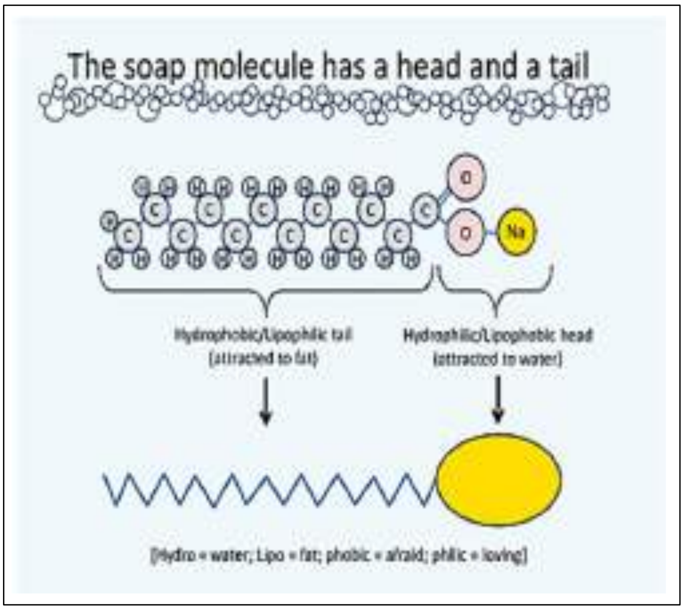
Summary of Soap Formation and its Action



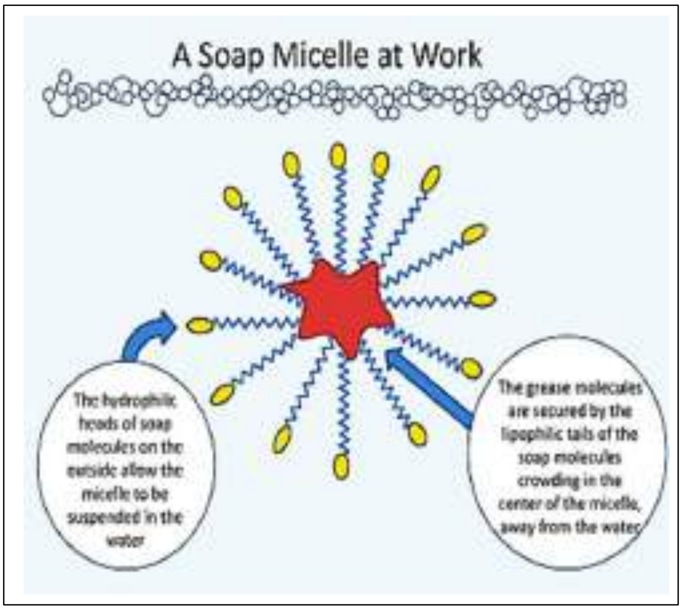
Ingredients for making soap:
Oil (Triglyceride) and NaOH



NaOH breaks down oil into soap and glycerol



The soap molecule has a head and a tail.



The long non-polar hydrophobic 'tail' dissolves is oil and grease to for a micelle and the micelle is surrounded by polar hydrophilic 'head 'of the soap molecules which make it soluble in water.

Image source: http://sustainablescientist.net

Exercise

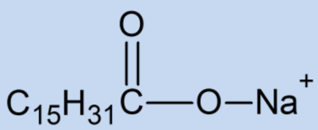
1. Give the purpose of the following:
   i. Pesticides
   ii. Fertilizers
2. Ammonium sulphate is a common soil fertilizer.
   i. What is the formula of ammonium sulphate?
   ii. What nutrient does this fertilizer supply to the soil?
   iii. Ammonium sulphate increases soil acidity. Name a common agricultural chemical which can reduce soil acidity.
3. Give some precautions that must be taken while using/discarding batteries.
4. What are fuels used for and give some examples of fuels.

5. i. Complete the word equation for making of soap given below.

