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Unit 1: Gas Laws

Introduction

In Form 1 under the Kinetic theory of matter, we learnt that substances such as solids, liquids and gases are made up of small particles and that they occupy space and have volume. We also learnt that these particles possess Kinetic energy and are in constant random motion.

In gases, for example, the molecules in the vessel move randomly at high speed colliding with one another and with the walls of the container. We learnt that this leads to some force being exerted on the walls of the container. This force that a gas exerts on the walls of the container per unit area is what is known as gas pressure. This pressure is constant for every unit area of the walls of the container and that it can either be raised or reduced by increasing or decreasing such factors like temperature, the amount of gas, or the volume of the container.

In this unit, we will learn how changing the above factors affect the behaviour of the gas in the container. This will be done through learning about important gas laws which include Boyle’s law, Charles’ law, Combined gas law and the Graham’s law of diffusion.

1.1: Boyle’s law

In the 1700’s a scientist called Robert Boyle investigated the relationship between the volume of a dry gas and its pressure. He did not change the mass of gas and its temperature during his investigations. He found out that when the pressure increased the volume decreased and when volume increased the pressure decreased. Robert Boyle expressed his results in the law named after him.

Boyle’s law states that the volume of a given mass of gas is inversely proportional to the pressure applied to it provided that the temperature remains constant.

When we say that the volume (V) is inversely proportional to the pressure (P); this means that when pressure increases, volume decreases; and when pressure decreases, volume increases.

To understand Boyle’s law, let us consider the following situation.

Let us compare the words “inversely proportional” with the following analogy:

- A number of people in a particular household.
- The period it takes them to finish completely ten packets of maize flour.
In this household, each person can finish 1 kg of maize flour per day. In the beginning of the month there are only two people (husband and wife) in the household. With time, nephews, cousins and other relatives join them. As the number of people living in the house increases, the time taken for the ten packets of maize flour in the household to be used up reduces. This means that the number of people in the household is inversely proportional to the time taken to finish all the packets of maize flour. See Fig. 1.2 below.

Fig. 1.2 An analogy explaining Boyle’s law

Using the above analogy, let us see how change in pressure affects the volume of a given gas.

According to Boyle’s law, the volume of a gas decreases as the pressure applied to it increases. If pressure applied decreases, the volume of the gas increases. This will happen only if the temperature and the mass of the gas remain constant.

Boyle’s law can be represented as:

\[ P \propto \frac{1}{V} \quad \text{or} \quad PV = \text{constant} \]

If a certain mass of gas has a volume \( V_1 \) at a pressure \( P_1 \) then: \( P_1V_1 = \text{constant} \). If the same gas has a volume \( V_2 \) at pressure \( P_2 \) then: \( P_2V_2 = \text{a constant} \).

Combining the two expressions, we get \( P_1V_1 = P_2V_2 = \text{constant} \). Therefore \( P_1V_1 = P_2V_2 \)

Where
- \( P_1 \) represents initial pressure.
- \( V_1 \) represents initial volume.
- \( P_2 \) represents final pressure.
- \( V_2 \) represents final volume.
A plot of the graph of pressure (P) against volume (V) gives a curve similar to the one shown below.

![Graph of pressure against volume](image)

*Fig. 1.3 A graph of pressure against volume*

The graph shows that as the volume increases, the pressure of the gas decreases.

**Note:** Both pressure and volume never reach zero hence the curve does not touch both the x and y axes.

Now, let us carry out the following activity to demonstrate how pressure increases as volume decreases.

**Activity 1.1**

1. In your notebook, draw a graph of Boyle’s law as shown below.

![Graph of Boyle’s law](image)

2. At any point on the vertical axis (y-axis), in the above graph draw a horizontal line until it touches the curve. From this point draw a vertical line until it touches the horizontal axis (x-axis) as shown below.
3. Label $P_1$ and $V_1$ respectively.

4. Repeat the above procedure at a different position of pressure as shown below.

From the graphs in 4 above

I. What happens to the volume of the gas if:
   (a) pressure is increased?
   (b) pressure is decreased?
II. What happens to the pressure of the gas if:
(a) volume is increased?
(b) volume is decreased?

From the above activity we can see that when the volume of a gas is increased, the pressure decreases and if the volume is decreased the pressure increases. Therefore, the volume of a gas is inversely proportional to the pressure applied on the gas and vice versa.

However if a graph of pressure (P) against \(\frac{1}{\text{volume}}\) is plotted, the graph will be a straight line instead of a curve as shown in Fig. 1.4 below. How would you expect a graph of volume against \(\frac{1}{\text{pressure}}\) to look like? Sketch the graph in your notebook.

![Graph of pressure against the reciprocal of volume](image)

*Fig. 1.4 A graph of pressure against the reciprocal of volume*

Your graph of V against \(\frac{1}{P}\) should look like the one in Fig. 1.4 above. This is because if pressure is doubled the volume is halved and vice versa. Diagrammatically, Boyle’s law can be represented as shown in Fig. 1.5.

![Diagrammatical representation of Boyle’s law](image)

*Fig. 1.5 Diagrammatical representation of Boyle’s law*
• What do you observe in Fig. 1.5?
• What do you conclude from your observation?

When the pressure applied to a gas is high, the volume of the gas is small. As the pressure decreases, the volume of the gas increases. Therefore, pressure is inversely proportional to volume at a constant temperature. Note from Fig. 1.5 that a small \( v \) corresponds to a big \( P \) and a big \( V \) corresponds to a small \( p \). This means that \( Pv = PV \)

Hence \( P_1V_1 = P_2V_2 \).

**Units of pressure**

The standard unit of pressure is the **Pascal**. One pascal is equivalent to a force of one newton exerted on an area of one square metre.

\[
1 \text{ Pascal (Pa)} = 1 \text{ N/m}^2
\]

Pressure is also measured in a unit called **standard atmosphere** or **atm**.

1 atm = 101 325 N/m\(^2\). Another unit that is used for pressure is the millimetre of mercury (mm Hg)

1 atm = 760 mm Hg

### 1.2: Applications of Boyle’s law

**Example 1**

- If 60 cm\(^3\) of oxygen is compressed from 20 atmospheres pressure to 40 atmospheres pressure, what is the new volume of the gas at constant temperature?

**Solution**

Step 1: Identify the variable conditions

- Initial pressure, \( P_1 = 20 \text{ atms} \)
- Initial volume, \( V_1 = 60 \text{ cm}^3 \)
- Final pressure, \( P_2 = 40 \text{ atms} \)
- Final volume, \( V_2 = ? \)

Step 2: Make unknown variable the subject of the formula.

Since \( P_1V_1 = P_2V_2 \); then \( V_2 = \frac{P_1V_1}{P_2} \)

Step 3: Solve by substituting the known variables i.e

\[
V_2 = \frac{1}{40 \text{ atm}} \times \frac{30}{60 \text{ cm}^3} = 30 \text{ cm}^3
\]
The new volume \( (V_2) \) at constant temperature is 30cm\(^3\).

**Note:** Remember always to include the units when solving these problems.

**Example 2**
- If a sample of a gas in a balloon has a volume of 200cm\(^3\) and a pressure of 6 atms is compressed and its new pressure increased to 20 atms, what would be its new volume? Assume the temperature remains constant.

Identify the variables

\[
P_1 = 6 \text{ atms} \\
V_1 = 200 \text{ cm}^3 \\
P_2 = 20 \text{ atms} \\
V_2 = ?
\]

\[
P_1V_1 = P_2V_2
\]

Substituting the known variables:

\[
V_2 = \frac{6 \text{ atms} \times 200 \text{ cm}^3}{20 \text{ atms}} = 60 \text{ cm}^3
\]

Its new volume would be 60 cm\(^3\).

**Example 3**
- A certain mass of a gas occupies 500 cm\(^3\) at 100,000 Pa. What will be its pressure at 1000 cm\(^3\) volume?

**Solution**

\[
P_1 = 100,000 \text{ Pa} \\
V_1 = 500 \text{ cm}^3 \\
P_2 = ? \\
V_2 = 1000 \text{ cm}^3 \\
P_1V_1 = P_2V_2
\]
\[
P_2 = \frac{P_1V_1}{V_2}
\]
\[
= \frac{100,000 \text{ Pa} \times 500 \text{ cm}^3}{1000 \text{ cm}^3}
\]
\[
= 50,000 \text{ Pa}
\]

Its pressure will be 50,000 Pa.

**Exercise 1.1**

1. State Boyle’s law
2. How is Boyle’s law represented mathematically?
3. If a gas has a volume of 50 cm\(^3\) at a certain pressure, what would be its new volume if the pressure is doubled?
4. If the pressure of a gas is 3 atmospheres and the volume is increased four times. What would be its final pressure?
5. At constant temperature, a gas with a pressure of 750 mm Hg occupies a volume of 80 cm\(^3\). What will be the pressure of the gas if it is expanded to occupy a volume of 340 cm\(^3\).
6. Complete the following table:

<table>
<thead>
<tr>
<th></th>
<th>Initial Pressure</th>
<th>Initial volume</th>
<th>Final pressure</th>
<th>Final volume</th>
</tr>
</thead>
<tbody>
<tr>
<td>(i)</td>
<td>760 mmHg</td>
<td>300 cm(^3)</td>
<td>—</td>
<td>5000 cm(^3)</td>
</tr>
<tr>
<td>(ii)</td>
<td>—</td>
<td>481 cm(^3)</td>
<td>2 atm</td>
<td>281 cm(^3)</td>
</tr>
<tr>
<td>(iii)</td>
<td>1 atm</td>
<td>— dm(^3)</td>
<td>5 atm</td>
<td>80 dm(^3)</td>
</tr>
<tr>
<td>(iv)</td>
<td>770 mmHg</td>
<td>25.5 dm(^3)</td>
<td>790 mmHg</td>
<td>—</td>
</tr>
<tr>
<td>(v)</td>
<td>760 mmHg</td>
<td>400 cm(^3)</td>
<td>—</td>
<td>8000 cm(^3)</td>
</tr>
</tbody>
</table>

**1.3: Charles’ law**

Another scientist, Jacques Charles, investigated the relationship between the volume of a gas and temperature change. He noted that the volume of a gas increases with increase in temperature. He expressed his results in a law named after him.

Charles’ law states that *at constant pressure, the volume of a fixed mass of a gas is directly proportional to its absolute temperature.* Absolute temperature is a temperature measured with respect to absolute zero on the Kelvin scale. It is the lowest temperature where the volume of a gas is assumed to be zero.
When we say that the volume (V) is directly proportional to temperature (T), it means that as T increases, V also increases and as T decreases, V also decreases. To understand Charles’ law better, let us consider the following analogy:

**Fig. 1.6: Jacques Charles**

Take the following two things into consideration:
- The number of chapattis, C
- The time(T) taken to finish eating a given number of chapattis.

Let us compare the term ‘directly proportional’ to the analogy.

In a competition where several teams are competing to eat chapattis, a member of the chapatti eating team employed by chapatti eating.com can eat four chapatis in five minutes. In a chapatti eating competition, how long would it take the competitor to eat the following number of chapattis?
- 16
- 32
- 8

We can observe that as the quantity of chapattis on the table increase, Fig. 1.7 the time taken by the competitor to eat the chapattis increases. If the number of chapattis on the table is decreased, the time the person takes to finish the chapattis will also decrease. This implies that the time(T) taken by the person to complete eating the chapattis is proportional to the number of chapattis, C, the person is assigned to eat. See Fig. 1.7.

**Fig. 1.7 An illustration to explain Charles’ law.**

Using the above analogy let us see how change in temperature affects volume in Charles’ law.
According to Charles’ law, the volume of the gas increases as its absolute temperature increases. However, if its absolute temperature is decreased, its volume will consequently decrease. We say that the volume of a gas is directly proportional to its absolute temperature. This can only happen if the pressure of the gas and its mass remain constant. Diagrammatically, Charles’ law can be represented as illustrated in Fig. 1.8.

**Fig. 1.8 A diagramatic representation of Charles’ law**

- What do you observe?
- What can you conclude from this?

Fig. 1.8 shows that when $v$ is small, $t$ is also small. As $V$ increases, $T$ also increases. Therefore when the volume increases, temperature also increases.

Charles’ law is expressed as

$$V \alpha T \text{ or } \frac{V}{T} = \text{constant}$$

Therefore if a fixed mass of a gas with volume $(V_1)$ and temperature $(T_1)$ has its temperature changed to $(T_2)$ at constant pressure, then its volume will change to $(V_2)$. This can be represented as follows:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Where
- $V_1$ represents initial volume
- $T_1$ represents initial temperature.
- $V_2$ represents final volume
- $T_2$ represents final temperature
A plot of the graph of volume against absolute temperature gives a straight line as shown in Fig. 1.9.

Note that the graph is a straight line with a positive gradient and does not pass through the origin (zero).

The graph implies that a fixed mass of a gas expands or contracts by \( \frac{1}{237} \) of its volume at 0°C for each degree Celsius rise or fall in temperature. This graph also shows that the volume of a gas at –273°C would become zero if the gas does not liquify or solidify as happens to gases at low temperature.

Gases liquify or solidify at low temperatures to liquids or solids respectively. Gas volumes cannot therefore, be measured at temperatures right down to –273°C, but if the volume of a given mass of gas is measured at different temperatures, by extrapolation (extending the line backwards) we can show that the gas would have no volume if it could be obtained at –273°C.

This temperature, at which, theoretically, a gas would have no volume is known as **absolute zero** and it provides the starting point for a scale of temperature known as the **absolute scale**. The unit of temperature on the absolute scale is called **Kelvin** (K) and is related to Celsius scale whose unit is degree centigrade °C, by the formula: Kelvin temperature (K) = 273 + °C.

**Example**

Convert 25°C to Kelvin.

\[
K = 273 + 25 \\
= 298 \text{ K}
\]
1.4: Application of Charles’ law

Example 1

A gas occupies a volume of 100 cm$^3$ at 27°C. What will be the temperature of the gas if it occupied a volume of 200 cm$^3$?

**Solution**

All temperatures in °C must be converted to Kelvin (K).

\[ 0°C = 273K \]
\[ 27°C = 273 + 27 = 300K \]

Identify the variables; and substitute in Charles’ law equation then solve for the unknown.

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2}
\]

\[
\begin{align*}
V_1 &= 100 \text{ cm}^3 \\
T_1 &= 300K \\
V_2 &= 200 \text{ cm}^3 \\
T_2 &= ?
\end{align*}
\]

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2} \therefore \frac{100 \text{ cm}^3}{300K} = \frac{200 \text{ cm}^3}{T_2}
\]

\[
T_2 = \frac{V_2 \times T_1}{V_1}
\]

\[
= \frac{200 \text{ cm}^3 \times 300K}{100 \text{ cm}^3} = 600K
\]

The final temperature will be 600K.

Example 2

The volume of a gas at 30°C is 400 cm$^3$, what will be its volume at 50°C?

**Solution**

Convert the temperatures in °C to K.

\[
\begin{align*}
30°C &= (30 + 273)K \\
&= 303K \\
50°C &= (50 + 273)K \\
&= 323K.
\end{align*}
\]

\[
\begin{align*}
V_1 &= 400 \text{ cm}^3 \\
T_1 &= 303K
\end{align*}
\]
\[ V_2 = \ ? \]
\[ T_2 = \ 323K \]
\[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \]
\[ 400 \text{ cm}^3 = \frac{V_2}{303K} \]
\[ V_2 = \frac{V_1 \times T_2}{T_1} \]
\[ V_2 = \frac{323K \times 400 \text{ cm}^3}{303K} \]
\[ V_2 = \ 426.4 \text{ cm}^3 \]

The volume of the gas at 50\(^{\circ}\)C will be 426.4 cm\(^3\).

**Exercises 1.2**
2. How is Charles’ law represented mathematically?
3. A plastic bag containing potato chips, is enlarged with 100 cm\(^3\) of air at 20\(^{\circ}\)C. If the packet is placed in the sunlight where the temperature is 30\(^{\circ}\)C, what will be the final volume of the air?

**1.5: Combined Gas Law (The general gas equation)**
Boyle’s law and Charles’ law can be combined into a general gas equation. Boyle’s law, \[ P_1 V_1 = \text{constant} \] and Charles’ law, \[ \frac{V_1}{T_1} = \text{constant} \].

When the two equations are combined, the general gas equation is obtained. Thus
\[ \frac{P_1 V_1}{T_1} = \text{constant} \]
\[ \frac{P_1}{T_1} \]
\[ \frac{V_1}{T_1} \]

Note
- Boyle’s Law
- Charles’ law

Different physical conditions of a gas can be related. For example, if the initial conditions of pressure, volume and absolute temperature of a fixed mass of gas change from \( P_1, V_1 \) and \( T_1 \) to final conditions \( P_2, V_2 \) and \( T_2 \) it follows that:
\[ \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \]

\( P_1 \) – initial pressure
\( V_1 \) – initial volume
\( T_1 \) – initial absolute temperature
\( P_2 \) – final pressure
\( V_2 \) – final volume
\( T_2 \) – final absolute temperature
Note: Standard temperature and pressure is abbreviated as s.t.p. These are 273 K(0°C) and 760 mm Hg respectively.

Using the following formulae you can calculate any unknown variable.

\[
\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2},
\]

for example let us make \(P_1\) (pressure) the subject of the formula as shown below.

\[
P_1 = \frac{P_2V_2T_1}{T_2V_1}
\]

\[
V_1 = \frac{P_2V_2T_1}{T_2P_1}
\]

\[
T_1 = \frac{P_1V_1T_2}{P_2V_2}
\]

\[
P_2 = \frac{P_1V_1T_2}{T_1V_2}
\]

\[
V_2 = \frac{P_1V_1T_2}{P_2V_2}
\]

\[
T_2 = \frac{P_2V_2T_1}{P_1V_1}
\]

1.6: Application of combined gas law

Example 1
A certain mass of gas occupies 400 cm\(^3\) at 25°C and 760 mm Hg. What will be its volume at 0°C and 740 mm Hg?

Solution

\[
\begin{align*}
P_1 &= 760 \text{ mm Hg} \\
V_1 &= 400 \text{ cm}^3 \\
T_1 &= (25 + 273) = 298 \text{ K} \\
P_2 &= 740 \text{ mm Hg} \\
V_2 &= ? \\
T_2 &= (0 + 273) = 273 \text{ K}
\end{align*}
\]
Substitute the values in the equation

\[
V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}
\]

\[
V_2 = \frac{760 \text{ mm Hg} \times 400 \text{ cm}^3 \times 273 \text{ K}}{740 \text{ mm Hg} \times 298}
\]

\[
V_2 = \frac{82992000}{220520} \text{ cm}^3
\]

\[
V_2 = 376.3 \text{ cm}^3
\]

**Example 2**

In a class experiment, it was found that a gas occupied 48 dm\(^3\) at 25°C and 2 atmospheres pressure. What volume would it occupy at s.t.p?

**Solution**

\[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}
\]

From the above equation:

- \(P_1 = 2\) atmospheres
- \(V_1 = 48\) dm\(^3\)
- \(T_1 = (25 + 273) = 298\) K.
- \(P_2 = 1\) atmosphere
- \(V_2 = ?\)
- \(T_2 = 273\) K

Substitute the variables in the equation

\[
V_2 = \frac{P_1 V_1 T_2}{T_1 P_2}
\]

\[
V_2 = \frac{2 \text{ atm} \times 48 \text{ dm}^3 \times 273 \text{ K}}{298 \times 1 \text{ atm}}
\]

\[
V_2 = \left(\frac{26208}{298}\right) \text{ dm}^3
\]

\[
V_2 = 87.9 \text{ dm}^3
\]

**Example 3**

A certain mass of a gas occupies 220 cm\(^3\) at 18°C and 740 mm Hg pressure. At what temperature will the volume of the gas be 180 cm\(^3\) and the pressure 770 mm Hg?

\[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}
\]
From the above equation
\[ P_1 = 740 \text{ mm Hg} \]
\[ V_1 = 220 \text{ cm}^3 \]
\[ T_1 = (18 + 273) = 291 \text{K} \]
\[ P_2 = 770 \text{ mm Hg} \]
\[ V_2 = 180 \text{ cm}^3 \]
\[ T_2 = ? \]

Substitute in the equation
\[ T_2 = \frac{P_2 V_2 T_1}{P_1 V_1} \]
\[ T_2 = \frac{770 \text{ mm Hg} \times 180 \text{ cm}^3 \times 291 \text{K}}{740 \text{ mm Hg} \times 220 \text{ cm}^3} \]
\[ T_2 = \left( \frac{40,332,600}{162,800} \right) \text{K} \]
\[ T_2 = 247.7 \text{K} \]

**Exercises 1.3**
1. Write the combined gas law.
2. A certain mass of gas occupies 250 cm\(^3\) at 20°C and a pressure of 740 mm Hg. Calculate its volume at standard temperature and pressure.
3. Copy and complete the following table.