   Glyceryl tristearate + ------------------ → Soap + ------------------

   ii. If the molecular formula of the soap formed in the above reaction is $C_{17}H_{35}COONa$, predict the formula of the glyceryl tristearate.

5. Briefly explain what is observed when soap is used in hard water. Write an ionic equation for the reaction.
6. Label the hydrophilic and the hydrophobic group of a soap molecule shown below.



7. Briefly explain why cleaning with water alone has little effect when clothes has grease or oil stain.
8. Describe the cleaning actions of soaps.
9. Briefly explain why soap cannot lather in hard water.
10. Most bottles of bleach have the following information printed on them:
    i. Use in well ventilated areas.
    ii. Must be diluted before adding to clothes.

    Give a reason for the above precautions.

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Glossary

A **Accuracy-** Refers to the closeness of a measured value to a standard, known or true value.

Acid - Is a compound which can provide protons (H^+) in aqueous solutions. It is also an electron pair acceptor and neutralizes a base to form salt and water.

Acidic oxides - Are the oxides of non-metals. They can also be termed as inorganic chemicals that react with water to form an acid; or react with a base to form a salt.

Acid-base indicator- A dye (mostly an organic weak acid) which has one distinct color in acid and another distinct color in base. It is used in acid – base titration.

Activation energy- Is the minimum kinetic energy that must be possessed by the reactants in order to start a chemical reaction.

Addition reaction- A reaction in which reactants are added to an unsaturated compound containing a carbon to carbon double or triple bond.

Alcohol- Is an organic compound which contains the hydroxyl group, (OH), attached to an alkane framework.

Aliquot- A portion of a chemical substance. The term is most commonly used in titration to refer to the volume of liquid delivered by the pipette.

Alkane – A hydrocarbon belonging to the series with the general formula C_nH_{2n+2}.Contains C-C single bonds.

Alkene- A hydrocarbon belonging to the series with the general formula C_nH_{2n}.Contains C=C double bonds.

Alkyl group – A group of atoms derived from the alkane series with the general formula C_nH_{2n+1}.

Alkyne - A hydrocarbon belonging to the series with the general formula C_nH_{2n-2}.Contains C≡C triple bonds.

Amphoteric - Are substances which are able to react as both an acid and a base. These are particles capable of being a proton donor and a proton acceptor.

Analgesics- An analgesic is a medicine that takes away physical pain.

Anode- The electrode where oxidation occurs in an electrochemical change. It is the electrode used in electrolytic cells that is connected to the positive terminal of the power source.

Antacids - An antacid is a type of medication that can control acid levels in the stomach or neutralizes stomach acidity.

Antibiotics- Antibiotics are medicines used to treat infections or diseases caused by bacteria.

Antipyretics- An antipyretic is a substance that reduces fever.

Avogadro's number- Is the number of elementary particles or formula units (which can be molecules, atoms, compounds or ions) in one mole of a substance. The number is $6.22 \times 10^{23.}$

B **Base -** Is a chemical compound that releases hydroxide ions (OH^-) in aqueous solutions. It is also a proton acceptor, an electron pair donor and neutralizes an acid to form salt and water.

Basic oxides -Are the oxides of metals. They react with water to form a base or reacts with an acid to form a salt and water.

Bond angle - The angle formed by two bonds that extends from the same atom.

Bond pair -A pair of electrons forming a covalent bond.

Bond polarity- Is a measure of how equally the electrons in a bond are shared between the two atoms of the bond.

C

Catalyst- Is a substance that increases the rate of chemical reaction by providing an alternative pathway for the reaction (pathway that has a lower activation energy) without being used up or without being chemically changed in the process.

Carboxylic acid- An organic compound containing the carboxyl group (- COOH).

Cathode – The electrode where reduction occurs. It is also the electrode used in electrolytic cells that is the negative terminal of the power source.

Cis fats- In *cis* configuration, the carbon chain extends from the same side of the double bond.

Coefficients- Numbers in front of formulas in chemical equations.

Collision theory – States that the rate of a reaction is proportional to the number of effective collisions that occur each second between the reactants.

Combustion- A chemical reaction with oxygen in which heat and light energy are released.

Concentration – Is a measure of the amount of solute in a fixed quantity of solvent.

Condensation reaction- Is a chemical process by which two or more molecules are joined together to make a larger, more complex, molecule, with the loss of water.

Conjugate acid- The species in a conjugate acid-base pair that has the greater number of H^+ units.

Conjugate acid-base pair- Two substances (ions or molecules) whose formulas differ by only one H^+ unit.

Conjugate base - The species in a conjugate acid-base pair that has the fewer number of H^+ units.

Constitutional isomerism (or structural isomerism)-A type of isomerism where compounds have the same number and type of atoms arranged in different ways.

Covalent bond- A chemical bond which results when two atoms share electron pairs.

D

Dilution- Is the process of making a concentrated solution less concentrated.

Dilute solution- A solution in which the ratio of the quantities of solute to solvent is small.

Dimensional Analysis- A problem solving technique that uses the correct cancellation of the units of physical quantities to get the correct set up of the solution to the problem.

Dipeptide - A dipeptide is a molecule consisting of two amino acids joined by a single peptide bond

Diprotic acid- An acid that has two H^+ per molecule.

Disaccharide- A carbohydrate whose molecules can be hydrolyzed to two monosaccharides.

Dissociation- Is a term used to explain the splitting of molecules into its simpler units or ions.

Drug - Any substance that has an effect on the body — positive, negative, intentional, or side-effect.

Dynamic equilibrium- An equilibrium where the rate of the forward reaction is equal to the rate of the backward reaction.

E

Electron configuration- Shows the arrangement or distribution of electrons in an atom's orbital.

Electron group geometry- Shows the arrangement of electron groups about the central atom of a molecule.

Electronegativity - Is a measure of the strength of an atom to attract the shared pair of electrons in forming a covalent bond.

Empirical formula - Expresses the simplest whole number ratio of atoms or ions in a compound.

Endothermic reactions- Are reactions which absorbs heat energy from the surrounding.

End point- Is the point during a titration experiment at which the indicator changes color, signifying the end of the reaction.

Enthalpy content - Is the heat energy content of a substance.

Enthalpy change- Is the heat energy content change occurring during a chemical reaction.

Enthalpy of Reaction- Is the change in the heat energy (enthalpy) of a chemical reaction that occurs at a constant pressure.

Esterification - Is the process of producing an ester from carboxylic acid and an alcohol.

Equivalence point- The point in a titration where the reactants have reacted according to their stoichiometric equation.

Exact numbers – Are numbers whose values are known exactly.

Exothermic reactions – Are reactions which releases heat energy to the surrounding.

F

Fat: An ester formed from glycerol and fatty acids.

Fatty acids - Are compounds that have a long hydrocarbon (carbon and hydrogen) chain with a carboxyl (acid) group.

Functional group- Is an atom or group of atoms within a molecule that shows a characteristic set of predictable physical and chemical behaviors.

G

Geometric isomer- A compound with the same molecular formula as another compound and having the same structural arrangement; however the spatial arrangement is different which is usually related to the presence of a double bond in the molecule.

Glycosidic bond - Is the type of chemical linkage between the monosaccharide units of disaccharides and polysaccharides which is formed by the removal of a molecule of water.

Gravimetric Analysis- Is a method of quantitative analysis where the constituent needed to be analyzed is converted into a substance of known composition that can be separated from the sample and reweighed.

H

Haber process – Is a process in which ammonia is manufactured by combining nitrogen and hydrogen.

Half-equation- An equation indicating either a reducing or an oxidising process.

Halogenation- A reaction involving a halogen which either substitutes into or adds to a molecule.