<table>
<thead>
<tr>
<th>Initial Pressure</th>
<th>Initial Volume</th>
<th>Initial Temperature</th>
<th>Final Pressure</th>
<th>Final Volume</th>
<th>Final Temperature</th>
</tr>
</thead>
<tbody>
<tr>
<td>(i) 760 mm Hg</td>
<td>200 cm(^3)</td>
<td>0°C</td>
<td>— mm Hg</td>
<td>400 cm(^3)</td>
<td>0°C</td>
</tr>
<tr>
<td>(ii) 2 atm</td>
<td>— cm(^3)</td>
<td>25°C</td>
<td>0.5 atm</td>
<td>200 cm(^3)</td>
<td>100°C</td>
</tr>
<tr>
<td>(iii) 1 atms</td>
<td>20l cm(^3)</td>
<td>— °C</td>
<td>1 atm</td>
<td>251 cm(^3)</td>
<td>18°C</td>
</tr>
<tr>
<td>(iv) 300 mm Hg</td>
<td>30 dm(^3)</td>
<td>100°C</td>
<td>700 mm Hg</td>
<td>— dm(^3)</td>
<td>50°C</td>
</tr>
<tr>
<td>(v) — mm Hg</td>
<td>48l cm(^3)</td>
<td>20°C</td>
<td>750 mm Hg</td>
<td>46l cm(^3)</td>
<td>25°C</td>
</tr>
</tbody>
</table>

**1.7: Graham’s law of diffusion**

**Diffusion in gases and liquids**

We have learnt that we can change the volume of a fixed mass of a gas by varying the pressure or temperature. We also earlier learnt that the particles of a gas or liquid are in a continuous random motion. Particles of gases and liquids move from areas of high concentration to areas of low concentration through a process called diffusion. We can define diffusion as the movement of molecules from regions where they are more concentrated to regions where they are less concentrated. Diffusion therefore means spreading out.
Diffusion is experienced in our everyday life. For example, if a bottle of perfume is opened or sprayed at one corner of a room, within a few seconds, a person at the other end of the room will be able to smell the perfume, Fig. 1.10. How does the scent reach the other parts of the room?

In the bottle of the perfume, the molecules of the perfume vapour are very concentrated. When the perfume is sprayed at one corner of the room, the molecules move from the original area (where they are more concentrated) to other parts of the room (where they are less concentrated). That is why someone at the other end of the room is able to smell the perfume, though the person is not close to the perfume bottle.

*Fig. 1.10 Diffusion of gases*

**Experiment 1.1**
To investigate diffusion of bromine fumes.

*Caution:* This experiment should be done in a fume chamber or an open space as bromine is poisonous.

**Apparatus and chemicals**
- gas jars
- bromine liquid
- teat pipette

**Procedure**
1. Using a teat pipette place about 2 to 3 cm³ of bromine liquid into a gas jar.
2. Invert another gas jar over it as illustrated in Fig. 1.11.
3. Leave it undisturbed for about 5 minutes.
   - What do you observe?
   - Record your observation in your notebook.
   - What do you conclude?
• What would happen if bromine water was used instead of bromine liquid? When bromine liquid is placed in a gas jar and another gas jar inverted over it, red-brown fumes of bromine are observed in the upper gas jar after a few seconds.

Diffusion of bromine

Bromine molecules diffuse from the lower gas jar into the upper gas jar. How do you think bromine moved into the upper jar?

Diffusion occurs because of the random movement of molecules in liquid and gaseous substances. Although bromine fumes are denser than air they are capable of moving to the upper gas jar.

Experiment 1.2
To investigate diffusion in liquids

Apparatus and chemicals
• glass tube
• potassium manganate(VII) or copper(II) sulphate crystals
• water
• beaker
Procedure
1. Put water in a 250 cm\(^3\) beaker.
2. Place a glass tube at the middle of the beaker and hold it tightly so that there is no disturbance.
3. Put a small amount of potassium manganate(VII) or copper(II) sulphate at the middle of the water in a beaker through the glass tube.
4. Close the glass tube with first finger carefully to avoid disturbing the potassium manganate(VII) crystals or copper(II) sulphate crystals as shown in Fig. 1.12. (the tube is closed to hold the coloured liquid inside the tube thereby preventing it from spreading out). Remove the tube from the beaker slowly.

![Fig. 1.12 Demonstrating diffusion in liquids](image)

5. Leave the beaker to stand for about 30 minutes. What do you observe? Record your observation in your notebook.
   - Suppose you were given crystals of sugar instead of potassium manganate(VII) or copper(II) sulphate crystals, what would you observe? Explain.
   - Suggest how we would know that diffusion has taken place. What would you conclude from these experiments?

When potassium manganate(VII) is put in water, the water slowly turns purple; but with copper(II) sulphate the water turns blue.

Liquids are composed of particles that are less closely packed and indeed move freely. Since there is more space between the particles of a liquid, diffusion occurs at a faster rate than in solids. From the above experiment the coloured potassium manganate(VII) and copper(II) sulphate crystals which are also made of particles dissolve as a result of particles moving slowly from the crystals into the water. These particles keep on colliding with water particles as they move. Eventually homogenous coloured solutions are obtained.
Diffusion in terms of Kinetic theory
The kinetic theory explains the forces between molecules and the energy that they possess. This theory has two basic assumptions:

- Matter is composed of small particles.
- The particles are in constant random motion.

1. Particles in solids are held close to each other by strong forces of attraction operating between the particles. They are held in fixed positions in a regular arrangement (lattice). They vibrate within the same positions.

2. Particles in liquids also have some moderate forces of attraction between them but they are free to move past each other yet, they tend to stick together. They are also constantly in random motion.

3. Particles in gases have negligible forces of attraction between them because they are far apart and are free to move faster in random motion. Consequently they have a lot of kinetic energy (K.E.).

Fig. 1.13: States of matter

Rate of diffusion of gases
The rate (speed) of diffusion of a gas is a measure of the quantity of the gas that passes through a space after a particular period of time.

Experiment 1.3
To investigate the rate of diffusion of hydrogen chloride and ammonia gases.

Caution: These gases are corrosive to the skin and the eyes. They are potentially respiratory hazards if inhaled in large quantities. However ammonia is used in smelling salts since it acts on the heart and prevents fainting and dizziness.

Apparatus and chemicals
- 2 clamps and stands
- long glass tube
• cotton wool
• concentrated ammonia solution
• concentrated hydrochloric acid

**Procedure**
1. Hold the long glass tube firmly horizontally with the clamps.
2. At the same time, plug one end of the long glass tube with cotton wool soaked in concentrated ammonia solution and the other end with other cotton wool soaked in concentrated hydrochloric acid as illustrated in Fig 1.14(a)

![Diagram of glass tube with cotton wool plugs](image)

*Fig. 1.14(a) Diffusion of ammonia and hydrogen chloride gases*

3. Leave the set up for about 10-12 minutes.
   - What do you observe in the glass tube? Record the observations in your notebook.
   - What do we conclude from this experiment?

Hydrogen chloride and ammonia gases diffuse from opposite ends of the long glass tube. When they meet they react to produce ammonium chloride which appears as a white ring as illustrated in Fig. 1.14(b)

\[
\text{Ammonia} + \text{hydrogen chloride} \rightarrow \text{ammonium chloride}
\]

\[
\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl(s)}
\]

*Fig. 1.14(b) Distances moved by ammonia and hydrogen chloride gases*
The rate of diffusion of the gases depends on the molecular mass of the gas. It also depends on the density of the gas, since the density of the gas increases with its molecular mass. The relationship between rate of diffusion of a gas with its density was first discovered in 1833 by an English chemist named Graham Fig. 1.15.

He carried out experiments using gases with different densities. He found out that a less dense gas would diffuse through a medium much faster than a denser gas. Based on this, he formulated a law, which is known as Graham’s law of diffusion.

Graham’s law of diffusion

Graham’s law of diffusion states that at a constant temperature and pressure, the rate of diffusion of a gas is inversely proportional to the square root of its density. Mathematically, this relationship is expressed as follows:

\[ \text{Rate of diffusion} (r) \propto \frac{1}{\sqrt{\rho}} \]

where \( \rho \) is density.

Therefore:

\[ r \propto \frac{1}{\sqrt{\rho}} \]

\[ r = \frac{\text{constant}}{\sqrt{\rho}} \]

\[ r \sqrt{\rho} = \text{constant} \]

If the rates of diffusion of two gases, 1 and 2 are compared, the ratio of their rates of diffusion will be equal to the ratio of the inverses (reciprocal) of the square roots of their densities. This statement is expressed mathematically as follows:
In the above expression
\[ \frac{r_1}{r_2} = \sqrt{\frac{\rho_2}{\rho_1}} \]

\( r_1 \) represents the rate of diffusion of gas 1
\( r_2 \) represents the rate of diffusion of gas 2
\( \rho_1 \) represents the density of gas 1
\( \rho_2 \) represents the density of gas 2.

Certain modification can be made to this law. This modification is made depending on the relationship between the rate of diffusion of a gas and the time taken to diffuse. And also the relationship between the rate of diffusion and molecular mass.

- The rate of diffusion of the gas increases as its time(t) of diffusion decreases.
- The density (\( \rho \)) of a gas increases as its molecular mass(M) increases.

Mathematically, the statements can be summarised as shown below.

\[ \frac{\rho_2}{\rho_1} = \frac{M_2}{M_1} \]

\( \rho_1 \) represents density of gas 1
\( \rho_2 \) represents density of gas 2
\( M_1 \) represents molecular mass of gas 1
\( M_2 \) represents molecular mass of gas 2

\[ \frac{r_2}{r_1} = \sqrt{\frac{M_1}{M_2}} \]

\( r_1 \) represents rate of diffusion of gas 1
\( r_2 \) represents rate of diffusion of gas 2
\( M_1 \) represents molecular mass of gas 1
\( M_2 \) represents molecular mass of gas 2

\[ \frac{r_2}{r_1} = \frac{t_2}{t_1} \]

\( r_1 \) represents the rate of diffusion of gas 1
\( r_2 \) represents the rate of diffusion of gas 2
\( t_1 \) represents the time of diffusion of gas 1
\( t_2 \) represents the time of diffusion of gas 2
For equal volumes of two gases, the rate of diffusion $\alpha \frac{1}{\text{time}}$

\[ \therefore \frac{\text{Time of diffusion of gas 1}}{\text{Time of diffusion of gas 2}} = \sqrt{\frac{\text{molecular mass of gas 1}}{\text{molecular mass of gas 2}}} \]

\[ \frac{t_1}{t_2} = \sqrt{\frac{M_1}{M_2}} \]

$t_1$ = time of diffusion of gas 1
$t_2$ = time of diffusion of gas 2

You may use any of the above equations for calculations depending on the variables that you are given and what you are looking for.

**Example 1**

If 10 cm$^3$ of an unknown gas takes 6.3 seconds to pass through a small orifice while 1 cm$^3$ of oxygen, O$_2$, takes 5.6 seconds through the same orifice under the same conditions of temperature and pressure. What will be the molecular mass of the unknown gas? (O = 16)

Use the following equation:

\[ \frac{t_1}{t_2} = \sqrt{\frac{M_1}{M_2}} \]

From the question above we can substitute the variables

$t_1 = 6.3$ seconds
$t_2 = 5.6$ seconds

$M_1$ (unknown gas) = ?

$M_2$ (O$_2$) = (16 + 16) = 32

\[ \frac{6.3}{5.6} = \sqrt{\frac{M_1}{32}} \]

\[ \left(\frac{6.3}{5.6}\right)^2 = \frac{M_1}{32} \]

\[ \frac{39.69}{31.36} = \frac{M_1}{32} \]
25

\[
\frac{32 \times 39.69}{31.36} = M_1
\]

\[
M_1 = 40.5
\]

∴ The molecular mass of the unknown gas is 40.5 g

**Example 2**

A given volume of nitrogen gas, N\(_2\), requires 68.3 seconds to diffuse through a tiny hole in a chamber. Under the same conditions, another unknown gas requires 85.6 seconds for the same volume to diffuse. What is the molecular mass of this gas?

\[
M_1 (N_2) = 14 + 14 = 28
\]

\[
M_2 \text{ (unknown gas)} = ?
\]

\[
t_1 = 68.3\text{s}
\]

\[
t_2 = 85.6\text{s}
\]

\[
\frac{t_1}{t_2} = \sqrt{\frac{28}{M_2}}
\]

\[
\frac{68.3}{85.6} = \sqrt{\frac{28}{M_2}}
\]

\[
\left(\frac{68.3}{85.6}\right)^2 = \frac{28}{M_2}
\]

\[
\frac{4664.89}{7327.36} = \frac{28}{M_2}
\]

\[
0.637 = \frac{28}{M_2}
\]

\[
M_2 = \frac{28}{0.637}
\]

\[
= 43.98 \text{ g}
\]

**Exercises 1.4**

1. The relative molecular masses of ammonia and hydrogen chloride gases are 17 and 36.5 respectively. If the two gases are subjected to the same conditions of temperature and pressure, find the rate at which ammonia will diffuse faster than hydrogen chloride.
2. A sample of an unknown gas X is shown by analysis to contain only sulphur and oxygen. The gas requires 28.3s to diffuse through an orifice into a vacuum. An identical number of oxygen molecules, O₂, passes through the same orifice in 20.0s. Determine the formula and molecular mass of the gas.
\(O = 16, S = 32\).

3. A gas G diffuses through a small aperture at the rate of 2 cm³ s⁻¹. Oxygen diffuses through the same aperture at rate of 3 cm³ s⁻¹. Calculate the molar mass of gas G. \(O = 16\)

**Summary**

1. All matter is composed of tiny particles. In gases their particles exert pressure on the walls of the container.

2. A gas may be readily compressed since the particles of the gas are far apart.

3. Boyle’s law states that the **volume of a fixed mass of a gas is inversely proportional to its pressure provided that the temperature remains constant.** This law can be stated mathematically as

\[
P \propto \frac{1}{V} \quad \text{or} \quad PV = \text{a constant}
\]

4. Charles’ law states that the **volume of a fixed mass of a gas is directly proportional to its absolute temperature provided the pressure remains constant.**

5. Direct proportionality is a mathematical expression meaning that the volume increases in the same proportion as the Kelvin temperature increases.

6. Absolute temperature is a temperature measured with respect to absolute zero on the Kelvin scale — absolute zero is the lowest possible temperature that can be achieved.

7. Charles’ law can be expressed mathematically as

\[
V \propto T
\]

or

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2}
\]

\(V_1 = \text{initial volume}\)

\(T_1 = \text{initial temperature}\)

\(V_2 = \text{Final volume}\)

\(T_2 = \text{Final temperature}\)

8. Kinetic theory of gases accounts for the fact that:

- gases consists of tiny particles (molecules) which are in constant motion.
- the motion of the particles is random in all directions.
- during this random motion the particles collide with each other and with the walls of the container.
9. When we combine Boyle’s and Charles’ Laws we get the general gas equation which can be written as follows

\[ \frac{PV}{T} = \text{constant} \]

\[ \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \]

\( V_1 \) is the initial volume at initial pressure \( P_1 \) and initial temperature \( T_1 \).
\( V_2 \) is the final volume at final pressure \( P_2 \) and final temperature \( T_2 \).

**Note**

- The physical factors which when changed affect the volume of a gas are temperature and pressure
- Pressure at s.t.p.
  - 1 atmosphere pressure = 760 mm Hg = 101325 Pa
- Temperature
  - Kelvin (K) and degree Celsius(°C)
  - Only Kelvin is used in calculations. To convert °C to K;
    \[ \text{Kelvin} = \text{°C} + 273. \]
- If pressure is given in mm Hg use mm Hg all along. If it is given in atmospheres, use atmospheres all along.

**Revision Exercise 1**

1. State and explain Boyle’s law.
2. A given mass of gas occupies 30 cm\(^3\) at 790 mm Hg pressure. What will be its volume at 760 mm Hg pressure at constant temperature?
3. The pressure of 250 cm\(^3\) of a gas X is decreased from 790 mm Hg to 760 mm Hg at constant temperature. Find the new volume.
4. Draw a sketch graph of \( P \) against \( V \) of a gas at constant temperature.
5. At constant temperature, the pressure of 150 cm\(^3\) of gas was increased from 720 mm Hg to 780 mm Hg. What is the new volume?
6. 380 cm\(^3\) of a gas R has a pressure of 85 kpa. What will be its new volume if the pressure is reduced to 83 kpa, while the temperature remains constant?
7. Draw a sketch graph of \( P \) against \( \frac{1}{V} \) of a gas X at constant temperature.
8. A student wants to find out whether a balloon would burst. The balloon can hold a maximum of 1000 cm\(^3\) of air. The balloon contains 980 cm\(^3\) of air at 10°C. Will the balloon burst if the student enters in to a room with a temperature of 25°C, Assume the pressure of the air remains constant.
9. A sample of gas occupies 225 cm\(^3\) at 30°C. Find the new volume at 100°C if the pressure remains constant.
10. Draw a sketch graph of \( V \) against \( \frac{1}{P} \) of a gas at constant temperature.
12. Plot a sketch graph of V against T(K) of a gas X at constant pressure.
14. A fixed mass of a certain gas X occupies 75 cm$^3$ at 36˚C and 85 kPa. Calculate the volume the gas would occupy at s.t.p.
15. A fixed mass of a certain gas occupies 500 cm$^3$ at 28˚C and 101325 Pa. Find its volume at 0˚C and 100 000 Pa.
16. To what temperature must 2 litres of air at 17˚C be heated at constant pressure in order to double the volume.
17. Imagine you are travelling on a long journey from Mombasa to Kisumu. Before you start the journey, your tyre pressure is 3 \times 10^4 \text{ Pa} and the temperature is 16˚C. At the end of the journey, the pressure has increased to 3.2 \times 10^4 \text{ Pa}. Calculate the temperature of the air in the tyre at the end of the journey. Assume the volume of the tyre remains constant.
Unit 2: The Mole: Formulae and chemical equations

Introduction

What is a mole?
The following are some meanings of the word “mole”.

(a) a small burrowing mammal
(b) a small dark mark on the skin

Fig. 2.1: Some meanings of moles

2.1: The mole as a basic unit in chemistry

In the previous unit we learnt that matter consists of very tiny particles which are in continuous random motion. These particles include: atoms, molecules, ions, electrons among others. It is impossible to weigh a single particle because it is too small. This is why the mole is used as a weighing unit. Before we discuss the mole in detail, let us start with some of the counting units we are familiar with in everyday life as shown in Table 2.1.

Table 2.1: Some of the counting units

<table>
<thead>
<tr>
<th>Units</th>
<th>Items</th>
<th>Number of items</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pair of</td>
<td>• shoes</td>
<td>2</td>
</tr>
<tr>
<td></td>
<td>• socks</td>
<td>2</td>
</tr>
<tr>
<td></td>
<td>• feathers</td>
<td>2</td>
</tr>
<tr>
<td></td>
<td>• earrings</td>
<td>2</td>
</tr>
<tr>
<td></td>
<td>• stockings</td>
<td>2</td>
</tr>
<tr>
<td>Dozen of</td>
<td>• exercise books</td>
<td>12</td>
</tr>
<tr>
<td></td>
<td>• pencils</td>
<td>12</td>
</tr>
<tr>
<td></td>
<td>• eggs</td>
<td>12</td>
</tr>
</tbody>
</table>
From the table above, we can see that a pair represents 2 items whereas a dozen represents 12 items. In Chemistry, the mole represents a fixed number of items. This number is $602\,300\,000\,000\,000\,000\,000\,000$. This number is written to 4 significant figures as $6.023 \times 10^{23}$. The mole therefore is such a large number. As you can see it is more than the people on planet, Earth! Astronomers estimate that there is a mole of stars in the universe. This is a “magic number” in Chemistry and it is called **Avogadro’s number or constant**, named in honour of an Italian chemist – Amedeo Avogadro. It is represented by letter $L = 6.023 \times 10^{23}$ particles.

In Chemistry, we deal with elementary particles such as **atoms, molecules, ions** and **electrons**. Each mole of these particles contains Avogadros number of particles i.e. $6.023 \times 10^{23}$. For example one mole of sodium (Na) atoms contains $6.023 \times 10^{23}$ atoms. Similarly one mole of electrons contains $6.023 \times 10^{23}$ electrons and one mole of nitrogen molecules ($N_2$) also contains $6.023 \times 10^{23}$ molecules. See Table 2.2 below.

**Table 2.2: Number of particles in a mole**

<table>
<thead>
<tr>
<th>Particles</th>
<th>Element</th>
<th>Number of moles</th>
<th>Number of particles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atoms</td>
<td>Na</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
</tr>
<tr>
<td></td>
<td>Mg</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
</tr>
<tr>
<td></td>
<td>O</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
</tr>
<tr>
<td>Molecules</td>
<td>$O_2$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
</tr>
<tr>
<td></td>
<td>$N_2$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
</tr>
<tr>
<td></td>
<td>$CO_2$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
</tr>
<tr>
<td>Ions</td>
<td>Na$^+$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
</tr>
<tr>
<td></td>
<td>SO$_4^{2-}$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
</tr>
<tr>
<td></td>
<td>O$_2^-$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
</tr>
<tr>
<td>Electron</td>
<td>e$^-$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
</tr>
</tbody>
</table>

**Definition of a mole**

The mole of a substance is the amount which contains as many elementary units as there are atoms in 12.0 g of carbon – 12 isotope.
The mole is abbreviated as mol.

Just as we have one, two, three, $\frac{1}{2}$ or $\frac{1}{4}$, pairs or dozens of things we can also have 1, 2, 3, $\frac{1}{2}$ or $\frac{1}{4}$, mole(s) of particles.

Now, let us find the number of particles in several moles. One mole contains $6.023 \times 10^{23}$ particles. Copy Table 2.3 and 2.4 in your notebook and fill in the number of items and particles.