Heat of Reaction - *See enthalpy of reaction.*

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Hydration- Is a chemical process in which water is added to a compound to form a hydrated compound.

Hydrocarbons- Are large family of organic compounds containing only carbon and hydrogen atoms.

Hydrogenation- A reaction which adds hydrogen to a compound.

Hydrolysis- A hydrolysis reaction separates a compound into its constituents through the addition of a water molecule.

Hydrophilic group - A polar molecular unit which is capable of having attractions or hydrogen bonds with water molecules.

Hydrophobic group- A non-polar molecular unit with no affinity or attraction for the molecules of a polar solvent such as water.

Hydronium ion- A hydrated hydrogen ion (proton), H_3O^+.

I

Inexact numbers – Are those numbers whose values have some uncertainty.

Indicator - A chemical put in a solution being titrated and whose change in color signals the end point.

Inhalants- Inhalants are volatile substances that produce chemical vapors that can be inhaled.

Intermolecular forces (bond) – Force of attraction between two molecules.

Intramolecular forces (bond) - Forces that hold the atoms together in a molecule, such as the covalent and ionic bonds. Usually referred to as a bond within a molecule.

Ionic bond – Is a bond which results when an atom with far greater electronegativity combines with an atom of lower electronegativity, and the result is the complete transfer of valence electrons to the higher electronegative atom.

Ionic product- The mathematical product of the molar concentrations of ions of a substance.

K **K_c** – Is the equilibrium constant for a reaction involving the concentrations of the reagents in mol L^{-}.

K_w- The equilibrium constant for the dissociation of water. It has the approximate value of 1 x 10^{-14} mol^2L^{-2}.

L **Le Chatelier's principle-** A principle which states that if a change is applied to a system at dynamic equilibrium, the position of the equilibrium shifts to counteract the change and reestablish equilibrium.

Lewis structures - Are visual representations of the bonds (or electrons) between atoms and illustrate the lone pairs of electrons in molecules.

Lipid – Any class of organic compounds that are fatty acids or their derivatives and are insoluble in water but soluble in organic solvents.

M **Medicine-** A substance that are taken with the intention of healing or improving health conditions.

Molar mass- Is the mass of one mole of any chemical compound.

Molecular formula - Expresses the actual number of atoms of each type in the compound.

Molecular geometry (shape) - Shows the arrangement of the atoms in space.

Monoprotic acid- An acid which has only one ionisable proton or hydrogen ion.

Monosaccharide- A simplest form of carbohydrate that cannot be hydrolyzed.

Monounsaturated fat- A fat that have one carbon-carbon double bond in its structure.

N **Non-polar covalent bonds-** A type of bond which forms when two atoms of equal electronegativity are bonded where both will have the same tendency to attract the shared pair of electrons.

O **Octet rule-** An atom tends to gain or lose electrons until its outer shell has eight electrons.

Oxidation- Is a process in which there is a gain of oxygen or loss of hydrogen or loss of electrons or simply an increase in oxidation number of the atom, ion or molecule.

Oxidation state - Is a number that is assigned to an element in a chemical reaction to show the total number of electrons which have been removed from an element (a positive oxidation state) or added to an element (a negative oxidation state) to get to its present state.

Oxidation number- *See oxidation state*.

Oxide- A chemical compounds with one or more oxygen atoms combined with another element.

P **Peptide bond-** a peptide bond is a chemical bond formed between two molecules when the carboxyl group of one molecule reacts with the amino group of the other molecule, releasing a molecule of water (H_2O).

Periodic table- Table containing the symbol of the chemical elements arranged in order of increasing Atomic number and arranged so that the elements with similar properties lie in the same column.

pH- Is a measure of how acidic or basic a substance is.

Polar covalent bond- A bond which forms when atoms with high electronegativity are covalently bonded to atoms of lower electronegativity.

Polypeptide- Is chain of many amino acids linked together by peptide bonds.

Polyprotic acids-Acids which have more than one ionisable protons or hydrogen ions.