Table 2.3: Number of particles in a mole

<table>
<thead>
<tr>
<th>Number of pairs</th>
<th>Number of items</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>$1 \times 2 = 2$</td>
</tr>
<tr>
<td>2</td>
<td>$2 \times 2 = 4$</td>
</tr>
<tr>
<td>3</td>
<td>—</td>
</tr>
<tr>
<td>0.5</td>
<td>—</td>
</tr>
</tbody>
</table>

Table 2.4: Number of particles in several moles

<table>
<thead>
<tr>
<th>Number of moles</th>
<th>Number of particles</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>$1(6.023 \times 10^{23}) = 6.023 \times 10^{23}$</td>
</tr>
<tr>
<td>2</td>
<td>$2(6.023 \times 10^{23}) = 12.046 \times 10^{23}$</td>
</tr>
<tr>
<td></td>
<td>$= 1.2046 \times 10^{24}$</td>
</tr>
<tr>
<td>3</td>
<td>—</td>
</tr>
<tr>
<td>4</td>
<td>—</td>
</tr>
<tr>
<td>$\frac{1}{2}$</td>
<td>—</td>
</tr>
</tbody>
</table>

Note: The mole is just a quantity of a substance. But it deals with very tiny particles. It is necessary to state the type of particles we are dealing with, i.e. are they atoms, molecules, ions, and so on. For example, if we are dealing with oxygen, we should state whether we are referring to one mole of oxygen atoms (O) or oxygen molecules (O$_2$). We have seen that 1 mole of any substance contains very many particles.
How then can such large amount of particles be counted?
- Imagine you are employed in a bank. You are given the task of counting one-shilling coins amounting to one billion shillings. If you took one second to count two coins, how long will it take you to count one billion shillings?
- How do you think coins are counted in banks?

You may have noticed from your calculations that it takes a long time to count one billion shilling coins. In the bank, coins are counted by weighing. This takes a shorter time compared to physical counting. This counting method is called counting by weighing.

**Activity 2.1**
To count the number of particles in samples of sugar and maize flour.

**Apparatus and chemicals**
- Maize flour
- Sucrose (sugar)
- Petri dishes or saucers
- Weighing balance

**Procedure**
1. Place approximately the same amount of flour and sugar in petri dishes as shown in Fig. 2.3.

![Fig. 2.3: Counting particles in samples of flour and sugar](image)

32
2. Try to count the particles.
   • What do you observe?
   • What can you conclude?
   • How can we know the number of particles in each sample?

3. Weigh each sample and record the mass as shown below.

   | Mass of flour  | $x\ g$ |
   | Mass of sugar  | $y\ g$ |

You may have noticed that it is impossible to count the particles in a sample of flour or sugar. However, we can find out the number of particles in a sample of a substance from the mass of that sample. This mass is determined by weighing.

It has been determined experimentally that 12 g of carbon-12 contains $6.023 \times 10^{23}$ atoms. This simply means that if we place $6.023 \times 10^{23}$ atoms of carbon–12 on a balance, they would weigh 12 g.

One mole ($6.023 \times 10^{23}$) of carbon-12 atoms = 12 g. So instead of counting $6.023 \times 10^{23}$ atoms of carbon, we can simply weigh 12 g. This means that if you have 36 g or 3 moles of carbon-12, then you have $1.8069 \times 10^{24}$ atoms of carbon–12.

### 2.2: Relative Atomic mass

As earlier mentioned atoms are so tiny that it is difficult to obtain the mass of a single atom of any element. However, often times, scientists would like to know how masses of atoms of different elements compare.

To achieve this the mass of an atom of hydrogen (the lightest known element) has been used by scientists as the standard mass with which atoms of other elements can be compared. The mass of a hydrogen atom (or atomic mass of hydrogen) is 1 (that is, H=1). Therefore to get the relative atomic mass of an element, the mass of its atom is compared with that of a hydrogen atom.

#### Examples

An atom of oxygen has a mass 16 times heavier than that of hydrogen. An atom of carbon has a mass 12 times heavier than a hydrogen atom. Potassium and sodium atoms are 39 and 23 times heavier than the mass of a hydrogen atom respectively. The values 16, 12, 39 and 23 are referred to as Relative Atomic Masses (R.A.M or $A_r$).

Therefore $\text{R.A.M of an element} = \frac{\text{mass of 1 atom of that element}}{\text{mass of 1 atom of hydrogen}}$
However in 1961, the International Union of Pure and Applied Chemistry (IUPAC), choose a carbon atom to be the standard atom. The atom chosen is that of carbon – 12\(^{12}\text{C}\), the most abundant of carbon isotopes. On the carbon – 12 scale, carbon is assigned a mass of exactly 12.000.

Therefore R.A.M \[ \text{mass of 1 atom of an element} \times \frac{1}{12} \text{ of the mass of 1 atom of carbon–12} \]

**Note**

All other atoms are compared with this standard atom, using an instrument called a mass spectrometer.

Standard carbon atom

\[ \text{Mass} = 12.000 \]

Hydrogen atom is \( \frac{1}{12} \) as heavy as standard atom. Its R.A.M = 1

Magnesium atom is twice as heavy as the standard atom. Its R.A.M = 24

Use a periodic table to find how many times a potassium atom is as heavy as the standard atom.

**Fig. 2.4 Comparing masses of different elements with carbon–12**

The mass of an atom that we get by comparing it with the \( \frac{1}{12} \)th of the carbon – 12 \(^{12}\text{C}\) atom is called its relative atomic mass or R.A.M as shown in Table 2.5.

**Relative Atomic Mass has no units since it is a comparison of two masses.**

**Relative atomic mass and isotopes**

All atoms of an element are not exactly the same. We know that some have more neutrons than others. For example, \( \frac{1}{4} \) of chlorine isotopes have 20 neutrons hence have a mass of 37. The rest, \( \frac{3}{4} \), have 18 neutrons and a mass of 35.

Let us calculate the average mass of a chlorine atom.
Most elements have more than one isotope, and we must consider this when finding their relative atomic masses.

The relative atomic mass of an element is defined as the average mass of one atom of the element compared to \( \frac{1}{12} \) of the mass of one atom of carbon – 12 \(^{12}\text{C}\).

**Table 2.5 Relative Atomic Masses of some common elements**

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Relative atomic mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>H</td>
<td>1</td>
</tr>
<tr>
<td>Carbon</td>
<td>C</td>
<td>12</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N</td>
<td>14</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O</td>
<td>16</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na</td>
<td>23</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg</td>
<td>24</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl</td>
<td>35.5</td>
</tr>
<tr>
<td>Potassium</td>
<td>K</td>
<td>39</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca</td>
<td>40.1</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe</td>
<td>55.8</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu</td>
<td>63.5</td>
</tr>
</tbody>
</table>

**Note:** The relative atomic masses of elements are usually rounded off to whole numbers, to make calculations much easier.

**Application of relative atomic mass**

We have already seen that a pair is made up of 2 items, whereas one dozen consists of 12 items, and 1 mole consists of \(6.023 \times 10^{23}\) particles. Let us now find out whether 1 pair of different items weigh the same.

**Experiment 2.1**

To investigate whether a pair of different items have the same mass

**Apparatus**
- beam balance
- potatoes
- tomatoes
- oranges
Procedure

1. Choose one pair of potatoes and one pair of tomatoes or oranges of the same size as the potatoes.
2. Place the pair of tomatoes/oranges on one side of the beam balance and the pair of potatoes on the other as shown in Fig. 2.5. Do they balance? If not, which pair is heavier than the other?

3. Repeat procedure 2 using other items such as earrings, pens and pencils. What do you observe?
4. Record your observations in your notebook. What do you conclude?

A student was supplied with a dozen chicken feathers of the same size. Explain briefly how the student would find the number of pieces of paper each measuring 6 cm × 2 cm that have the same mass as the feathers.

You may have observed that, although a pair consists of two items, the items may have different masses. In the same way, though a dozen has twelve items, the items may have different masses. Do you expect 1 mole of different particles to have different or similar masses?

If we weigh 12 g of carbon atoms and 1 g of hydrogen atoms, they will contain the same number of atoms. 24 g of magnesium will also contain the same number of atoms as 1 g of hydrogen atoms. In other words, the relative atomic mass of all elements expressed in grams contain the same number of atoms. Experiments have shown this number to be $6.023 \times 10^{23}$. It means that if we placed these number of atoms on a balance, they would weigh the same as the relative atomic mass as illustrated in Fig. 2.6.
Examples

Fig. 2.6: *Masses of 1 mole of different atoms*

Therefore, elements have the same number of particles in one mole but different relative atomic masses as shown in Tables 2.6, 2.7 and 2.8 below.

### Table 2.6: Relative Atomic Mass of some elements

<table>
<thead>
<tr>
<th>Atom</th>
<th>No. of particles (atoms)</th>
<th>Relative Atomic Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>N</td>
<td>$6.023 \times 10^{23}$</td>
<td>14</td>
</tr>
<tr>
<td>S</td>
<td>$6.023 \times 10^{23}$</td>
<td>32</td>
</tr>
<tr>
<td>Cl</td>
<td>$6.023 \times 10^{23}$</td>
<td>35.5</td>
</tr>
<tr>
<td>K</td>
<td>$6.023 \times 10^{23}$</td>
<td>39</td>
</tr>
<tr>
<td>Cu</td>
<td>$6.023 \times 10^{23}$</td>
<td>63.5</td>
</tr>
</tbody>
</table>

### Table 2.7: Relative Atomic Mass of some ions

<table>
<thead>
<tr>
<th>Ion</th>
<th>No. of moles</th>
<th>No. of particles (ions)</th>
<th>Relative mass of ions</th>
</tr>
</thead>
<tbody>
<tr>
<td>K$^+$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
<td>39</td>
</tr>
<tr>
<td>Na$^+$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
<td>23</td>
</tr>
<tr>
<td>Zn$^{2+}$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
<td>65</td>
</tr>
<tr>
<td>Cl$^-$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
<td>35.5</td>
</tr>
<tr>
<td>CO$_3^{2-}$</td>
<td>1</td>
<td>$6.023 \times 10^{23}$</td>
<td>60</td>
</tr>
</tbody>
</table>

**Note:** Gain or loss of electrons during ion formation does not change the R.A.M because the mass of an electron is negligible.
Scientists count atoms, ions and molecules by weighing. If we could count $6.0 \times 10^{23}$ (1 mole of atoms, and then place them on a balance, then we would get atomic mass of the element in grams. From Table 2.6, we can express the values in grams only after weighing. Otherwise they should remain as relative atomic masses with no units.

**Note**
For purposes of calculation we will use $L = 6.0 \times 10^{23}$.

### 2.3: Conversions in mole

**(i) Conversion of grams to moles and number of particles**

**Example 1**
Suppose during the laboratory preparation of oxygen we collected 32 g of oxygen. How many atoms would we have collected? ($O = 16; L = 6.0 \times 10^{23}$)

**Solution**
**Note** the values in the brackets are always given and must be used as follows

**Method 1 : First principle method**

\[
1 \text{ mole of oxygen atoms(O)} = 16 \text{ g}
\]
\[
x \text{ mol of oxygen atoms(O)} = 32 \text{ g}
\]
Cross multiply and solve for $x$
\[
16 \text{ g} \times x = 32 \text{ g} \times 1 \text{ mole}
\]
\[
x = \frac{32 \text{ g} \times 1 \text{ mole}}{16 \text{ g}}
\]
\[
x = 2 \text{ moles}
\]
Then, we say:-

1 mole of atoms = \(6.0 \times 10^{23}\) atoms

\[\therefore \text{2 moles} = 2 \times 6.0 \times 10^{23}\text{ atoms}\]
\[= 12.0 \times 10^{23}\text{ atoms}\]
\[= 1.2 \times 10^{24}\text{ atoms}\]

**Method 2: Formula method**

Moles of atoms \(= \frac{\text{Mass}}{\text{R.A.M}}\)

\[= \frac{32\text{ g}}{16\text{ g}}\]
\[= 2 \text{ moles}\]

1 mole = \(6.0 \times 10^{23}\) atoms

\[\therefore 2 \text{ moles} = 2 \times 6.0 \times 10^{23}\text{ atoms}\]
\[= 1.2 \times 10^{24}\text{ atoms}\]

*Note* You can re-arrange the equation and find any other unknown.

E.g mass = moles \(\times\) R.A.M

---

**Example 2**

If 54 g of aluminium were obtained from aluminium ore, how many atoms were obtained? (Al = 27; L = \(6.0 \times 10^{23}\))

**Solution**

**Method 1: First principle method**

\[1 \text{ mole of Al} \times 27\text{ g.}\]

\[\therefore y \text{ mole of Al} = 54\text{ g}\]

Let us cross multiply and solve for y

\[27\text{ g} \times y = 54\text{ g} \times 1 \text{ mole}\]

\[y = \frac{54\text{ g}}{27\text{ g}} = 2 \text{ moles}\]

1 mole = \(6.0 \times 10^{23}\) atoms

\[\therefore 2 \text{ moles} = 2 \times 6.0 \times 10^{23}\text{ atoms}\]
Example 3
How many ions are there in 13 g of K\(^+\) ions? (K = 39, L = 6.0 \times 10^{23})

**Solution**

**First principle method**

1 mole of K\(^+\) ions = 39 g
y moles of K\(^+\) ions = 13 g

Let us cross multiply and solve for y

\[
\frac{13 \text{ g}}{39 \text{ g}} = \frac{1}{3} \text{ mole}
\]

1 mol contains 6.0 \times 10^{23} ions

\[
\frac{1}{3} \text{ mole contains } y \text{ ions}
\]

Let us cross multiply and solve for y

\[
y = \frac{1}{3} \times 6.0 \times 10^{23} \text{ ions}
\]

**Formula method**

\[
\text{Moles} = \frac{\text{Mass}}{\text{R.A.M.}}
\]

\[
= \frac{13 \text{ g}}{39 \text{ g}} = \frac{1}{3} \text{ mole}
\]

1 mole contains 6.0 \times 10^{23} ions

\[
\frac{1}{3} \text{ moles contains } \frac{1}{3} \times 6.0 \times 10^{23} = 2.0 \times 10^{23} \text{ ions}
\]
Example 4
How many grams of sulphur contain:
(a) $6.0 \times 10^{20}$ atoms
(b) $1.5 \times 10^{23}$ atoms
\[(L = 6.0 \times 10^{23}, S = 32)\]

Solution
Remember that 1 mole of sulphur(S) contains $6.0 \times 10^{23}$ atoms.

Therefore:
(a) 1 mol of S atoms weighs 32 g
32 g of S contain $6.0 \times 10^{23}$ atoms thus:-

\[x \text{ g of S contain } 6.0 \times 10^{20} \text{ atoms}\]
32 g of S contain $6.0 \times 10^{23}$ atoms
\[x \text{ g of S contain } 6.0 \times 10^{20}\]
Let us cross multiply and solve for x.
\[
\therefore \quad x \text{ g } \times 6.0 \times 10^{23} = \frac{32 \text{ g } \times 6.0 \times 10^{20} \text{ atoms}}{6.0 \times 10^{23} \text{ atoms}}
\]
\[
x = 32 \times 10^{20} \times 10^{-23}
= 3.2 \times 10^{21} \times 10^{-23}
= 3.2 \times 10^{-2} \text{ g}
\]

Therefore $6.0 \times 10^{20}$ atoms of S = 0.032 g

(b) \[x \text{ g of S contain } 1.5 \times 10^{23} \text{ atoms}\]

Let us cross multiply and solve for x.
\[
x = \frac{32 \text{ g } \times 1.5 \times 10^{23} \text{ atoms}}{6.0 \times 10^{23} \text{ atoms}} = 8.0 \text{ g}
\]

\[
\therefore \quad 1.5 \times 10^{23} \text{ atoms of S = 8.0 g}
\]

Exercise 2.1
1. Find the number of atoms contained in the following elements:

<table>
<thead>
<tr>
<th>Relative atomic masses</th>
</tr>
</thead>
<tbody>
<tr>
<td>S = 32</td>
</tr>
<tr>
<td>Ca = 40</td>
</tr>
<tr>
<td>Al = 27</td>
</tr>
<tr>
<td>Fe = 56</td>
</tr>
<tr>
<td>L = 6.0 \times 10^{23}</td>
</tr>
</tbody>
</table>

(a) 8 g sulphur
(b) 80 g calcium
(c) 14.5 g aluminum
(d) 56 g iron
2. Given Avogadro’s constant, \( L = 6.0 \times 10^{23} \text{ mol}^{-1} \)
Calculate the number of atoms in each of the following.

(a) 11.5 g sodium
(b) 8 g magnesium
(c) 14 g iron
(d) 17.75 g chlorine

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Valency</th>
<th>Name</th>
<th>Symbol</th>
<th>Valency</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium</td>
<td>Na</td>
<td>1</td>
<td>Iron(III)</td>
<td>Fe^{3+}</td>
<td>3</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg</td>
<td>2</td>
<td>Chloride</td>
<td>Cl^{-}</td>
<td>1</td>
</tr>
<tr>
<td>Silver</td>
<td>Ag</td>
<td>1</td>
<td>Ammonium</td>
<td>NH_{4}^{+}</td>
<td>1</td>
</tr>
<tr>
<td>Barium</td>
<td>Ba^{2+}</td>
<td>2</td>
<td>Nitrate</td>
<td>NO_{3}^{-}</td>
<td>1</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca^{2+}</td>
<td>2</td>
<td>Hydrogensulphate</td>
<td>HSO_{4}^{-}</td>
<td>1</td>
</tr>
<tr>
<td>Copper (II)</td>
<td>Cu^{2+}</td>
<td>2</td>
<td>Carbonate</td>
<td>PO_{4}^{3-}</td>
<td>3</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg^{2+}</td>
<td>2</td>
<td>Sulphate</td>
<td>SO_{3}^{2-}</td>
<td>2</td>
</tr>
<tr>
<td>Oxide</td>
<td>O^{2-}</td>
<td>2</td>
<td>Sulphate</td>
<td>SO_{4}^{2-}</td>
<td>2</td>
</tr>
<tr>
<td>Sulphide</td>
<td>S^{2-}</td>
<td>2</td>
<td>Phosphat</td>
<td>PO_{4}^{3-}</td>
<td>3</td>
</tr>
<tr>
<td>Aluminium</td>
<td>Al^{3+}</td>
<td>3</td>
<td>Hydroxide</td>
<td>OH^{-}</td>
<td>1</td>
</tr>
</tbody>
</table>

3. Calculate the mass of
(a) 1 atom of oxygen
(b) 10 atoms of sodium
(c) 5 atoms of aluminium
(d) 3 atoms of potassium

4. An atom of a certain element X, has a mass of \( 2.5 \times 10^{-22} \text{ g} \). Find the relative atomic mass of X  (\( L = 6.0 \times 10^{23} \text{ mol}^{-1} \))

(ii) Molar Mass
Molar mass is the mass of one mole of a substance. Sometimes, it is called the relative formula mass. In order to calculate the molar mass of a compound, we must first write a correct formula. To write the correct formula, we must remember the following:
- Symbols of elements and formulae of radicals
- Valencies of elements and radicals
  Some of these are listed in Table 2.9

<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Valency</th>
<th>Name</th>
<th>Symbol</th>
<th>Valency</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>H^{+}</td>
<td>1</td>
<td>Iron(III)</td>
<td>Fe^{3+}</td>
<td>3</td>
</tr>
<tr>
<td>Potassium</td>
<td>K^{+}</td>
<td>1</td>
<td>Aluminium</td>
<td>Al^{3+}</td>
<td>3</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na^{+}</td>
<td>1</td>
<td>Ammonium</td>
<td>NH_{4}^{+}</td>
<td>1</td>
</tr>
<tr>
<td>Silver</td>
<td>Ag^{+}</td>
<td>1</td>
<td>Nitrate</td>
<td>NO_{3}^{-}</td>
<td>1</td>
</tr>
<tr>
<td>Barium</td>
<td>Ba^{2+}</td>
<td>2</td>
<td>Hydrogensulphate</td>
<td>HSO_{4}^{-}</td>
<td>1</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca^{2+}</td>
<td>2</td>
<td>Carbonate</td>
<td>CO_{2}^{2-}</td>
<td>2</td>
</tr>
<tr>
<td>Copper (II)</td>
<td>Cu^{2+}</td>
<td>2</td>
<td>Sulphite</td>
<td>SO_{3}^{2-}</td>
<td>2</td>
</tr>
<tr>
<td>Iron (II)</td>
<td>Fe^{2+}</td>
<td>2</td>
<td>Sulphate</td>
<td>SO_{4}^{2-}</td>
<td>2</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg^{2+}</td>
<td>2</td>
<td>Phosphat</td>
<td>PO_{4}^{3-}</td>
<td>3</td>
</tr>
<tr>
<td>Zinc</td>
<td>Zn^{2+}</td>
<td>2</td>
<td>Hydroxide</td>
<td>OH^{-}</td>
<td>1</td>
</tr>
</tbody>
</table>
Once we write the correct formula of the compound, we find its relative formula mass by adding up the relative atomic masses of the atoms or ions in the formula.

Let us now consider the following examples:-

**Example 1**

1. Find the Relative Formula Mass (R.F.M) of the following compounds.
   (a) Calcium chloride
   (b) Copper(II) hydroxide.

<table>
<thead>
<tr>
<th>Relative atomic masses</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ca = 40</td>
</tr>
<tr>
<td>Cl = 35.5</td>
</tr>
<tr>
<td>Cu = 63.5</td>
</tr>
<tr>
<td>O = 16</td>
</tr>
<tr>
<td>H = 1</td>
</tr>
</tbody>
</table>

**Solution**

First write the correct formula of the compounds. Secondly add up their Relative atomic masses.

(a) Each formula of calcium chloride contains 1 calcium ion and 2 chloride ions.

\[
\text{CaCl}_2 \\
40 + (35.5 \times 2) = 111
\]

So, the R.F.M for calcium chloride = 111

(b) The formula of copper(II) hydroxide contains 1 copper ion, 2 oxide ions and 2 hydrogen ions. i.e.

\[
\text{Cu(OH)}_2 \quad \text{(Note: You have to double whatever is inside the brackets).}
\]

\[
64 + (16 \times 2) + (1 \times 2) = 98
\]

So, the R.F.M of copper(II) hydroxide = 98

The relative formula mass of some molecules, for example, water, diatomic gases like hydrogen, oxygen, nitrogen, chlorine and others are referred to as Relative Molecular Mass (R.M.M) because they exist as simple molecules.