Polyunsaturated fat- Fat which have more than one carbon–carbon double bond in their structure.

Polysaccharides- Are long chains of monosaccharide units bound together by glycosidic linkages.

Precision- Refers to the closeness of two or more measurements to each other.

Preservatives- are a group of chemical substances which are used in food production to slow down spoilage, discoloration, or contamination by bacteria and other disease forming organisms or pathogens.

Primary standard - Is a salt or compound that is used to prepare a standard solution.

Proteins - Are polymers consisting of a large number of simple building units called amino acids.

Q

Quantitative analysis - Is the determination of the amount of a substance present in a given sample.

R

Reduction- Is a process in which there is a loss of oxygen or gain of hydrogen or gain of electrons or simply a decrease in oxidation number of the atom, ion or molecule.

REDOX- Abbreviated term for oxidation and reduction reactions.

Refining of fats- Is a process of removing impurities from extracted fats and oils.

S

Saturated fats- These fats contain only single carbon-carbon double bonds in their structure.

Smoke point- Is the temperature at which the oil is decomposed.

Stoichiometry - Is the quantitative relationship among reactants and products.

Standard solution- Is a solution whose concentration is accurately known.

T

Titration- Is a technique in which a solution of known concentration is used to determine the unknown concentration of another solution

Titre- Is the volume of liquid delivered from a burette during a titration experiment.

Trans fat- In *trans* configuration, the carbon chain extends from opposite sides of the double bond.

Triglycerides- Are sets of three fatty acids bound together.

U

Unsaturated fat- A fat that contain at least one carbon-carbon double bonds in its structure.

W

Water of crystallization - Is the number of water molecules, chemically combined in a definite molecular proportion with the salt, in its crystalline state.

Appendix

PERIODIC TABLE OF THE ELEMENTS

Row	Group I	II											III	IV	V	VI	VII	VIII
1	1 H 1																	2 He 4
2	3 Li 7	4 Be 9											5 B 11	6 C 12	7 N 14	8 O 16	9 F 19	10 Ne 20
3	11 Na 23	12 Mg 24											13 Al 27	14 Si 28	15 P 31	16 S 32	17 Cl 35.5	18 Ar 40
4	19 K 39	20 Ca 40	21 Sc 45	22 Ti 48	23 V 51	24 Cr 52	25 Mn 55	26 Fe 56	27 Co 59	28 Ni 59	29 Cu 64	30 Zn 65	31 Ga 70	32 Ge 73	33 As 75	34 Se 79	35 Br 80	36 Kr 84
5	37 Rb 85	38 Sr 88	39 Y 89	40 Zr 91	41 Nb 93	42 Mo 96	43 Tc (98)	44 Ru 101	45 Rh 101	46 Pd 106	47 Ag 108	48 Cd 112	49 In 115	50 Sn 119	51 Sb 122	52 Te 128	53 I 127	54 Xe 131
6	55 Cs 133	56 Ba 137	57* La 139	72 Hf 178	73 Ta 181	74 W 184	75 Re 186	76 Os 190	77 Ir 192	78 Pt 195	79 Au 197	80 Hg 201	81 Tl 204	82 Pb 207	83 Bi 209	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra 226	89+ Ac 227	104 Ku 260														

* Lanthanide Series	58 Ce 140	59 Pr 141	60 Nd 144	61 Pm (145)	62 Sm 150	63 Eu 152	64 Gd 157	65 Tb 159	66 Dy 162	67 Ho 165	68 Er 167	69 Tm 169	70 Yb 173	71 Lu 175
+ Actinide Series	90 Th 232	91 Pa 231	92 U 238	93 Np 237	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (252)	99 Es (254)	100 Fm (257)	101 Md (256)	102 No (259)	103 Lr (260)

( ) mass number of most stable isotope.

Key :

Z	Atomic number
X	symbol of element
A	Relative atomic mass