**Note:** Sometimes the R.M.M is abbreviated as M_r.
Example 2
What is the relative molecular mass of:
(a) one molecule of water?
(b) one molecule of oxygen?

Solution
First, we must write the correct formula of water and oxygen molecules.
(a) The formula of water is $\text{H}_2\text{O}$.

Then we add up the relative atomic masses of the elements that make up water.

$$\text{H}_2\text{O}$$

$$(1 \times 2) + 16 = 18$$

Therefore the relative molecular mass (R.M.M) of water is 18.

(b) Oxygen is a diatomic gas whose formula is $\text{O}_2$.
Therefore the R.M.M of oxygen is $16 \times 2 = 32$

Fig. 2.7: Some common diatomic gases

Example 3
What is the mass of one mole of
(a) Magnesium sulphate
(b) Calcium carbonate
Write the correct formula of magnesium sulphate then get the relative formula mass by adding up all the relative atomic masses of the atoms in the formula.

\[
\text{R. F. M of MgSO}_4 = 24.0 + 32.0 + (16.0 \times 4) = 120.0
\]

But the problem is asking us to find the mass of 1 mole. Therefore the mass of 1 mole of MgSO\(_4\) = 120 g

(b) The formula of calcium carbonate is CaCO\(_3\).

Therefore the R.F.M of CaCO\(_3\) = 40 + 12 + (16 \times 3) = 100
The molar mass of CaCO\(_3\) = 100 g

**Example 4**

Find the mass of 0.2 moles of calcium carbonate (Ca = 40, C = 12, O = 16)
The relative molecular mass of CaCO\(_3\) = 40 + 12 + 48 = 100

\[
1 \text{ mole of CaCO}_3 = 100 \text{ g}
\]
Therefore \[
0.2 \text{ moles of CaCO}_3 = x \text{ g.}
\]
Let us cross multiply and solve for \(x\)

\[
x = \frac{0.2 \text{ moles} \times 100 \text{ g}}{1 \text{ mole}} = 20 \text{ g}
\]

**Example 5**

What is the mass of:
(a) 0.25 moles of ammonium carbonate?
(b) 0.5 moles of copper(II) oxide?
(R.A.M of N = 14.0; H = 1.0; C = 12.0; Cu=64.0; O = 16.0)

(a) 1 mole of \((\text{NH}_4)_2\text{CO}_3\) \[
= (14 + 4) 2 + 12 + 48 = 96 \text{ g}
\]
0.25 moles of \((\text{NH}_4)_2\text{CO}_3\) \[
= x \text{ g}
\]
Let us cross multiply and solve for \(x\)

\[
x = \frac{0.25 \text{ moles} \times 96 \text{ g}}{1 \text{ mole}} = 24 \text{ g}
\]
(b) 1 mole of CuO \((64 + 16)\) = 80 g  
0.5 moles of CuO = xg  
Let us cross multiply and solve for x  
\[
x = \frac{0.5 \text{ moles} \times 80 \text{g}}{1 \text{ mole}} = 40 \text{ g}
\]

Exercise 2.2
1. Write down the formula of the following compounds:
   (i) Copper(II) nitrate  
   (ii) Iron(II) sulphate  
   (iii) Ammonium phosphate  
   (iv) Sodium sulphate  
   (v) Copper(II) sulphate crystals  
   (vi) Magnesium hydrogen carbonate

2. Find the molar mass of the following compounds:
   (i) \(\text{Fe(OH)}_2\)  
   (ii) \(\text{CaSO}_4\)  
   (iii) \((\text{NH}_4)_2\text{SO}_4\)  
   (iv) \(\text{Cu}_2\text{O}\)  
   (v) \(\text{Al}_2(\text{SO}_4)_3\)  
   (vi) \(\text{NaNO}_3\)

3. Find the formula mass of the following:
   (i) \(\text{C}_6\text{H}_12\text{O}_6\)  
   (ii) \(\text{CH}_3\text{CH}_2\text{OH}\)  
   (iii) \(\text{NH}_4\text{NO}_3\)  
   (iv) \(\text{Fe}_2\text{O}_3\)  
   (v) \(\text{CaCl}_2\)

4. Find the mass of:-
   (a) 2 moles of lead(IV) oxide  
   (b) 0.01 moles copper(II) carbonate  
   (c) 0.5 moles carbon(IV) oxide  
   (d) 3 moles water.

Example 6
How do we calculate the mass of one atom?  
Calculate the mass of 1 atom of copper. \((\text{Cu} = 63.5; \ L = 6.0 \times 10^{23})\)

\[
\text{1 mole of atoms of Cu} = 63.5 \text{ g}
\]
\[
\therefore \text{Mass of 1 atom of Cu} = x \text{ g}
\]
Let us cross multiply and solve for x  
\[
x \text{ g} \times 6.0 \times 10^{23} \text{ atoms} = 63.5 \text{ g}
\]
\[
x = \frac{63.5 \text{ g}}{6.0 \times 10^{23} \text{ atoms}}
\]
\[
\therefore 1 \text{ atom of copper has mass} = 1.06 \times 10^{-22} \text{ g}
\]
(iii) **Conversion of mass in grams to moles**

Usually conversion of a given mass of an element into moles is quite easy as long as we write the right formula. Let us now study the following examples to find out how this can be done.

**Examples 1**

Convert the masses of the following substances to moles.

(a) 0.12 g of carbon atoms
(b) 0.24 g of magnesium atoms
(c) 0.16 g of oxygen molecules
(d) 0.8 g of oxygen atoms

\(C = 12, \text{Mg} = 24, \text{O} = 16\)

**Solution**

**First principle method**

(a) 12 g of carbon atoms contain 1 mole.
\[\therefore 0.12 \text{ g of carbon contain } x \text{ mole}\]
Cross multiply and solve for \(x\).
\[12 \text{ g} \times x \text{ mol} = 0.12 \text{ g} \times 1 \text{ mole}\]
\[\therefore x = \frac{0.12 \text{ g} \times 1 \text{ mole}}{12 \text{ g}}\]
\[= 0.01 \text{ moles}\]

\[\therefore 0.12 \text{ g carbon atom contain 0.01 moles}\]

**Formula method**

\[
\begin{align*}
\text{moles} & = \frac{\text{Mass}}{\text{R.A.M}} \\
& = \frac{0.12}{12} \\
& = 0.01 \text{ moles}
\end{align*}
\]

**First principle method**

(b) 24 g of magnesium contains 1 mole
\[\therefore 0.24 \text{ g magnesium contains } y \text{ mole}\]
Again cross multiply and solve for \(y\).
\[\therefore y \text{ moles} = \frac{0.24}{24}\]
\[= 0.01 \text{ moles}\]
Try (c) and (d) using formula method.

### Formula method

\[
\text{moles} = \frac{\text{Mass}}{\text{R.A.M}}
\]

\[
= \frac{0.24}{24}
\]

\[
= 0.01 \text{ moles}
\]

(c) Oxygen is a diatomic gas. 1 molecule of oxygen contains two atoms i.e \( \text{O}_2 \).

The mass of 1 mole of oxygen atoms (O) is \( 16 \text{ g} \)

The mass of 1 mole of oxygen molecules, \( \text{O}_2 \), = \( 16 \times 2 = 32 \text{ g} \)

\[
\therefore \quad 32 \text{ g of oxygen contain 1 mole of oxygen (O}_2\text{) molecules}
\]

0.16 g of oxygen contain \( x \) moles of \( \text{O}_2 \) molecules

Cross multiply and solve for \( x \)

\[
\begin{align*}
32 \text{ g} & \times x \text{ moles} = 0.16 \text{ g} \times 1 \text{ mole} \\
32 \text{ g} & \times x \text{ moles} = \frac{0.16 \text{ g} \times 1 \text{ mole}}{32 \text{ g}} \\
\therefore \quad x \text{ mole} & = 0.005 \text{ moles}.
\end{align*}
\]

(d) 16 g of oxygen atoms (O) contain 1 mole

0.8 g oxygen atoms contain \( y \) moles

Cross multiply and solve for \( y \)

\[
\begin{align*}
16 \text{ g} & \times y \text{ moles} = 0.8 \text{ g} \times 1 \text{ mol} \\
\therefore \quad y \text{ mole} & = \frac{0.8 \text{ g} \times 1 \text{ moles}}{16 \text{ g}} \\
\therefore \quad y \text{ mole} & = 0.05 \text{ moles}
\end{align*}
\]

It is very easy to convert grams to moles. Remember in Chemistry, you do not just write numbers only except for R.A.Ms and R.M.Ms. Once again use statements, before you solve for your unknown. \textit{In chemistry it is wrong to do the following even if the answer is correct.}

\[
\begin{align*}
16 \text{ y} & = 0.8 \\
y & = \frac{0.8}{16} = 0.05
\end{align*}
\]

Try to use first principle method since you do not have to memorise many formulae.
Example 2
Dear sister,
I noticed on your certificate that you did well in Chemistry. Help! I can convert grams of elements to moles quite comfortably. Can you tell me how I can convert the following masses of ions to moles. I’ve quoted R.A.Ms for you in case you have forgotten.
(a) 71 g of chloride (Cl\(^-\)) ions
(b) 20 g of Ca\(^{2+}\) ions
(c) 0.56 g of Fe\(^{3+}\) ions.

(R.A.M \ Cl = 35.5, \ Ca = 40, \ Fe = 56)
Yours

Ionic brother

Solution
Easy! Just as you did with the elements. Use the first principle method. Remember to start with statements.

(a) The mass of 1 mole of Cl\(^-\) ions is 35.5 g
∴ the mass of x moles of Cl\(^-\) ions is 71 g
Cross multiply and solve for x

\[
\text{35.5 g} \times x \text{ moles} = 71 \text{ g} \times 1 \text{ mole}
\]
∴ \(x \text{ moles} = \frac{71 \text{ g} \times 1 \text{ mole}}{35.5 \text{ g}} = 2 \text{ moles}

(b) The mass of 1 mole of Ca\(^{2+}\) ions is 40 g
∴ The mass of y moles of Ca\(^{2+}\) ions is 20 g
Cross multiply solve for y

\[
y \text{ moles} \times 40 \text{ g} = 20 \text{ g} \times 1 \text{ mole}
\]
∴ \(y \text{ moles} = \frac{20 \text{ g} \times 1 \text{ mole}}{40 \text{ g}} = 0.5 \text{ moles}\)
The mass of 1 mole of Fe$^{3+}$ ions is 56 g.

∴ The mass of x mol of Fe$^{3+}$ is 0.56 g

Cross multiply and solve for x

\[
x \text{ moles} \times 56 \text{ g} = 0.56 \text{ g} \times 1 \text{ mole}
\]

\[
x = \frac{0.56 \text{ g} \times 1 \text{ mole}}{56 \text{ g}} = 0.01 \text{ moles}
\]

**Note:** Remember the mass of an ion is equal to the R.A.M of the element.

Your sister

Linda

Now repeat a, b and c using formula method.

**Exercises 2.3**

1. How many moles of atoms are present in:
   (a) 3.9 g of potassium.
   (b) 0.8 g of sulphur.
   (c) 0.8 g of Aluminium.
   (d) 2.8 g of iron.
   (e) 71 g of chlorine molecules.

2. Calculate the number of moles of atoms in the following.
   (a) 4 g of oxygen molecules.
   (b) 30 g of helium.
   (c) 28 g of iron.
   (d) 14 g of nitrogen atoms.

3. Calculate the number of moles of ions in each of the following.
   (a) 14 g of iron(II) ions.
   (b) 48 g of magnesium ions.
   (c) 13 g of potassium ions.
   (d) 4000 g of calcium ions.

(R.A.Ms at the back of the book)

(iv) **Conversion of mass (in grams) of compounds to moles**

**Examples 1**

Suppose you have:
   (a) 32 g of sulphur(IV) oxide
   (b) 15 g of aluminum sulphate
   (c) 3.92 g of hydrated ammonium iron(II) sulphate

How many moles of each compound do you have?
Solution

**First principle method**

First write the correct formula of the compound and work out the mass of 1 mole also known as molar mass.

(a) $1$ mole of sulphur(IV) oxide, $\text{SO}_2$ = $32 + (16 \times 2) = 64 \text{ g}$

$x$ moles of sulphur(IV) oxide, $\text{SO}_2$ = $32 \text{ g}$

$\therefore$ $x$ moles = \(\frac{32 \text{ moles} \times 1 \text{ mole}}{64 \text{ moles}}\)

= $0.5$ moles

(b) $1$ mole of $\text{Al}_2(\text{SO}_4)_3$ = $(27 \times 2) + (32 \times 3) + 3(16 \times 4) = 342 \text{ g}$

$x$ g mole of $\text{Al}_2(\text{SO}_4)_3$ = $15 \text{ g}$

$\therefore$ $x$ moles = \(\frac{15 \text{ g} \times 1 \text{ mole}}{342 \text{ g}}\)

= $0.044$ moles

(c) $1$ mole of $(\text{NH}_4)_2 \text{Fe(SO}_4)_2 \cdot 6\text{H}_2\text{O}$ = $392 \text{ g}$

$x$ mole of $(\text{NH}_4)_2 \text{Fe(SO}_4)_2 \cdot 6\text{H}_2\text{O}$ = $3.92 \text{ g}$

$\therefore$ $x$ moles = \(\frac{3.92 \text{ g} \times 1 \text{ moles}}{392 \text{ g}}\)

= $0.01$ moles

---

**Formula method**

(a) moles = \(\frac{\text{Mass}}{\text{R.F.M.}}\)

= \(\frac{32 \text{ g}}{\text{R.F.M. of SO}_2}\)

= \(\frac{32}{64}\)

= $0.5$ moles

(b) moles = \(\frac{15 \text{ g}}{\text{R.F.M. of Al}_2(\text{SO}_4)_3}\)

= \(\frac{15}{342}\)

= $0.044$ moles
(c) moles \[= \frac{\text{Mass}}{\text{R.F.M}} = \frac{3.92 \text{ g}}{(\text{NH}_4)_2 \text{Fe(SO}_4)_2 \cdot 6\text{H}_2\text{O}}\]
\[= \frac{3.92}{392 \text{ g}}\]
\[= 0.01 \text{ moles}\]

**Exercise 2.4**

How many moles are in:

(a) 2 g of sodium hydroxide.
(b) 0.96 g of ammonium carbonate.
(c) 126 g of calcium sulphate.
(d) 0.106 g of sodium carbonate.
(e) 31 g of aluminium sulphate.
(f) 9.9 g of zinc hydroxide.
(g) 0.17 g of ammonia.
(h) 13.2 g of carbon(IV) oxide.

**Conversion of moles to mass in grams**

It is also easy to convert moles to grams. We will again cross multiply and solve for the unknown value. Let us start with a general problem that we can easily solve. Remember it becomes even easier when we use statements.

**General example**

One dozen contains 12 items. How many items are in 12 dozens?

**Answer**

1 dozen contain 12 items
\[\therefore 12 \text{ dozens contain } x \text{ items.}\]
Cross multiply and solve for \(x\).
\[x \times 1 \text{ dozen } = (12 \times 12) \text{ items}\]
\[\therefore x \text{ items } = 144 \text{ items}\]

There are 144 items in 12 dozens. Compare the following examples with this one.

**Example 1**

A student bought \(x\) g of charcoal (carbon).

This was the same as 0.1 mole. Find the value of \(x\). (\(C = 12\)).

**First principle method**

**Solution**

1 mole of carbon atoms \[= 12 \text{ g}\]
\[\therefore 0.1 \text{ moles of carbon atoms } = x \text{ g}\]
\[
1 \text{ mol} \times x \text{ g} = 0.1 \text{ mole} \times 12 \text{ g}
\]
\[
\therefore x = \frac{0.1 \text{ mole} \times 12 \text{ g}}{1 \text{ mole}}
\]
\[
= 1.2 \text{ g}
\]

**Formula method**

\[
\text{Moles} = \frac{\text{Mass}}{\text{R.A.M}}
\]

Mass = moles \times R.A.M

\[
\therefore \text{Mass of 0.1 mole of carbon} = 0.1 \times 12 = 1.2 \text{ g}.
\]

Try examples 2, 3 and 4 by formula method and see whether you get same answer.

**Example 2**
Find the mass of 0.5 mole of bromine atoms. (Br = 80)

**Solution**

The mass of 1 mole of Br atoms = 80 g

\[
\therefore \text{The mass of 0.5 mole of Br atoms} = x \text{ g}
\]

Cross multiply and solve for x

\[
1 \text{ mol} \times x \text{ g} = 0.5 \text{ moles} \times 80 \text{ g}
\]

\[
\therefore x \text{ moles} = \frac{0.5 \text{ moles} \times 80 \text{ g}}{1 \text{ mol}}
\]

\[
= 40 \text{ g}
\]

**Example 3**
Calculate the mass of \( \frac{1}{4} \) moles of zinc (Zn = 65)

The mass of 1 mole of zinc = 65 g

\[
\therefore \text{The mass of} \ \frac{1}{4} \ \text{moles of zinc} = x \text{ g}
\]

Cross multiply and solve for x

\[
1 \text{ mole} \times x \text{ g} = \frac{1}{4} \text{ mole} \times \frac{65 \text{ g}}{1 \text{ mole}}
\]

\[
\therefore x = 16.25 \text{ g}
\]
Example 4
Find the mass of:
(a) 2 moles of hydrogen atoms
(b) 20 moles of oxygen atoms
(H = 1, O = 16)

Solution
(a) The mass of 1 mole of H atoms = 1 g
   ∴ The mass of 2 moles of H atoms = y g
   Cross multiply and solve for y
   \[ y \text{ g} \times 1 \text{ mole} = 2 \text{ moles} \times 1\text{g} \]
   \[ y = 2 \text{ g} \]

(b) The mass of 1 mole of O atoms = 16 g
   ∴ the mass 20 moles of O atoms = x g
   Cross multiply and solve for x
   \[ 1 \text{ mole} \times x = 20 \text{ moles} \times 16 \text{ g} \]
   \[ x = 320 \text{ g} \]

Exercise 2.5
1. What is the mass of:
   (a) 0.01 moles of C atoms.
   (b) 5 moles of N atoms
   (c) 0.5 moles of F atoms.
   (d) \( \frac{1}{4} \) moles of Ca atoms.

2. What is the mass of \( \frac{1}{5} \) of a mole of atoms of the following elements:
   (a) helium
   (b) chlorine
   (c) manganese
   (d) aluminium

(vi) Conversion of moles of molecules to grams
Sometimes you may be required to convert moles of molecules to grams. Before you do that you must:
(i) write the correct formula of the compound or molecule,
(ii) add up the R.A.M of the atoms of elements in the compound or molecule i.e. find the formula mass.
(iii) express the number of moles in grams.
Example 1
How many grams are there in
(a) 9 moles of hydrogen gas?
(b) 4 moles of water?

Solution
First principle method
(a) The correct formula of hydrogen gas is H₂.
   R.A.M of H = 1
   1 mole of H₂ = 1 × 2 = 2 g
   9 moles of H₂ = x g
   ∴ x = 2 × 9
        = 18 g.

(b) The formula of water is H₂O.
   (R.A.M: H = 1, O = 16)
   Mass of H atoms, 2H = 1 × 2 = 2
   Mass of O atoms, = 1 × 16 = 16
   1 mole of H₂O, = 2 +16 = 18 g
   ∴ 4 moles of H₂O = x g
        x = 18 × 4
        = 72 g

Formula method
\[
\text{Moles} = \frac{\text{Mass}}{\text{R.F.M}}
\]
(a) Mass of H₂ = moles × R.F.M
    = 9 × 2
    = 18 g

(b) Mass of 4 moles of water = 4 × 18
    = 72 g
Try examples 2 and 3 using the formula method.

Example 2
Work out the number of grams in 0.5 mole of ethanol.
(R.A.M of) C = 12, H = 1, O = 16)
Solution
(i) The formula of ethanol is C\textsubscript{2}H\textsubscript{5}OH.
(ii) The formula mass of ethanol is as follows:

<table>
<thead>
<tr>
<th>Moles of atoms in formula</th>
<th>R.A.M</th>
<th>Formula mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 C</td>
<td>12</td>
<td>2 \times 12 = 24</td>
</tr>
<tr>
<td>6 H</td>
<td>1</td>
<td>6 \times 1 = 6</td>
</tr>
<tr>
<td>1 O</td>
<td>16</td>
<td>1 \times 16 = 16</td>
</tr>
</tbody>
</table>

Total formula mass of ethanol is 46 g

\[
1 \text{ mole of ethanol } C_2H_5OH = 46 \text{ g} \\
\therefore 0.5 \text{ moles of ethanol } C_2H_5OH = x \text{ g} \\
\text{Cross multiply and solve for } x \\
1 \text{ mol } \times x \text{ g} = 0.5 \times 46 \text{ g} \\
x = 23.0 \text{ g} \\
\therefore \text{ mass ethanol in 0.5 mole} = 23.0 \text{ g}
\]

Example 3
Calculate the mass (in grams) in \(\frac{1}{10}\) moles of magnesium nitrate.
(R.A.M.) of Mg = 24, N = 14, O = 16)

Solution
(i) The formula of magnesium nitrate is Mg(NO\textsubscript{3})\textsubscript{2}
(ii) Work out the formula of Mg(NO\textsubscript{3})\textsubscript{2} as follows:

<table>
<thead>
<tr>
<th>Moles of atoms in formula</th>
<th>R.A.M</th>
<th>Formula mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 Mg</td>
<td>24</td>
<td>1 \times 24 = 24</td>
</tr>
<tr>
<td>2 N</td>
<td>14</td>
<td>2 \times 14 = 28</td>
</tr>
<tr>
<td>6 O</td>
<td>16</td>
<td>6 \times 16 = 96</td>
</tr>
</tbody>
</table>

Total formula mass of magnesium nitrate = 148 g

This means that 1 mole of Mg(NO\textsubscript{3})\textsubscript{2} = 148 g
\[
\therefore \frac{1}{10} \text{ mole of Mg(NO}_3\text{)}_2 = x \text{ g} \\
\text{cross multiply and solve for } x \\
1 \text{ mol} \times x = \frac{1}{10} \times 148 \text{ g} \\
x = 14.8 \text{ g} \\
\therefore \text{ mass} = 14.8 \text{ g}
\]
Exercises 2.6
1. Calculate the mass of:
   (i) 0.5 moles of chlorine molecules.
   (ii) 0.25 moles of nitrogen molecules.
   (iii) 3 moles of iodine molecules.
   (iv) 0.01 moles of fluorine molecules.
2. How many grams are in each of the following compounds:
   (i) 0.1 moles of copper(II) oxide.
   (ii) 10 moles of ammonia.
   (iii) 0.2 moles of ammonium nitrate.
   (iv) 2.5 moles of calcium phosphate.
   (v) 2.5 moles of aluminum sulphate.
3. Calculate the mass in 5 moles of ethane, $\text{C}_2\text{H}_6$. 

(vii) Conversion of moles to number of particles in an element or compound
So far, we have learnt how to convert moles to mass in grams and mass in grams to moles. Now let us find out how many particles are in a given number of moles or in given mass in grams. Let us start with the particles in a mole.

(a) Particles in a number of moles
First let us start with familiar things by answering the following familiar questions. How many pencils are in:
   (i) 1 dozen?
   (ii) 2 dozens?
   (iii) 0.5 dozen?
   (1 dozen = 12)

Solution
Using the value given in brackets, let us calculate using our unknown value $x$.
   (i) 1 dozen contains 12 items
   (ii) 2 dozens contain $x$ items
   
   By cross multiplying and solving for $x$ we get
   
   \[1 \text{ dozen} \times x \text{ items} = 2 \text{ dozens} \times 12 \text{ items}\]

   \[\therefore x = \frac{2 \text{ dozens} \times 12 \text{ items}}{1 \text{ dozen}} \]

   \[= 12 \times 2 = 24 \text{ items}\]
(iii) 1 dozen contains 12 items

\[ \therefore 0.5 \text{ dozens contain } x \text{ items} \]

Cross multiply and solve for \(x\)

\[ \therefore x = \frac{0.5 \text{ dozen} \times 12 \text{ items}}{1 \text{ dozen}} = 6 \text{ items} \]

In Chemistry, we deal with tiny particles. Therefore instead of using dozen as earlier mentioned we use the mole. Using the mole let us follow similar procedure to solve the following problems.

**Example 1**

Find the number of atoms in

(i) 3 moles of iron

(ii) 0.25 moles of calcium

(iii) \(\frac{3}{4}\) moles of oxygen atoms

\((L = 6.0 \times 10^{23})\)

**Solution**

(i) 1 mole of Fe contains \(6.0 \times 10^{23}\) atoms

3 moles of Fe contain \(x\) atoms

Cross multiply and solve for \(x\)

\[ \therefore x \text{ atoms} = \frac{3 \text{ moles} \times 6.0 \times 10^{23} \text{ atoms}}{1 \text{ mole}} = 18.0 \times 10^{23} \text{ atoms} = 1.8 \times 10^{24} \text{ atoms} \]

**Note:** We multiply the ‘6’ part only. Do not multiply with the exponential i.e. ‘\(10^{23}\)’ part.

(ii) 1 mole of Calcium contains \(6.0 \times 10^{23}\) atoms

\[ \therefore 0.25 \text{ moles of Ca contain } x \text{ atoms} \]

Cross multiply and solve for \(x\)

\[ x \text{ atoms} = \frac{0.25 \text{ moles} \times 6.0 \times 10^{23} \text{ atoms}}{1 \text{ mole}} = 1.5 \times 10^{23} \text{ atoms} \]

(iii) 1 mole of oxygen \(= 6.0 \times 10^{23}\) atoms

\[ \frac{1}{4} \text{ of a moles of oxygen } = x \text{ atoms} \]
Cross multiply and solve for x
\[ x \text{ atoms} = \frac{1}{4} \text{ moles} \times \frac{6.0 \times 10^{23}}{1 \text{ mole}} \]
\[ = 1.5 \times 10^{23} \text{ atoms} \]

Example 2
Find the number of atoms in the following:
(i) 0.2 moles of oxygen molecules.
(ii) 2 moles of sulphur(IV) oxide.
(iii) 0.5 moles of CaO. 
\( (L = 6.0 \times 10^{23}) \)

Solution
(i) 1 mole of oxygen molecules, \((O_2)\), contain 2 moles of O atoms
\[ \therefore 0.2 \text{ mole of oxygen molecules, } (O_2), \text{ contain } x \text{ moles of O atoms} \]
Let us cross multiply and solve for \(x\)
\[ x = 0.2 \times 2 = 0.4 \text{ moles of O atoms} \]
But 1 mole contains \(6.0 \times 10^{23}\) O atoms
\[ \therefore 0.4 \text{ moles contain } y \text{ O atoms} \]
Cross multiply and solve for \(y\)
\[ y = 0.4 \times 6.0 \times 10^{23} \text{ atoms} \]
\[ = 2.4 \times 10^{23} \text{ atoms} \]

(ii) 1 mole of sulphur(IV) oxide, \(SO_2\) contains 3 moles of atoms
(one for sulphur, two for oxygen)
\[ \therefore 2 \text{ moles of } SO_2 = 2 \times 3 = 6 \text{ moles of atoms} \]
But 1 mole \[ = 6.0 \times 10^{23} \text{ atoms} \]
\[ \therefore 6 \text{ moles } = x \text{ atoms} \]
Cross multiply and solve for \(x\)
\[ x = 6 \times 6.0 \times 10^{23} \]
\[ = 36.0 \times 10^{23} \]
\[ = 3.6 \times 10^{24} \text{ atoms} \]

(iii) 1 mole of CaO contain 2 moles of atoms (one Ca and one O)
\[ \therefore 0.5 \text{ moles contain } x \text{ moles} \]
Cross multiply and solve for \(x\)
\[ x = 0.5 \times 2 \]
\[ = 1.0 \text{ mole} \]
1 mole will contain \[ = 1.0 \times 6.0 \times 10^{23} \text{ atoms} \]
\[ = 6.0 \times 10^{23} \text{ atoms} \]
Example 3

(i) How many magnesium ions are in 1 mole of magnesium chloride?

(ii) How many chloride ions are there in 0.5 mole of magnesium chloride?

\[(L = 6.0 \times 10^{23})\]

Solution

The formula of magnesium chloride is, \(\text{MgCl}_2\). Ions are set free when the salt is dissolved or melted.

(i) \[
\text{MgCl}_2(aq) \rightarrow \text{Mg}^{2+}(aq) + 2\text{Cl}^-(aq) \\
1 \text{ mole} \rightarrow 1 \text{ mole} \quad 2 \text{ moles}
\]

1 mole of magnesium chloride when dissolved in water produces 1 mole of \(\text{Mg}^{2+}\) ions and 2 moles of \(\text{Cl}^-\) ions.

But 1 mole \(= 6.0 \times 10^{23}\) \(\text{Mg}^{2+}\) ions

1 mole of \(\text{Mg}^{2+}\) ions \(= x\) ions

Cross multiply and solve for \(x\)

\[
1 \text{ mole} = 1 \times 6.0 \times 10^{23}\text{ ions} \\
\therefore x = 6.0 \times 10^{23} \text{ ions} \\
= 6.0 \times 10^{23}\text{ Mg}^{2+}\text{ions}
\]

(ii) \[
\text{MgCl}_2(aq) \rightarrow \text{Mg}^{2+}(aq) + 2\text{Cl}^-(aq) \\
0.5 \text{ moles} \rightarrow 0.5 \text{ mole} \quad 1 \text{ mole}
\]

0.5 mole of \(\text{MgCl}_2\) produce 1 mole of \(\text{Cl}^-\) ions

But 1 mole \(= 6.0 \times 10^{23}\) \(\text{Cl}^-\) ions

1 mole \(= y\) \(\text{Cl}^-\) ions

Cross multiply and solve for \(y\)

\[
1 \text{ moles} \times y = 1 \times 6.0 \times 10^{23}\text{ Cl ions} \\
\therefore y = 6.0 \times 10^{23}\text{ Cl}^-\text{ions}
\]

Exercise 2.7

1. Find the number of atoms in:

   (a) \(\frac{1}{2}\) mole of sodium.

   (b) 0.16 g of oxygen molecules.

   (c) 19.5 g of potassium.

2. How many iron(II) ions and chloride ions are in aqueous solutions of the following:

   (a) 1 mole of iron(II) chloride.

   (b) 0.5 mole iron(II) chloride.

   (c) 81.25 g of iron(II) chloride.
(viii) **Conversion of mass in grams to number of particles in an element or compound**
From a given number of grams of an element or compound, we may also be required to work out the number of particles in that given mass. Study the following examples.

**Example 1**
Calculate the number of atoms in 12 g of carbon.
\( (C = 12, \text{L} = 6.0 \times 10^{23}) \)

**Solution**

**First principle method**
Always, as mentioned earlier, the values in the brackets are given to be used. Therefore we must use them! Let us start with \( C = 12 \).

\[
\begin{align*}
1 \text{ mole of carbon atoms} &= 12 \text{ g} \\
x \text{ mole of carbon atoms} &= 12 \text{ g} \\
\text{Cross multiply and solve for } x &= 1 \times 12 \text{ g} \\
x \text{ mole} &= \frac{12 \text{ g}}{12 \text{ g}} = 1 \text{ mole} \\
\text{But, 1 mole of carbon will have} &= 1 \times 6.0 \times 10^{23} \text{ atoms.} \\
&= 6.0 \times 10^{23} \text{ atoms}
\end{align*}
\]

**OR**
We can say, since 1 mole contain \( 6.0 \times 10^{23} \) atoms

\[
\begin{align*}
\text{12 g of carbon} &= 6.0 \times 10^{23} \text{ atoms} \\
\therefore \text{The number of atoms in 12 g of carbon} &= x \text{ atoms}
\end{align*}
\]

Cross multiply and solve for \( x \) = \( \frac{12 \text{ g} \times 6.0 \times 10^{23} \text{ atoms}}{12 \text{ g}} \)

\[
x = 6.0 \times 10^{23} \text{ atoms}
\]

**Formula method**

\[
\text{Moles} = \frac{\text{Mass}}{\text{R.A.M}}
\]

\[
= \frac{12}{12} = 1 \text{ mole}
\]

The number of atoms in one mole = \( 1 \times 6.0 \times 10^{23} \) = \( 6.0 \times 10^{23} \)

Try example 2 using the formula method.
Examples 2
How many atoms are in 20 g of calcium
\((Ca = 40, \text{L} = 6.0 \times 10^{23})\)

\[
\begin{align*}
1 \text{ mole of Ca atoms} &= 40 \text{ g} \\
x \text{ moles of Ca atoms} &= 20 \text{ g} \\
\text{Cross multiply and solve for } x \text{ to get moles} \\
\therefore x &= \frac{20}{40} = 0.5 \text{ moles} \\
1 \text{ mole of Ca atoms} &= 6.0 \times 10^{23} \text{ atoms} \\
\therefore 0.5 \text{ mole of Ca atoms} &= 0.5 \times 6.0 \times 10^{23} \text{ atoms} \\
&= 3.0 \times 10^{23} \text{ atoms}.
\end{align*}
\]

OR

\[
\begin{align*}
40 \text{ g of Ca contain } 6.0 \times 10^{23} \text{ atoms} \\
20 \text{ g of Ca contain } y \text{ atoms} \\
\text{Cross multiply and solve for } y = \frac{20 \times 6.0 \times 10^{23} \text{ atoms}}{40 \text{ g}} \\
\therefore y &= 3.0 \times 10^{23} \text{ atoms}
\end{align*}
\]

Exercise 2.8
1. How many atoms are there in:
   (a) 1 mole of ammonia.
   (b) 20 g of sodium hydroxide.
2. How many zinc ions and nitrate ions are in 0.5 mole of zinc nitrate solution?
3. Calculate the number of sulphate ions in 0.2 mole ammonium sulphate solution.
4. Find the number of moles of \(Ca^{2+}\) and \(PO_4^{3-}\) ions in 310 g of calcium phosphate.
5. How many copper(II) ions and chloride ions are in 2 moles of copper(II) chloride?

2.4: Determination of formulae of compounds
We have seen that we can write the correct chemical formula of a compound by considering the correct symbols and valencies. A correct formula gives us the ratio of moles of the elements in that compound. For example, the formula of calcium carbonate is \(CaCO_3\). This formula carries the following information, 1 mole of calcium carbonate contains, 1 mole of Ca atoms, 1 mole of C atoms and 3 moles of oxygen atoms.

Example 1
We can calculate the percentage by mass of each element in the compound as follows:

The formula mass of \(CaCO_3\) = 40 + 12 + 48 = 100
∴ The % of 1 mole of Ca = \( \frac{\text{Ca}}{\text{CaCO}_3} \times 100 = \frac{40}{100} \times 100 = 40\% \)

The % of 1 mole of C = \( \frac{\text{C}}{\text{CaCO}_3} \times 100 = \frac{12}{100} \times 100 = 12\% \)

The % of 3 moles O = \( \frac{3(\text{O})}{\text{CaCO}_3} \times 100 = \frac{3 \times 16}{100} \times 100 = 48\% \)

Example 2
Calculate the percentage of water of crystallisation in copper(II) sulphate crystals
(R.A.M Cu = 63.5, S = 32, O = 16, H = 1)

Solution
The formula of copper(II) sulphate crystals is CuSO\(_4\) \( \cdot \) 5H\(_2\)O.
The formula mass is obtained by adding all the R.A.M e.g.
- 1 mole of Cu atoms = 63.5 g
- 1 mole of S atoms = 32 g
- 4 moles of O atoms (4 \times 16) = 64 g
- 5 moles of H\(_2\)O molecules 5(2 + 16) = 90 g
∴ 1 mole of CuSO\(_4\) \( \cdot \) 5H\(_2\)O crystals = 249.5g

The % of water of crystallisation (5H\(_2\)O) = \( \frac{5\text{H}_2\text{O}}{\text{CuSO}_4 \cdot 5\text{H}_2\text{O}} \times 100 \)
= \( \frac{90}{249.5} \times 100 = 36.07\% \)
= 36.1%

Exercise 2.9
1. Calculate the percentage by mass of each element in each of the following compounds.
   (a) Iron(II) oxide, (Fe\(_2\)O\(_3\))
   (b) Sodium hydroxide, (NaOH)
   (c) Aluminum sulphate, Al\(_2\)(SO\(_4\))\(_3\)
   (d) Magnesium oxide, (MgO).

2. Calculate the percentage by mass of the following compounds:
   (a) Carbon in ethanol, (C\(_2\)H\(_5\)OH)
   (b) Chromium in potassium dichromate(VI), (K\(_2\)Cr\(_2\)O\(_7\))
   (c) Chlorine in Iron(II) chloride, (FeCl\(_2\)).
   (d) Oxygen in copper(II) sulphate, (CuSO\(_4\)).
3. Write the formulae of the following compounds and calculate the percentage composition by mass of nitrogen in each one of them.
   (a) Ammonium sulphate.
   (b) Lead(II) nitrate.
   (c) Potassium nitrate.
   (d) Ammonium carbonate.

4. Calculate the percentage of water of crystallisation in iron(II) sulphate – 7 – water.

Number of moles of a compound can be used to calculate its formula. This is possible if we can get or if we are given the mass of each element in that compound. This mass is usually in grams. We can also find the formula of a compound if we know the percentage composition by mass of the elements that constitute the compound. We have already learnt how to calculate the percentage of elements in a compound.

Let us see how we can find formulae of compounds.

**Formulae of compounds**
As mentioned earlier, we have already learnt how to write formulae of compounds using symbols and valencies. For example, water is represented as:

$$\text{H}_2\text{O}$$

What does this formula mean?
1 oxygen atom combines with 2 atoms of hydrogen. It implies that 1 mole of oxygen atoms combines with 2 moles of hydrogen atoms.

Given the R.A.M of H = 1, O = 16, we can say that 2 g of hydrogen combine with 16 g of oxygen atoms to form 1 molecule of water.

Therefore the formula shows:
- The ratio of moles of atoms that constitute the compound.
- The quantity in grams of the element in the compound.

So, if we know the number of grams of each element that form the compound, using the relative atomic masses, we can change grams to moles. This will give us the ratio of moles in which the atoms combine and consequently enable us to write down a formula. Such a formula is usually referred to as the **empirical formula (E.F)**.

**Empirical formula (E.F)**
We are all aware of the fact that when masons construct stone houses, they must mix cement, sand and water in certain proportions.
From the Fig. 2.8 above when the sand, cement and water are not mixed in correct proportions (ratios), the walls of the building are bound to crack Fig. 2.8(b) and finally collapse. This is analogous to the composition of atoms in a compound. The atoms in the compound are combined in certain fixed ratios for them to remain stable, otherwise such compounds would be unstable hence would not exist. Now, let us consider the following experiment.

**Experiment 2.2**
To determine the empirical formula (E.F) of magnesium oxide

**Apparatus and chemicals**
- crucible and lid
- magnesium ribbon
- balance
- bunsen burner
- emery paper/steel wool
- tripod stand

**Procedure**
1. Weigh an empty crucible plus the lid and record the mass.
2. Clean about 15 cm of magnesium ribbon thoroughly using emery paper or steel wool. What do you observe? Why is it necessary to clean magnesium ribbon?
3. Coil up the magnesium ribbon and place it into the crucible.
4. Weigh the crucible and the lid again and record this mass.
5. Heat gently for a few seconds followed by strong heating for just a few minutes as shown in Fig. 2.9 below.

**Fig. 2.8: An illustration to explain proportions**

From the Fig. 2.8 above when the sand, cement and water are not mixed in correct proportions (ratios), the walls of the building are bound to crack Fig. 2.8(b) and finally collapse. This is analogous to the composition of atoms in a compound. The atoms in the compound are combined in certain fixed ratios for them to remain stable, otherwise such compounds would be unstable hence would not exist.

**Experiment 2.2**
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1. Weigh an empty crucible plus the lid and record the mass.
2. Clean about 15 cm of magnesium ribbon thoroughly using emery paper or steel wool. What do you observe? Why is it necessary to clean magnesium ribbon?
3. Coil up the magnesium ribbon and place it into the crucible.
4. Weigh the crucible and the lid again and record this mass.
5. Heat gently for a few seconds followed by strong heating for just a few minutes as shown in Fig. 2.9 below.

**Fig. 2.9: Heating magnesium ribbon in a crucible**
6. Remove the bunsen burner, lift the lid slightly to let in air and quickly replace the lid to avoid loss of the white smoke. The white smoke is the magnesium oxide (MgO).
7. Repeat procedure number 6 several times until no more magnesium flares.
8. Remove the lid and heat strongly.
9. Stop heating and allow the crucible plus its contents to cool and weigh them. Record this mass.
   - Can you suggest any experimental errors that may contribute to getting a wrong empirical formula?
   - What do you conclude from the result after re-weighing?

The following is a sample of the results and how we use the results to calculate the empirical formula.

<table>
<thead>
<tr>
<th>Mass of crucible + lid</th>
<th>= 25.00 g</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of crucible + lid + magnesium before heating</td>
<td>= 25.24 g</td>
</tr>
<tr>
<td>Mass of crucible + lid + magnesium oxide after heating</td>
<td>= 25.40 g</td>
</tr>
<tr>
<td>Mass of magnesium</td>
<td>= 0.24 g</td>
</tr>
<tr>
<td>Mass of crucible + lid + magnesium oxide after cooling</td>
<td>= 25.40 g</td>
</tr>
<tr>
<td>Mass of crucible + lid + magnesium before heating</td>
<td>= 25.24 g</td>
</tr>
</tbody>
</table>

∴ The mass of oxygen in the oxide = 0.16 g

(R.A.M Mg = 24, O = 16)

We therefore calculate the empirical formula of magnesium oxide as follows:

<table>
<thead>
<tr>
<th>Elements</th>
<th>Mg</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Composition by mass in g</td>
<td>0.24</td>
<td>0.16</td>
</tr>
<tr>
<td>Get moles reacting by dividing by the R.A.M</td>
<td>( \frac{0.24}{24} = 0.01 )</td>
<td>( \frac{0.16}{16} = 0.01 )</td>
</tr>
<tr>
<td>Divide by smaller number to get whole number moles</td>
<td>0.01</td>
<td>0.01</td>
</tr>
<tr>
<td>Mole ratio</td>
<td>1</td>
<td>1</td>
</tr>
</tbody>
</table>

This means that 1 mole of magnesium atoms combined with 1 mole of oxygen atoms to form the compound magnesium oxide. Therefore the empirical formula of magnesium oxide is MgO.

**Note**
Due to experimental errors, in weighing and the reaction of magnesium with nitrogen in the atmosphere forming magnesium nitride, there may be variation in the moles of magnesium and oxygen atoms.
We can see that 1 mole of magnesium atoms combines with 1 mole of oxygen atoms in the ratio of 1:1 to form magnesium oxide, MgO. 1:1 is the simplest whole-number ratio of the moles of atoms in the compound. Its formula is MgO. Such a formula is known as empirical formula (E.F).

The empirical formula of a compound shows the simplest whole-number ratio of the moles of atoms of the different elements in the compound.

Note: Empirical formula does not usually give exact number of the different atoms in one molecule of the compound.

Now let us find the empirical formula of copper(II) oxide.

**Experiment 2.3**
To determine the empirical formula of copper(II) oxide.

**Apparatus and chemicals**
- porcelain boat
- bunsen burner
- balance
- combustion tube
- copper(II) oxide
- source of dry hydrogen.

**Procedure**
1. Weigh a clean, dry empty porcelain boat.
2. Place a little copper(II) oxide into the boat and replace it on the balance. Add more copper(II) oxide a little at a time until the amount of copper(II) oxide is about 1.5 g.
3. Place the boat tube plus its contents into the combustion tube and arrange the apparatus as shown in Fig. 2.10.
4. Pass a stream of hydrogen before you start heating.
5. Heat the combustion tube at the point where you have placed the porcelain boat.

![Fig 2.10: Experimental set up to find the empirical formula of copper(II) oxide](image-url)
**Note:** The combustion tube must be clamped slanting so that the water formed does not flow back to the hot part of the combustion. This can cause the combustion tube to crack.

- A stream of hydrogen gas is passed through the combustion tube to drive out all air to avoid explosion of the oxygen-hydrogen mixture when ignited.
- When we stop heating, we must let a stream of hydrogen to continue passing through the combustion tube and its contents to avoid re-oxidation of the metal as it cools down.
- When the reaction is complete, black copper(II) oxide changes to a red-brown solid.

**Nature of results**

| Mass of empty porcelain boat | = ____ g |
| Mass of empty porcelain boat + copper(II) oxide | = ____ g |
| Mass of porcelain boat + copper after reaction | = ____ g |
| Mass of copper | = ____ g |
| Mass of oxygen | = ____ g |

**Conclusion**

- ____ g of copper(II) oxide used
- ____ g of copper formed
- ____ g of oxygen removed

*(O = 16, Cu = 63.5)*

**Using the results, complete the table below**

<table>
<thead>
<tr>
<th>Elements</th>
<th>Cu</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Composition by mass in (g)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Get moles reacting by dividing by R.A.M</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Divide by smaller number to get whole number moles</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mole ratio</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

The mole ratio of copper: oxygen is ____

∴ The empirical formula of copper(II) oxide is ____

Now compare your results with the following example.

Dear friend,

Our class did an experiment to find the E.F. of an oxide of copper. In our experiment we reduced copper(II) oxide with dry hydrogen.

I have quoted data for you as well as the R.A.Ms of Cu and O. Please assist.
Mass of empty boat = 25.0 g
Mass of empty boat + oxide of copper = 29.0 g
Mass of boat + copper (after reaction) = 28.2 g
(Cu = 64, O = 16)

Your, friend ……

Dear friend,

Proceed as follows to work out the empirical formula of that oxide of copper

Mass of Cu in g \(28.2 - 25.0 = 3.2 \text{ g}\)
Mass of oxygen \(29.0 - 28.2 = 0.8 \text{ g}\)

\[
\begin{array}{|c|c|c|}
\hline
\text{Elements} & \text{Cu} & \text{O} \\
\hline
\text{Composition by mass in g} & 3.2 & 0.8 \\
\hline
\text{Get reacting moles by dividing by R.A.M} & \frac{3.2}{64} = 0.05 & \frac{0.8}{16} = 0.05 \\
\hline
\text{Divide by the smaller value} & 0.05 = 1 & 0.05 = 1 \\
\hline
\text{Mole ratio} & 1 & 1 \\
\hline
\end{array}
\]

∴ The empirical formula is CuO
I hope this solution will be of assistance to you.

Yours sincerely,
Mole friend.

**Example 1**

When a sample of a hydrocarbon (X) was analysed, it was found to contain 12 g of carbon and 4 g of hydrogen. Work out the empirical formula of this hydrocarbon (R.A.M of C = 12, H = 1)

**Solution**

\[
\begin{array}{|c|c|c|}
\hline
\text{Elements} & \text{C} & \text{H} \\
\hline
\text{Composition by mass in g} & 12 & 4 \\
\hline
\text{Get reacting moles by dividing by R.A.M} & \frac{12}{12} = 1 & \frac{4}{4} = 4 \\
\hline
\text{Divide by the smallest value} & \frac{1}{1} = 1 & \frac{4}{1} = 4 \\
\hline
\text{Mole ratio} & 1 & 4 \\
\hline
\end{array}
\]

Therefore empirical formula (E.F) of the hydrocarbon (X) is CH₄
Example 2
When another compound containing carbon and hydrogen was analysed, it was found to contain 75% of carbon and 25% of hydrogen. Find its simplest formula. (R.A.M of C = 12 H = 1).

Solution

<table>
<thead>
<tr>
<th>Element</th>
<th>C</th>
<th>H</th>
</tr>
</thead>
<tbody>
<tr>
<td>Composition by mass in %</td>
<td>75</td>
<td>25</td>
</tr>
<tr>
<td>Get reacting moles by dividing by R.A.M</td>
<td>( \frac{75}{12} = 6.25 )</td>
<td>( \frac{25}{1} = 25 )</td>
</tr>
<tr>
<td>Divide by the smallest value</td>
<td>6.25 = 1</td>
<td>25 ( \div 6.25 = 4 )</td>
</tr>
<tr>
<td>Mole ratio</td>
<td>1</td>
<td>4</td>
</tr>
</tbody>
</table>

Therefore, the empirical formula is \( \text{CH}_4 \).

Example 3
A compound was found to contain 3.2 g of copper, 0.6 g of carbon and 2.4 g of oxygen. Find its empirical formula (Cu = 64, C = 12, O = 16).

Solution

<table>
<thead>
<tr>
<th>Element</th>
<th>Cu</th>
<th>C</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Composition by mass g</td>
<td>3.2</td>
<td>0.6</td>
<td>2.4</td>
</tr>
<tr>
<td>Get moles reacting by dividing by R.A.M</td>
<td>( \frac{3.2}{64} = 0.05 )</td>
<td>( \frac{0.6}{12} = 0.05 )</td>
<td>( \frac{2.4}{16} = 0.15 )</td>
</tr>
<tr>
<td>Divide by the smallest (to get whole number)</td>
<td>0.05 ( \div 0.05 = 1 )</td>
<td>0.05 ( \div 0.05 = 1 )</td>
<td>0.15 ( \div 0.05 = 3 )</td>
</tr>
<tr>
<td>Mole ratio</td>
<td>1</td>
<td>1</td>
<td>3</td>
</tr>
</tbody>
</table>

Therefore, the empirical formula of the compound is \( \text{CuCO}_3 \).

Example 4
14.82 g of a hydrocarbon contains 12.26 g of carbon. Find the simplest formula of this hydrocarbon.

Solution
A hydrocarbon has H and C atoms. If C = 12.26 g, the rest is hydrogen
\[
\therefore \text{mass of H} = (14.82 - 12.26) \text{ g} = 2.56 \text{ g}
\]
<table>
<thead>
<tr>
<th>Element</th>
<th>C</th>
<th>H</th>
</tr>
</thead>
<tbody>
<tr>
<td>Composition by mass in g</td>
<td>12.26</td>
<td>2.56</td>
</tr>
<tr>
<td>Get reacting moles by dividing by RAM</td>
<td>(\frac{12.26}{12} = 1.02)</td>
<td>(\frac{2.56}{1} = 2.56)</td>
</tr>
<tr>
<td>Divide by the smallest value to get whole numbers</td>
<td>(\frac{1.02}{1.02} = 1.0)</td>
<td>(\frac{2.56}{1.02} = 2.5)</td>
</tr>
<tr>
<td>Mole ratio</td>
<td>1.0</td>
<td>2.5</td>
</tr>
</tbody>
</table>

**Note:** We do not round off 2.5 to 3. To get a whole number, multiply both ratios by a certain factor to get whole numbers. In this case, we multiply both sides by 2.

\[\therefore \text{ Ratio of moles will be } 1.0 \times 2 = 2 \text{ and } 2.5 \times 2 = 5\]

\[\therefore \text{ The empirical formula is } \text{C}_2\text{H}_5\]

**The composition of a compound can also be in percentages**

**Example 5**

Calculate the empirical formula of a compound containing, 43.4% sodium, 11.3% carbon and the rest is oxygen. 
(Na = 23, C = 12, O = 16)

**Solution**

Oxygen \[= 100 - (43.4 + 11.3) = 45.3\%\]

<table>
<thead>
<tr>
<th>Element</th>
<th>Na</th>
<th>C</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Composition by mass in %</td>
<td>43.4</td>
<td>11.3</td>
<td>45.3</td>
</tr>
<tr>
<td>Get reacting moles by dividing by R.A.M</td>
<td>(\frac{43.4}{23} = 1.89)</td>
<td>(\frac{11.3}{12} = 0.94)</td>
<td>(\frac{45.3}{16} = 2.83)</td>
</tr>
<tr>
<td>Divide by the smallest to get whole numbers</td>
<td>(\frac{1.89}{0.94} = 2)</td>
<td>(\frac{0.94}{0.94} = 1)</td>
<td>(\frac{2.83}{0.94} = 3)</td>
</tr>
<tr>
<td>Mole ratio</td>
<td>2</td>
<td>1</td>
<td>3</td>
</tr>
</tbody>
</table>

Therefore, the empirical formula of this compound is \(\text{Na}_2\text{CO}_3\).
Example 6
When a 100 g of a sample of a hydrated salt was analysed, it was found to contain 25.6 g of copper, 12.8 g sulphur, 25.6 g of oxygen and the rest was water of crystallisation. Calculate the empirical formula of the salt.
(R.A.M, Cu = 64, S = 32, O = 16, H = 1)

Note: Water, (H₂O), is incorporated in the crystals as molecules. So we should calculate the moles of water in the crystals as shown below.

Mass of water = 100 – (25.6 + 12.8 + 25.6) = 36 g

<table>
<thead>
<tr>
<th>Symbol of element/compound present</th>
<th>Cu</th>
<th>S</th>
<th>O</th>
<th>H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Composition by mass in g</td>
<td>25.6</td>
<td>12.8</td>
<td>25.6</td>
<td>36.0</td>
</tr>
<tr>
<td>Moles that reacted</td>
<td>( \frac{25.6}{64} = 0.4 )</td>
<td>( \frac{12.8}{32} = 0.4 )</td>
<td>( \frac{25.6}{16} = 1.6 )</td>
<td>( \frac{36.0}{18} = 2 )</td>
</tr>
<tr>
<td>Divide by the smallest to get whole numbers</td>
<td>( \frac{0.4}{0.4} = 1 )</td>
<td>( \frac{0.4}{0.4} = 1 )</td>
<td>( \frac{1.6}{0.4} = 4 )</td>
<td>( \frac{2}{0.4} = 5 )</td>
</tr>
<tr>
<td>Mole ratio</td>
<td>1</td>
<td>1</td>
<td>4</td>
<td>5</td>
</tr>
</tbody>
</table>

Therefore, the empirical formula of the hydrated salt is \( \text{CuSO}_4 \cdot 5\text{H}_2\text{O} \)

Exercises 2.10
1. Calculate the empirical formulae of each of the compounds using the percentages of elements given.
   (a) C = 73.47%, H = 10.20%, O = 16.33%
   (b) Ca = 40%, C = 12%, O = 48%
   (c) Al = 52.9% S = 47.1%
   (d) Pb = 92.8%, O = 7.20%
   (e) C = 75%, H = 25%
   (f) Fe = 20.14%, S = 11.51%, O = 63.31% H = 5.04%

2. Glucose is composed of 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen by mass. What is its empirical formula?

3. Calculate the percentage composition of the elements in potassium hydroxide.
4. Work out the empirical formula of each of the following compounds made of:
   (a) 11.2 g of iron and 4.8 g of oxygen
   (b) 3.2 g of copper, 0.6 g of carbon and 2.4 g of oxygen
   (c) 12 g of carbon and 4 g of hydrogen.

**Molecular formula (M.F)**

To determine the molecular formula, we need to know the empirical formula and the relative molecular mass, R.M.M (or formula mass) of the compound.

The molecular formula of a compound shows the actual number of atoms of the various elements present in one molecule of that compound.

The following examples show how we find the molecular formula of compounds. **Note:** We get the molecular formula by multiplying the empirical formula by a whole number, n.

**Example 1**

Sodium metal burns in excess oxygen. It burns with a bright yellow flame and forms a yellow oxide which contains 59% sodium.

The relative molecular mass, $M_r$, of the oxide is 78. What is its molecular formula? (R.A.M Na = 23, O = 16)

The percentage of oxygen is 100 – 59 = 41%

<table>
<thead>
<tr>
<th>Element</th>
<th>Na</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Composition by mass in %</td>
<td>59</td>
<td>41</td>
</tr>
<tr>
<td>Mole reacting</td>
<td>$\frac{59}{23} = 2.57$</td>
<td>$\frac{41}{16} = 2.56$</td>
</tr>
<tr>
<td>Divide by the smaller number</td>
<td>2.57 = 1</td>
<td>2.56 = 1</td>
</tr>
<tr>
<td>Mole ratio</td>
<td>1</td>
<td>1</td>
</tr>
</tbody>
</table>

∴ The empirical formula (E.F) of the oxide is NaO.

The relative mass of empirical formula = (Na + O) = 23 + 16 = 39

The molecular mass is given by

$$(\text{NaO})_n = 78$$

This means that

$$39n = 78$$

$$n = \frac{78}{39} = 2$$
The molecular formula is \((\text{NaO})_2\) that is \(\text{Na}_2\text{O}_2\)

**Example 2**
A compound Q contains 73.5\% carbon 10.2\% hydrogen by mass and the rest is oxygen. Its relative molecular mass, \(M_r\), is 98. Calculate the molar mass of Q.

(R.A.M, \(\text{C} = 12, \text{H} = 1, \text{O} = 16\)).

**Note:** If you have understood how we proceed in these calculations, you may omit some details as shown in the following solution.

The percentage of oxygen = 100 – (73.5 + 10.2) = 16.3

<table>
<thead>
<tr>
<th>Element</th>
<th>C</th>
<th>H</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>% composition by mass</td>
<td>73.5</td>
<td>10.2</td>
<td>16.3</td>
</tr>
<tr>
<td>Moles reacting</td>
<td>(\frac{73.5}{12} = 6.1)</td>
<td>(\frac{10.2}{1} = 10.2)</td>
<td>(\frac{16.3}{16} = 1.0)</td>
</tr>
<tr>
<td>Divide by the smallest to get whole numbers</td>
<td>(\frac{6.1}{1.0} = 6.1)</td>
<td>(\frac{10.2}{1.0} = 10.2)</td>
<td>(\frac{1.0}{1.0} = 1.0)</td>
</tr>
<tr>
<td>Mole ratio</td>
<td>6</td>
<td>10</td>
<td>1</td>
</tr>
</tbody>
</table>

Empirical formula is \(\text{C}_6\text{H}_{10}\text{O}\).

The relative mass of empirical formula is:
\[6\times12 + 10\times1 + 16 = 98\]

The molecular formula \((\text{C}_6\text{H}_{10}\text{O})_n\) = 98
\[98n = 98\]
\[\therefore n = 1\]

The molecular formula is \((\text{C}_6\text{H}_{10}\text{O})_1 = \text{C}_6\text{H}_{10}\text{O}\).

\[\therefore\text{ The molar mass of Q is } = 6 \times 12 + 10 \times 1 + 16 = 98\ \text{g}.\]

**Exercise 2.11**
1. Calculate the molecular formula (M.F) of a compound which is made up of 85.7\% carbon and 14.3\% hydrogen. It’s molecular mass is 28.
2. An oxide of iron contains 70\% iron and the rest is oxygen. Calculate the empirical formula of the oxide.
3. 24.6g of a hydrogen salt of, MSO\(_4\)\(_x\) \(\times\) H\(_2\)O, on heating produced 12.0 g of anhydrous MSO\(_4\). Find the value of x. (M = 24, S = 32, O = 16)
4. 22.3 g of an oxide of lead was reduced with a gas X to produce 20.7 g of lead. Find the empirical formula of the oxide. If the relative molecular mass of the oxide is 223 g mol\(^{-1}\). What is the molar mass of the oxide?
5. An acid with a molar mass of 194 g contain 0.5 g of hydrogen, 16 g of sulphur and 32 g of oxygen. Calculate the molecular formula of the acid.

6. When a compound of nitrogen and hydrogen was analysed, it was found to contain 1.4 g of nitrogen and 0.3 g of hydrogen. The molecular mass of the compound is 17. What is the molecular formula of the compound?

2.5: Molar solution

We have at one time or another dissolved sodium chloride in water. We have also added sugar into a cup of tea and it dissolves. Study Fig. 2.11 and answer the questions that follow.

![Fig. 2.11: Adding sugar to cups of tea](image)

Assuming the amount of sugar that was added in the three cups above containing equal volume of tea is as follows; \( \frac{1}{2} \) teaspoonful, 1 teaspoonful, 2 teaspoonfuls.

- How can we tell the tea with a lot of sugar?
- How can we decrease the taste of sugar in each cup?

If we taste each sample of tea we should know the sweetest. This is the tea containing most sugar. By adding more tea that is, diluting it, we would reduce the taste of sugar. Scientists carry out experiments to find out how much of a substance is dissolved in a certain volume of solution. They do this by working out the amount dissolved and not by tasting. This way they can learn a lot more about chemical reactions. Many chemical reactions are performed with reagents in solution. It is necessary to know the amount of solute dissolved in a given amount of water. If a large volume of water is added to a small amount of solute, the solution is said to be diluted.

Concentration of solutions

The term concentration refers to the amount of solute dissolved in a specific volume of solvent. In a given amount of water, the more solute dissolved, the more concentrated the solution. If the solution contains a little solute, it is described as a dilute solution. Similarly, if a large volume of water is added to a small amount of solute, then the solution is dilute.
Molar solutions
The concentration of a solution is usually expressed in number of moles dissolved in 1000 cm$^3$, that is 1 cubic decimeter (dm$^3$) of solution.

$$1000 \text{ cm}^3 = 1 \text{ dm}^3 = 1 \text{ litre.}$$

Note
The total volume of the solute plus water must be equal to 1000 cm$^3$. This is why a 1 dm$^3$ volumetric flask is used.

When one mole of a solute is dissolved in water and the volume of the solution is made up to 1 dm$^3$, the solution is said to be a molar solution. A concentration of one mole per cubic decimeter is often written as 1 mol/dm$^3$ or 1 mol dm$^{-3}$ or 1M. Concentration expressed this way is sometimes referred to as the molarity of a solution, abbreviated as M. Therefore, a molar solution contains 1 mole of solute per dm$^3$ solution.

A solution containing 2 moles in 1000 cm$^3$ (1 dm$^3$), has a concentration of 2 mol dm$^{-3}$ or simply 2M. A solution which contains 0.5 moles of solute dissolved in 1 dm$^3$ solution would be written as 0.5 mol dm$^{-3}$ or 0.5 M. This shows the concentration is known. A solution whose concentration is known is called a standard solution.

Exercise 2.12(a)
1. What is the molarity of a solution containing:
   (a) 3 moles per dm$^3$?
   (b) 0.25 moles in 1000 cm$^3$?
   (c) 2 moles in 2 dm$^3$?

Preparation of molar solutions

Experiment 2.4
To prepare 1 mol dm$^{-3}$ sodium chloride solution (1M NaCl solution).

Apparatus and chemicals
- weighing balance
- 1 dm$^3$ volumetric flask
- stirrer/glass rod
- wash bottle
- distilled water
- beaker
- funnel
- sodium chloride
Procedure
1. Weigh 58.5 g of sodium chloride crystals.
2. Place 400 cm$^3$ of distilled water in a beaker.
3. Add a little salt into the water, stirring continuously until all the salt dissolves.
4. Using a filter funnel transfer the salt solution into a 1 dm$^3$ volumetric flask Fig. 2.12(b).
5. Rinse the beaker with distilled water carefully and add the washing into the flask.
6. Add more distilled water and shake the flask well. Add more water until the solution level is just below the calibration mark. Using a wash bottle add more distilled water drop by drop until the bottom of the meniscus is at the same level with calibration mark Fig. 2.12(c).
7. Use a stopper to cover the flask and invert it several times to make sure the solution mixes thoroughly.
8. Transfer the solution into a reagent bottle and label it 1M NaCl Fig. 2.12 (d).

Fig. 2.12: How to make 1M solution of sodium chloride

- Did all the salt dissolve in the water?
- Write the formula of sodium chloride.
• Using the relative atomic masses, find the molar mass of sodium chloride.
• What can we conclude from this experiment?

One mole of sodium chloride is 58.5 g. Therefore when we dissolve this amount of salt in 1000 cm$^3$ of distilled water, the molarity or concentration of the solution is 1 mol dm$^{-3}$ or 1M.

Suppose we did not wish to prepare so much solution i.e. 1000 cm$^3$ (1 dm$^3$), we needed only 100 cm$^3$, how much salt should we dissolve to make 1M solution? This is what we need to understand.

1M means that 1 mole of a substance, in this case salt, dissolved in 1 dm$^3$ (1000 cm$^3$).

But 1 mole of NaCl = 58.5 g
means 58.5 g were dissolved 1000 cm$^3$

∴ $x$ g are dissolved in 100 cm$^3$
Cross multiply and solve for $x$

∴ $x = 58.5 \text{ g} \times \frac{100 \text{ moles}}{1000 \text{ moles}}$

Therefore for us to make 1M NaCl; we add 5.85 g of NaCl in 100 cm$^3$ of distilled water.

Let us look at other examples.

**Example 1**
How many moles of sodium chloride are there in 100 cm$^3$ of a solution of a 2M solution?

**Solution**

**First principal method**
What we need to understand first is the meaning of 2M. 2M means 2 moles in 1000 cm$^3$.

Therefore start with a statement as follows:
2 moles are in 1000 cm$^3$ of solution
∴ $x$ moles are dissolved in 100 cm$^3$ of solution.
Then cross multiply and solve for $x$.

i.e. $x = 2 \text{ moles} \times \frac{100 \text{ cm}^3}{1000 \text{ cm}^3} = 0.2 \text{ moles}$
**Formula method**

<table>
<thead>
<tr>
<th>Concentration (Molarity)</th>
<th>[ \text{Concentration (Molarity)} = \frac{\text{no. of moles of solute}}{\text{Volume of solution (in dm}^3\text{)}} ]</th>
</tr>
</thead>
<tbody>
<tr>
<td>Substitute values and solve</td>
<td>[ 2 = \frac{\text{no. of moles of solute}}{\frac{100 \text{ cm}^3}{1000 \text{ cm}^3}} ]</td>
</tr>
<tr>
<td>no. moles</td>
<td>[ = 2 \times 0.1 ]</td>
</tr>
<tr>
<td></td>
<td>[ = 0.2 \text{ moles} ]</td>
</tr>
</tbody>
</table>

**Example 2**

How many moles are there in 24 cm\(^3\) of 2M sodium chloride solution.

**Solution**

Remember 2 M means 2 moles are dissolved in 1000 cm\(^3\) of solution. Then proceed as follows:

**First Principal method**

1000 cm\(^3\) contain 2 moles

\[ \therefore \quad 24 \text{ cm}^3 \text{ contain } x \text{ moles} \]

Cross multiply and solve for \(x\)

\[ x = \frac{24 \text{ cm}^3}{1000 \text{ cm}^3} \times 2 \text{ moles} \]

\[ = 0.048 \text{ moles} \]

Or

1000 cm\(^3\) contain \(\frac{2}{1000}\) moles

1 cm\(^3\) contain \(\frac{2}{1000}\) moles

\[ \therefore \quad 24 \text{ cm}^3 \text{ contain } \frac{2}{1000} \times 24 = 0.048 \text{ moles} \]

**Formula method**

Try example 2 using formula method.
Exercise 2.12(b)

How many moles are there in:

(a) 3000 cm$^3$ of 2 M nitric acid solution?

(b) 100 cm$^3$ of 0.25 M sodium hydroxide solution?

(c) 2 dm$^3$ of 1M sulphuric acid solution?

(d) 50 cm$^3$ of 0.5 M sodium chloride solution.

**Hint:** You do not need to work out the formula mass of substances.

Making solutions

Suppose your teacher asks you to prepare 250 cm$^3$ of 1 mol dm$^{-3}$ (1M) sodium hydroxide for a chemistry lesson. How would you prepare this solution?

**Solution**

**First Principle method**

*Step 1:* Work out the moles needed.

1 M means 1000 cm$^3$ contain 1 mole

250 cm$^3$ contain $x$ moles.

Cross multiply then solve for $x$

\[
\therefore x \text{ moles} = \frac{1 \text{ mole} \times 250 \text{ cm}^3}{1000 \text{ cm}^3} = 0.25 \text{ moles}.
\]

*Step 2:* Write the formula of sodium hydroxide and work out the R.F.M.

Sodium hydroxide (NaOH)

1 $\times$ Na = 1 $\times$ 23 = 23

1 $\times$ O = 1 $\times$ 16 = 16

1 $\times$ H = 1 $\times$ 1 = 1

R.F.M = 23 + 16 + 1 = 40

This means that

1 moles of NaOH = 40 g

0.25 moles of NaOH = $y$ g

Cross multiply and solve for $y$

1 moles $\times$ $y$ g = 40 g $\times$ 0.25 moles

\[
\therefore y = \frac{40g \times 0.25 \text{ moles}}{1 \text{ mole}} = 10 g
\]

This means you weigh 10 g of sodium hydroxide pellets and put in a 250 cm$^3$ volumetric flask, dissolve in distilled water. Make it up to 250 cm$^3$ mark with distilled water.
### Formula method

**Step 1**

Molarity \[ \text{Molarity} = \frac{\text{number of moles of solute}}{\text{volume of solution (in dm}^3\text{)}} \]

Substitute the values

\[
1 = \frac{\text{number of moles of solute}}{0.25}
\]

Cross multiply

Number of moles \[ = 1 \times 0.25 \]
\[ = 0.25 \text{ moles.} \]

**Step 2**

Number of moles \[ = \frac{\text{number of grams}}{\text{mass of 1 mole}} \]

\[
0.25 = \frac{\text{number of grams}}{40}
\]

Number of grams \[ = 0.25 \times 40 \]
\[ = 10 \text{ g} \]

Note that masses from the two methods are the same.

### Exercise 2.13

1. Calculate the mass in grams of each compound in each of the following solutions.
   - (a) 80 cm$^3$ of 0.5 M sodium hydroxide
   - (b) 100 cm$^3$ of 1 M potassium hydroxide
   - (c) 25 cm$^3$ of 1M sulphuric acid
   - (d) 13 cm$^3$ of 0.25 M ammonium phosphate
   - (e) 250 cm$^3$ of 0.5 M sodium chloride solution.

   **Hint:** You should check the R.A.M. of the elements at the back of the book.

2. What is the concentration, in mole dm$^3$ of a solution containing 9.8 g of sulphuric acid in 100 cm$^3$.

### Concentration in grams per dm$^3$

The concentration (or molarity) can also be expressed in grams per cubic decimeter (g dm$^{-3}$). We may first calculate the concentration in mol dm$^{-3}$ and then convert to g dm$^{-3}$. 
**Example 1**

53 g of anhydrous sodium carbonate were dissolved in 2 dm³. Calculate the concentration in g dm⁻³ (Na = 23, C = 12, O = 16).

**Solution**

**First principle method**

Remember the standard volume is 1 dm³.

The formula of sodium carbonate is, \( \text{Na}_2\text{CO}_3 \).

First work out the relative formula mass of \( \text{Na}_2\text{CO}_3 \).

\[
\begin{align*}
2 \times \text{Na} &= 2 \times 23 = 46 \\
1 \times \text{C} &= 1 \times 12 = 12 \\
3 \times \text{O} &= 3 \times 16 = 48 \\
\therefore \ R.F.M &= 106
\end{align*}
\]

Using the R.F.M obtained, calculate the number of moles as follows:

\[
1 \text{ mole of } \text{Na}_2\text{CO}_3 = 106 \text{ g}
\]

\[
\therefore x \text{ moles of } \text{Na}_2\text{CO}_3 = 53 \text{ g}
\]

Cross multiply and solve for \( x \)

\[
x \text{ moles} = \frac{53 \text{ g} \times 1 \text{ mole}}{106 \text{ g}} = 0.5 \text{ moles}.
\]

Using the number of moles let us calculate the concentration in mol dm⁻³.

0.5 mole was dissolved in 2 dm³.

\( y \) moles were dissolved in 1 dm³

Cross multiply and solve for \( y \)

\[
\therefore y = \frac{0.5 \text{ moles} \times 1 \text{ dm}^3}{2 \text{ dm}^3} = 0.25 \text{ moles dm}^{-3}
\]

Finally, we convert mol dm⁻³ to g dm⁻³ as follows:

1, mole of \( \text{Na}_2\text{CO}_3 \) = 106 g

\[
\therefore 0.25 \text{ moles of } \text{Na}_2\text{CO}_3 = x \text{ g}
\]

Cross multiply and solve for \( x \)

\[
x = 1 \text{ mole} \times xg = 0.25 \text{ moles} \times 106 \text{ g}
\]

\[
x = \frac{0.25 \text{ moles} \times 106 \text{ g}}{1 \text{ mole}} = 26.5 \text{ g}
\]

Concentration expressed in g per dm³ = 26.5 g dm⁻³
### Formula Method

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Number of moles</strong></td>
<td>( \frac{\text{Number of grams}}{\text{mass of 1 mole}} )</td>
</tr>
<tr>
<td><strong>Number of moles</strong></td>
<td>( \frac{53}{106} )</td>
</tr>
<tr>
<td></td>
<td>( = 0.5 ) moles</td>
</tr>
<tr>
<td><strong>Molarity</strong></td>
<td>( \frac{\text{Number of moles of solute}}{\text{Volume of solution (in dm}^3\text{)}} )</td>
</tr>
<tr>
<td></td>
<td>( \frac{0.5}{2} )</td>
</tr>
<tr>
<td></td>
<td>( = 0.25 \text{ M} )</td>
</tr>
<tr>
<td>**Mass dm}^3\text{ })</td>
<td>( = \text{molarity} \times \text{mass of 1 mole} )</td>
</tr>
<tr>
<td></td>
<td>( = 0.25 \times 106 )</td>
</tr>
<tr>
<td></td>
<td>( = 26.5 \text{ g dm}^3)</td>
</tr>
</tbody>
</table>

### Exercises 2.14
1. Express the following molarities as concentration in g dm\(^{-3}\).
   (a) 0.2 M potassium hydroxide
   (b) 0.1M nitric acid solution.
   (c) 0.1 M sodium carbonate solution.

2. An aqueous solution of barium hydroxide has a concentration of mol dm\(^{-3}\). What is its concentration in g dm\(^{-3}\)?
   \( \text{(Ba} = 137, \text{ O} = 16, \text{ H} = 1) \)

3. 0.1 mole of sodium chloride was dissolved in 100 cm\(^3\) of water. Calculate the concentration of this aqueous solution in g dm\(^{-3}\). \( \text{(Na} = 23, \text{ Cl} = 35.5) \)

### Calculation of molarity of ions in molar solutions

We should write an equation showing ions present first.

### Example 1

Calculate the molarity of sodium (\( \text{Na}^+ \)) and chloride, (\( \text{Cl}^- \)) ions in 0.5 M sodium chloride solution.

#### Solution

\[
\begin{align*}
\text{NaCl(aq)} & \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq). \\
1 \text{ mole} & \rightarrow 1 \text{ mole} \quad 1 \text{ mole}. \\
\therefore \quad 0.5 \text{ moles} & \rightarrow 0.5 \text{ moles} \quad 0.5 \text{ moles}
\end{align*}
\]

The molarity is 0.5 M Na\(^+\) ions and 0.5 M Cl\(^-\) ions.
Examples 2
Calculate the molarity of calcium, Ca\(^{2+}\) ions and chloride, (Cl\(^{-}\)) ions in 2M CaCl\(_2\) solution.

\[
\text{CaCl}_2(aq) \rightarrow \text{Ca}^{2+}(aq) + 2 \text{Cl}^{-}(aq)
\]

1 mole 1 mole 2 moles
2 moles 2 moles 4 moles

Since the mole ratio of Ca\(^{2+}\) : Cl\(^{-}\) is 1 : 2

Molarity of Ca\(^{2+}\) ions = 2M
∴ Molarity Cl\(^{-}\) ions = 4M

Examples 3
In example 2, what is the concentration of calcium ions and chloride ions in g dm\(^{-3}\)? (Ca = 40, Cl = 35.5).

2M means 2 mol in 1000 cm\(^3\) (1dm\(^3\))
But 1 mol of Ca\(^{2+}\) = 40 g
Molarity of Ca\(^{2+}\) ions in g per dm\(^3\) = 2 \times 40 = 80 g dm\(^{-3}\)

4 M means 4 moles in 1000 cm\(^3\) (1dm\(^3\))
1 mol of Cl\(^{-}\) = 35.5 g
4 moles of Cl\(^{-}\) = 35.5 g \times 4
= 142 g
= 142 g dm\(^{-3}\)

Note: Loss or gain of electrons has negligible change in mass of an atom. This is why for example irrespective of Ca\(^{2+}\) having lost 2 electrons, the mass still remains the same. The same applies to Cl and Cl\(^{-}\).

Exercise 2.15
1. Calculate the molarity of magnesium, Mg\(^{2+}\) ions and chloride, Cl\(^{-}\) ions in 0.5 moles dm\(^{-3}\) solution of magnesium chloride (Mg = 24, Cl = 35.3).
2. Calculate the molarity of Na\(^{+}\) ions and SO\(_{3}^{2-}\) ions in 0.2 M solution of sodium sulphite. (Na = 23, S = 32, O = 16).
3. What is the concentration in g dm\(^{-3}\) of ammonium and sulphate ions in 0.4 M solution of ammonium sulphate? (Na = 23, H = 1, S = 32, O = 16).
Concentration and dilution
We have learnt that concentration refers to the amount of solute in a specific volume of solvent. When more solvent is added to a concentrated solution it is said to be diluted. Let us now investigate what happens to a solution when different amounts of a solvent are added to it.

Experiment 2.4
Investigation of concentration and dilution

Apparatus and chemicals
• copper(II) sulphate crystals or potassium manganate(VII)
• beaker
• stirring rod

Procedure
1. Label four 250 cm$^3$ beakers A, B, C and D respectively.
2. Place equal amounts of crystals (either copper(II) sulphate or potassium manganate(VII)) into each beaker.
3. Place 50 cm$^3$ of water into beaker A, 100 cm$^3$ into beaker B, 150 cm$^3$ into beaker C and 200 cm$^3$ into beaker D as illustrated in Fig. 2.13.

4. Stir the mixture thoroughly.
   • What do you observe? What conclusion can you make from your observations?
   • Record your observation and conclusion in your notebook.

A student has a birthday party and has some juice. How would the student make sure the juice is enough for the visitors to enjoy?

We have seen that the concentration of a solution is expressed in moles per cubic decimeter (mol dm$^{-3}$) or g dm$^{-3}$. If we dilute a solution (add more solvent), the number of moles of solute in the solution does not change.

If we represent initial concentration as $C_1$ and the initial volume by $V_1$, we can get the number of moles of solute by multiplying the concentration by volume.
If we represent the concentration of the dilute solution by \( C_2 \) and the new volume by \( V_2 \), the number of moles can be obtained by again multiplying concentration by volume i.e.

\[ C_2 \times V_2 = \text{moles}. \]

Since moles remain the same after dilution then

\[ C_1 \times V_1 = C_2 \times V_2 \]

\[ \text{OR simply} \]

\[ C_1 \ V_1 = C_2 \ V_2 \]

We can use this relationship to calculate the concentration after dilution, \( C_2 \) or new volume, \( V_2 \).

**Example 1**

How much water must be added to 5 cm\(^3\) of 6.0 M HCl to dilute the acid to 0.1M?

**Solution**

\[ C_1 = 6.0 \text{M} \]
\[ V_1 = 5.0 \text{ cm}^3 \]
\[ C_2 = 0.1 \text{M} \]
\[ V_2 = \text{?} \]
\[ C_1 \times V_1 = C_2 \times V_2 \]

Make \( V_2 \) subject of the formula

\[ V_2 = \frac{C_1 \times V_1}{C_2} \]

Substitute the values

\[ V_2 = \frac{6.0 \times 5.0}{0.1} = 300 \text{ cm}^3 \]

**Note:** This does not mean that 300 cm\(^3\) of water must be added, but that the total volume of solution must be made up to 300 cm\(^3\). The volume of water to be added is 295 cm\(^3\) i.e. \((300 - 5)\text{cm}^3 = 295 \text{ cm}^3\)
Example 2
Suppose 180 cm$^3$ of a 2.0 M solution is diluted to 1.0 dm$^3$, what will the concentration (molarity) of the resulting solution?

Solution

\[ C_1V_1 = C_2V_2 \]

\[ C_2 = \frac{C_1V_1}{V_2} \]

Substituting values

\[ C_2 = \frac{180 \times 2.0}{1000} = 0.36 \text{ M or 0.36 mol dm}^{-3} \]

Exercise 2.16
1. If 2 dm$^3$ of 1.0 mol dm$^{-3}$ solution of a substance were diluted to 6 dm$^3$, find the concentration (molarity) of the diluted solution.
2. The concentration of sulphuric acid in the lab is 18.4 mol dm$^{-3}$. What volume of the acid should be used to prepare a 500 cm$^3$ solution of 2 mol dm$^{-3}$?
3. If 175 cm$^3$ of 2.0 M solution are diluted to 1.0 dm$^3$, what will be the molarity of the solution?

Stoichiometry of chemical reactions
Stoichiometry of chemical reactions means balancing chemical equations. A stoichiometric equation in chemistry is just a normal balanced chemical equation. In order to solve problems on stoichiometry of chemical reactions, the following should be done:

- Writing of correct chemical equations.
- Writing correct state symbols of reactants and products.
- Writing correct ionic equation.

Unbalanced equations are unacceptable in chemistry. Remember balanced equations help us to predict the mass of reactants and products. For example in a chemical factory, if a customer orders a product we can calculate how much raw materials are required to satisfy the customer’s needs. The numbers we use to balance chemical equations are actually the number of moles required for reactants and products. This is illustrated by Fig. 2.14 below.
Example 1
When magnesium ribbon is burned in oxygen, magnesium oxide is formed.

\[ \text{Mg}(s) + \text{O}_2(g) \rightarrow \text{MgO}(s) \]

We can balance the equation by writing numbers in front of some formula to make atoms of each element equal on both sides of the equation as follows:

\[ 2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s) \]

Example 2
When we heat an equal mixture of iron filings and sulphur, the two elements combine forming a new substance called iron(II) sulphide.

\[ \text{Fe}(s) + \text{S}(s) \rightarrow \text{FeS}(s) \]

From the equation, we can see that one mole of iron atoms in the solid state reacts with one mole of sulphur atoms in the solid state to produce one mole of iron(II) sulphide also in the solid state.

Remember, we must include the state symbols in all equations.
- \((s)\) means solid
- \((l)\) means liquid
- \((g)\) means gas
- \((aq)\) (aqueous) means the substance is dissolved in water.

A fully balanced chemical equation shows the
- reactants and the products
- symbols or formulae of all substances in the reaction
• number of moles of the reactants and products formed
• physical states of all reactants and products.

Chemical equations should first be remembered in words. Then, we should write equations in formulae form and finally balance the equations. We always write a number in front of a symbol or formula. Never balance an equation by writing a number as a subscript as follows:

Cu(s) + O_2(g) \longrightarrow CuO(s). This equation is unbalanced because we have one mole of oxygen atoms on the products side. The following is a wrongly balanced equation.

Cu(s) + O_2(g) \longrightarrow CuO_2(s) \quad (This \ is \ incorrect)

This is the correct way of balancing the equation:

2Cu(s) + O_2(g) \longrightarrow 2CuO(s) \quad (This \ is \ correct)

Exercise 2.17
1. Change the following word equations into symbols/formulae and balance the equations.
   (a) Zinc + hydrochloric acid \longrightarrow zinc chloride + hydrogen
   (b) Copper(II) oxide + hydrogen \longrightarrow copper + water
   (c) Sodium + sulphuric \longrightarrow sodium + water + carbon(IV)
      hydrogen- acid sulphate carbonate
   (d) Burning charcoal in a limited supply of air to form carbon(II)oxide.

2. Write balanced equations for the following reactions.
   (a) Neutralisation of sodium hydroxide solution with dilute sulphuric acid.
   (b) Burning methane (CH_4) in plentiful supply of air producing carbon(IV)
       oxide and water vapour.
   (c) Reducing iron(III) oxide with carbon(II) oxide to produce iron metal and
       carbon(IV) oxide.
   (d) Burning hydrogen in oxygen to form water.

3. Balance the following chemical equations:
   (a) BaCO_3(s) + HCl (aq) \longrightarrow BaCl_2(aq) + H_2O(l) + CO_2(g)
   (b) Pb(NO_3)_2 (aq) + NaI(aq) \longrightarrow PbI_2(s) + NaNO_3(aq)
(c) $\text{CuSO}_4(\text{aq}) + \text{Fe(s)} \rightarrow \text{Cu(s)} + \text{FeSO}_4(\text{aq})$

(d) $\text{Na}_2\text{CO}_3(\text{aq}) + \text{HCl(\text{aq})} \rightarrow \text{NaCl(\text{aq})} + \text{H}_2\text{O(1)} + \text{CO}_2(\text{g})$

(e) $\text{Zn(NO}_3\text{)_2(s)} \rightarrow \text{heat} \rightarrow \text{ZnO(s)} + \text{NO}_2(\text{g}) + \text{O}_2(\text{g})$

(f) $\text{C}_4\text{H}_{10(\text{g})} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O(1)}$

(g) $\text{Al}_2(\text{SO}_4\text{)_3(\text{aq})} + \text{NaOH(\text{aq})} \rightarrow \text{Al(OH)}_3(\text{s}) + \text{Na}_2\text{SO}_4(\text{aq})$

**Ionic equations**

An ionic equation represents a chemical change by means of ions. However, symbols and formulae of insoluble elements and compound sometimes appear in ionic equations. These type of equations usually can confuse us if we do not understand what we need to do. We have previously learnt about ionic (electrovalent) compounds. The particles that make ionic compounds are ions. They are arranged in a regular pattern in fixed positions in a solid.

When an ionic salt is dissolved in water, the ions are set free to move throughout the solution. In some reactions, like *precipitation*, ions are swapped as shown in Fig. 2.15.

![Fig. 2.15: How precipitation reactions take place](image)

**Experiment 2.5**

Making an insoluble salt by precipitation.

**Apparatus and chemicals**

- test tubes
- test tube rack
- filter paper
- funnel
- lead(II) nitrate solution.
- potassium iodide solution.

**Procedure**

1. Place 5 cm$^3$ of lead(II) nitrate solution in a test tube.
2. Add 10 cm$^3$ of potassium iodide solution to the lead(II) nitrate solution
   Fig. 2.16. What do you observe?
3. Filter your mixture. What is the colour of the solid left on the filter paper?
4. What is the name of this solid?
5. Rinse the solid with distilled water. Why should the solid be rinsed?
6. Repeat steps 1 – 4 using barium nitrate instead of lead(II) nitrate and sodium carbonate instead of potassium iodide.
   (i) What is the colour of the residue after filtration?
   (ii) Which ions formed the solids?
   (iii) Name the ions in the filtrate.
7. Record your observations in your notebook.
8. What can you conclude from your observations?

Fig. 2.16: Precipitation of lead(II) iodide

When the potassium iodide solution was added into the lead(II) nitrate solution a yellow precipitate of lead(II) iodide was formed. The precipitate was rinsed to wash away any soluble salt on it.

Note: How were the ions swapped? Lead(II) ions, (Pb$^{2+}$) exchanged nitrate (NO$_3^-$) ions for iodide ions, (I$^-$) and the two formed a yellow precipitate (ppt) PbI$_2$. Potassium, (K$^+$) and nitrate ions, (NO$_3^-$) remained free in the solution just as they were before mixing. They have not changed. Such ions that do not take part in the reaction are called spectator ions. When we write an ionic equation, we omit the spectator ions. So the ionic equation in the above reaction between Pb(NO$_3$)$_2$ (aq) and KI(aq) is as follows:

\[
Pb^{2+} (aq) + 2I^- (aq) \rightarrow \text{PbI}_2(s)
\]

Compare the reaction between barium nitrate solution and sodium carbonate solution with the following sport’s Fig. 2.17 then write the ionic equation.
In a reaction between sodium carbonate solution and barium nitrate solution the ionic equation is as follows:

\[ \text{Ba}^{2+}(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{BaCO}_3(s) \]

An ionic equation shows only those ions that change in some way during a reaction. From the above example, the ions that do not undergo any change are omitted from the ionic equation.

Some possibilities during ionic changes include:
1. Formation of a **precipitate**.
2. Evolution of a **gas**.
3. Formation of **water**.

Knowledge of solubility rules and the state of the reactants and the products is useful. Solids, liquids and gases do not have free ions. But when we dissolve ionic substances the ions are set free, for example when we dissolve sodium chloride in water, it disappears. This is because it breaks down into Na\(^+\) and Cl\(^-\) ions as follows:

\[ \text{NaCl}(s) + \text{water} \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq) \]

If you have difficulties in choosing the predominant ions to be written down, follow the steps outlined below.
**Steps to follow in writing ionic equations**

1. Write a balanced chemical equation of the reacting substances.
2. From the equation check for the substances which dissolve thereby setting ions free.
3. Write down the ions in the equation separately.
4. Cross out the common ions that appear on both sides of the equation.

**Note:** When in aqueous solution the following compounds are generally likely to be in form of free ions

- soluble metal salts and hydroxides
- ammonium compounds
- dilute acids

Let us illustrate the above steps by considering precipitation reactions.

**Double decomposition (precipitation)**

We learnt that a precipitate is an insoluble salt or hydroxide. Precipitation refers to the process of forming one soluble salt and one insoluble salt when two solutions of soluble salts are mixed.

(a) Write a balanced ionic equation for the reaction between aqueous sodium chloride solution and aqueous silver nitrate solution.

\[
\text{From } \text{NaCl}_{(aq)} \quad \rightarrow \quad \text{Na}^+(aq) + \text{Cl}^-(aq)
\]

\[
\text{From } \text{AgNO}_3_{(aq)} \quad \rightarrow \quad \text{Ag}^+(aq) + \text{NO}_3^-(aq)
\]

\[
\text{Na}^+(aq) + \text{Cl}^-(aq) + \text{Ag}^+(aq) + \text{NO}_3^-(aq) \rightarrow \text{Na}^+(aq) + \text{NO}_3^-(aq) + \text{AgCl(s)}
\]

Ionic equation: \( \text{Cl}^-(aq) + \text{Ag}^+(aq) \rightarrow \text{AgCl(s)} \)

(b) Write an ionic equation for the reaction between aqueous solution of barium chloride and sodium sulphate.

\[
\text{From } \text{BaCl}_2_{(aq)} \quad \rightarrow \quad \text{Ba}^{2+}(aq) + 2\text{Cl}^-(aq)
\]

\[
\text{From } \text{Na}_2\text{SO}_4_{(aq)} \quad \rightarrow \quad 2\text{Na}^+(aq) + \text{SO}_4^{2-}(aq)
\]

The ions that we are mixing are 4 types, which then swap partners.

\[
\text{Ba}^{2+}(aq) + 2\text{Cl}^-(aq) + 2\text{Na}^+(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s) + 2\text{Na}^+(aq) + 2\text{Cl}^-(aq)
\]
Cross the spectator ions

Therefore, the ionic equation is:-

$$\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$$

In these two examples we see that precipitates are formed by the predominant ions. Spectator ions remain in solution while the precipitate settles at the bottom as shown in Fig. 2.18.

Note: We usually omit charges on the ions when we write the formula of the compound eg. BaSO₄.

Study the information in the following table on solubility rules to remind you the ions that form precipitates.

**Table 2.10: Solubility of salts**

<table>
<thead>
<tr>
<th>Soluble</th>
<th>Insoluble</th>
</tr>
</thead>
<tbody>
<tr>
<td>• All sodium, potassium and ammonium salts</td>
<td>none</td>
</tr>
<tr>
<td>• All nitrates</td>
<td>none</td>
</tr>
<tr>
<td>• All chlorides</td>
<td>except Lead(II) chloride is soluble in hot water, Silver chloride</td>
</tr>
<tr>
<td>• All sulphates</td>
<td>except Lead(II) sulphate, Barium sulphate, Calcium sulphate is slightly soluble.</td>
</tr>
<tr>
<td>• Na, K, NH₄ carbonates</td>
<td>All other carbonates</td>
</tr>
<tr>
<td>• Na, K, NH₄ hydroxides</td>
<td>All other hydroxides, (calcium and magnesium hydroxides are slightly soluble)</td>
</tr>
</tbody>
</table>
In example (b) on page 93 we would not have chosen Na\(^+\)(aq) and \(SO_4^{2-}\)(aq) because Na\(_2\)SO\(_4\) is not among the insoluble sulphates in the table. But note, BaSO\(_4\) is among the insoluble sulphates.

**Application of double decomposition in calculations**

**Example**
20 g of zinc sulphate salt were dissolved in distilled water. Excess sodium hydroxide solution was added to the metal salt solution.

(i) Write an equation for the reaction.

(ii) Calculate the mass of the precipitate formed.

**Solution**

(i) \(ZnSO_4(aq) + 2NaOH(aq) \rightarrow Zn(OH)_2(s) + Na_2SO_4(aq)\)

(ii) From equation (i)

1 mole of \(ZnSO_4(aq)\) produces 1 mole of \(Zn(OH)_2(s)\)

1 mole of \(ZnSO_4\) expressed in grams = 65 + 32 + 4(16) = 161 g

1 mole of \(Zn(OH)_2\) = 65 + 2(16) + 2(1) = 99 g

This means 161g of \(ZnSO_4\) produce 99 g of \(Zn(OH)_2\)

∴ 20 g of \(ZnSO_4\) produce \(x\) g of \(Zn(OH)_2\)

Cross multiply and solve for \(x\)

\[
161 \times x = 99 \times 20 \ g
\]

\[
x = \frac{99 \times 20 \ g}{161 \ g} = 12.3 \ g
\]

∴ 20 g of \(ZnSO_4(aq)\) produce 12.3 g of \(Zn(OH)_2(s)\)

**Exercise 2.18**

1. 10 g of aluminum sulphate salt were dissolved in distilled water and the solution was reacted with excess sodium hydroxide solution to form a precipitate. After filtration the precipitate was washed with distilled water and dried.

   (a) Write an ionic equation for the reaction

   (b) Calculate the mass of precipitate formed.

2. A student reacted 10.8 g of aqueous silver nitrate and aqueous sodium chloride.

   (i) Write chemical and ionic equations for the reaction.

   (ii) Write the formulae of the spectator ions.

   (iii) Calculate the mass of silver chloride precipitated.
3. Calculate the mass of precipitate formed when a solution of zinc sulphate was reacted with sodium hydroxide solution. Write an ionic equation for the reaction.

4. Write ionic equations for the reaction between aqueous solutions of the following
   (i) Copper(II) sulphate and sodium hydroxide.
   (ii) Barium nitrate and sodium carbonate
   (iii) Lead(II) nitrate and potassium hydroxide
   (iv) Magnesium chloride and silver nitrate.

5. A solution of copper(II) sulphate reacts with sodium hydroxide to produce a precipitate of copper(II) hydroxide. What mass of sodium hydroxide would be needed to change 15.6 g copper(II) sulphate to copper(II) hydroxide and what mass of copper(II) hydroxide would be produced.

**Ionic equation for displacement reactions**
A metal which is higher in the reactivity series is more reactive than that lower than it. A more reactive metal displaces a less reactive one from its salt solution. In this process the more reactive metal atoms lose electrons when they dissolve. These electrons are then gained by the metal ions in the solution. This is a redox reaction as shown in Experiment 2.6.

**Experiment 2.6**
To displace copper from a solution of copper(II) sulphate with zinc.

**Apparatus and chemicals**
- beaker
- stirrer
- spatula
- zinc granules
- hydrated copper(II) sulphate

**Procedure**
1. Place about 50 cm$^3$ of distilled water in a beaker.
2. Add a little hydrated copper(II) sulphate into the water and stir.
3. Keep on dissolving copper(II) sulphate until you get a deep blue solution.
4. Now add pieces of zinc granules and leave the solution for some time.
5. What do you observe?
6. What do you conclude from the above observations?
   - Record your observations and conclusion in your notebook,
   - Write a chemical equation for the reaction.
   - From the equation, deduce the ionic equation.
   - Suggest another metal that you would use in place of zinc and write an ionic reaction for the displacement.
The solution turns from blue to colourless. Zinc dissolves and a brown solid which is copper is deposited at the bottom of the beaker as shown in Fig. 2.19.

![Fig. 2.19: Displacement of copper from copper(II) sulphate solution](image)

Zn(s) + CuSO$_4$(aq) $\rightarrow$ ZnSO$_4$(aq) + Cu(s)

Ionic equation

Zn(s) + Cu$^{2+}$(aq) $\rightarrow$ Zn$^{2+}$(aq) + Cu(s)

Note: Zinc goes into solution by losing 2 electrons (Oxidation). Copper(II), Cu$^{2+}$, ions gain the electrons (Reduction) to form copper (brown solid).

Zn(s) $\rightarrow$ Zn$^{2+}$(aq) + 2e$^-$ (loss of electron – oxidation)

Cu$^{2+}$(aq) + 2e$^-$ $\rightarrow$ Cu(s) (gain of electrons – reduction)

Overall reaction Zn(s) + Cu$^{2+}$(aq) $\rightarrow$ Zn$^{2+}$(aq) + Cu(s) (redox reaction)

Example 1

A student warmed copper(II) sulphate solution in a beaker. The student added 2.52 g of iron filings into the solution, stirring well until the reaction was complete.

(a) What observation did the student make?
(b) Write an ionic equation for the reaction.
(c) The student filtered the solution, rinsed the residue with distilled water and dried it. The mass of the residue was 2.86 g. Calculate the amount of residue that could be formed if the student used 1 mole of iron. (Fe = 56).

Solution

(a) The solution turned from blue to pale green because iron displaced copper from the solution. A brown solid was deposited at the bottom of the beaker.

(b) Cu$^{2+}$(aq) + Fe(s) $\rightarrow$ Fe$^{2+}$(aq) + Cu(s)

(Blue) (pale-green) (red-brown)
(c) 1 mole of iron = 56g.

If 2.52 of iron formed 2.86 g of copper then
56 g of iron would form x g of copper
Arms multiply and solve for x

\[
x \times 2.52 \, \text{g} = 2.86 \, \text{g} \times 56 \, \text{g}
\]

\[
x = \frac{2.86 \, \text{g} \times 56 \, \text{g}}{2.52 \, \text{g}} = 63.55 \, \text{g}
\]

\[
= 63.6 \, \text{g}
\]

1 mole of iron displaces 63.6 g of copper
\[\therefore\] The amount residue formed = 63.6 g

1. Ionic equations involving liberation of gases

Examples

(i) Reaction between a metal and a dilute acid

\[
\text{Zn(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g})
\]

\[
\text{Mg(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{H}_2(\text{g})
\]

(ii) Reaction between a soluble carbonate or hydrogen carbonate and an acid.

When a dilute acid reacts with a carbonate or hydrogen carbonate, a salt, water and carbon(IV) oxide are formed.

\[
\text{Na}_2\text{CO}_3(\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{H}_2\text{O}(1) + \text{CO}_2(\text{g})
\]

\[
\text{NaHCO}_3(\text{aq}) + \text{HCl}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(1) + \text{CO}_2(\text{g})
\]

Ionic equations

\[
\text{CO}_3^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) \rightarrow \text{H}_2\text{O}(1) + \text{CO}_2(\text{g})
\]

\[
\text{HCO}_3^-(\text{aq}) + \text{H}^+(\text{aq}) \rightarrow \text{H}_2\text{O}(1) + \text{CO}_2(\text{g})
\]

(iii) Reaction between an insoluble carbonate and dilute acid. For example

\[
\text{ZnCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2\text{O}(1) + \text{CO}_2(\text{g})
\]

Ionic equation

\[
\text{ZnCO}_3(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2\text{O}(1) + \text{CO}_2(\text{g})
\]
Note: Zinc carbonate solid contains ions but they are not free. So we would not write the ions separately.

2. Ionic equations for neutralisation reactions involving: A soluble base (alkali) and an insoluble base
   (i) Reaction with a soluble base (alkali)
       When an acid and an alkali are mixed the resulting mixture is a salt solution which appears just like water. Therefore we choose the ions that form water, when writing an ionic equation.
       \[ \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(L) \]
       Ionic equation:
       \[ \text{H}^+ (aq) + \text{OH}^- (aq) \rightarrow \text{H}_2\text{O}(l) \]
   (ii) Reaction with an insoluble base
       Copper(II) oxide is an insoluble base. When it is reacted with warm dilute sulphuric acid, a blue solution of copper(II) sulphate is formed.
       \[ \text{H}_2\text{SO}_4(aq) + \text{CuO}(s) \rightarrow \text{CuSO}_4(aq) + \text{H}_2\text{O}(l) \]
       Ionic equation:
       \[ 2\text{H}^+ (aq) + \text{CuO}(s) \rightarrow \text{Cu}^{2+} (aq) + \text{H}_2\text{O}(l) \]

Exercise 2.19
Write ionic equations for the following reactions:
   (i) Calcium carbonate and dilute nitric acid.
   (ii) Copper(II) oxide and dilute hydrochloric acid.
   (iii) Aqueous potassium hydroxide and dilute sulphuric acid.

Calculations of reacting quantities
In these calculations, we must always write balanced equations. These will help us see the mole ratios of the reactants and products hence calculate what is required.

Reacting masses

Example 1
What mass of magnesium oxide is produced when 48 g of magnesium is burned in air?
   (Mg = 24; O = 16)
Solution
Note: We do not need to work the mass of oxygen. Only magnesium and magnesium oxide are mentioned in the problem.

\[2\text{Mg(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO(s)}\]

2 moles of Mg react with one mole of oxygen to form 2 moles of MgO

\[
\begin{align*}
1 \text{ mole of Mg} & = 24 \text{ g} \\
2 \text{ moles of Mg} & = x \text{ g}
\end{align*}
\]
\[
\begin{align*}
1 \text{ mole of MgO} & = 24 + 16 = 40 \text{ g} \\
2 \text{ moles of MgO} & = y
\end{align*}
\]
\[
\therefore x = 2 \times 24 = 48 \text{ g} \\
\therefore y = 2 \times 40 = 80 \text{ g}
\]

This means 48 g of Mg when burned produce 80 g of MgO.

Example 2
Copper(II) oxide is reduced by heating with excess hydrogen to form copper metal. What mass of copper(II) oxide will react with hydrogen to produce 8 g of copper? \((\text{Cu} = 63.5; \text{O} = 16)\).

Solution
Use a pencil to underline what we are given and what we want to calculate from a balanced equation

Then write mole ratios as shown below.

Method (a): Using grams

\[\text{CuO(s)} + \text{H}_2\text{(g)} \rightarrow \text{Cu(s)} + \text{H}_2\text{O(g)}\]

1 mole of CuO produces 1 mole of Cu
But 1 mole of CuO = 63.5 + 16 = 79.5 g and 1 mole of Cu = 63.5 g

Let us re-write the equation and use grams to calculate.

\[79.5 \text{ g of CuO produce } 63.5 \text{ g of Cu}\]
\[x \text{ g CuO will produce } 8 \text{ g of Cu}\]

Cross multiply and solve for \(x\)

\[
\therefore x = \frac{79.5 \times 8}{63.5} = 10.02 \text{ g}
\]

Therefore 10.02 g of copper(II) oxide are needed to produce 8 g of copper metal.
**Method (b) Using moles**

\[ \text{CuO(s)} + \text{H}_2(\text{g}) \rightarrow \text{Cu(s)} + \text{H}_2\text{O(g)} \]

1 mole of CuO produces 1 mole of Cu

But 1 mole of Cu = 63.5 g

\[ x \text{ moles of Cu} = 8 \text{ g} \]

Cross multiply and solve for \( x \)

\[ x = \frac{8}{63.5} = 0.126 \text{ moles of copper} \]

Let us re-write the equation and use mole ratios

\[ \text{CuO(s)} + \text{H}_2(\text{g}) \rightarrow \text{Cu(s)} + \text{H}_2\text{O(g)} \]

1 mole of CuO produces 1 mole of Cu

\[ y \text{ moles CuO will produce 0.126 moles of Cu} \]

Cross multiply and solve for \( y \)

\[ y = 1 \times 0.126 = 0.126 \text{ moles of CuO} \]

But 1 mole of CuO = 79.5 g

0.126 moles of CuO = \( x \) g

\[ x = 79.5 \times 0.126 \]

\[ x = 10.02 \text{ g} \]

Sometimes we may want to find large quantities of substances, for example, a product from a factory.

**Example 3**

![Image of two mice discussing zinc](image)

*Fig. 2.20: Using moles to calculate large quantities*

A customer requires 115 tonnes of zinc. How many tonnes of zinc oxide do we require to be reduced by carbon? (Zn = 65; O = 16; 1 tonne = 1000 kg).

**Solution**

\[ \text{ZnO(s)} + \text{C(g)} \rightarrow \text{Zn(s)} + \text{CO(g)} \]

1 mole of ZnO produces 1 mole of Zn
1 mole of ZnO = 65 + 16 = 81 g  
1 mole of Zn = 65 g

It implies that 81 g of ZnO gives 65 g of Zn on reduction.
It also means 81 tonnes of ZnO would give 65 tonnes of zinc.
Therefore x tonnes of ZnO would give 115 tonnes of zinc.
Cross multiply and solve for x
\[
\frac{x}{115} = \frac{81}{65} \implies x = \frac{81 \times 115}{65} = 143.3 \text{ tonnes}
\]

Therefore 143.3 tonnes of zinc oxide are required.

You can also use kg and then finally convert them to tonnes as follows:

\[
\text{ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)}
\]

Mole ratio of ZnO : Zn is 1 : 1

0.081 kg ZnO produces 0.065 kg Zn
x kg produces 115000 kg
Cross multiply and solve for x
\[
x = \frac{0.081 \times 115000}{0.065} = 143307 \text{ kg}
\]
\[
\therefore \text{ no. of tonnes produced} = \frac{143307 \text{ kg}}{1000 \text{ kg}} = 143.3 \text{ tonnes}
\]

Example 4
10.0 g calcium carbonate are heated strongly to a constant mass. Calculate the mass of the solid residue produced.
(Ca = 40.0; C = 12.0; O = 16.0)

Solution
You can solve this problem by using any of the 2 methods (a) or (b).
Write the equation first. The relative atomic masses given must be used.

Method (a): Using moles

\[
\text{CaCO}_3(s) \rightarrow \text{CaO(s) + CO}_2(g)
\]

1 mole  
0.1 mole  
1 mole  
y moles

Covert 10 g to moles first.
1 mol of CaCO$_3$ = 100 g
x mole of CaCO$_3$ = 10.0 g
The mass of CaO is calculated as follows:

\[
1 \text{ mole of CaO} = 56 \text{ g.}
\]

\[
0.1 \text{ moles of CaO} = x \text{ g.}
\]

\[
\therefore x = 0.1 \times 56 = 5.6 \text{ g.}
\]

Therefore the mass of residue is 5.6 g.

**Method (b):** Using grams

\[
\text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g)
\]

100 g CaCO\(_3\) produce 56 g CaO

10.0 g CaCO\(_3\) produce x g CaO

Cross multiply and solve for x

\[
\therefore x = \frac{10.0 \text{ g}}{100 \text{ g}} \times 56 \text{ g}
\]

\[= 5.6 \text{ g}\]

The mass of the solid residue is 5.6 g.

**Example 5**

When excess carbon(IV) oxide is reacted with sodium hydroxide solution, aqueous sodium carbonate is formed. If excess water is evaporated and the saturated solution allowed to cool, crystals of Na\(_2\)CO\(_3\) . 10H\(_2\)O can be formed. Find the mass of crystals that would be formed from 5 g of aqueous sodium hydroxide in excess water.

(Na = 23.0; C = 12.0; O = 16.0; H = 1.0)

**Solution**

\[
2\text{NaOH(aq)} + \text{CO}_2(g) + 9\text{H}_2\text{O(l)} \rightarrow \text{Na}_2\text{CO}_3 . 10\text{H}_2\text{O(s)}
\]

2 moles of NaOH produce 1 mole of crystals

But 1 mole of NaOH = 40 g and 1 mole Na\(_2\)CO\(_3\) . 10H\(_2\)O = 286 g

\[\therefore 2 \text{ moles of NaOH} = 2 \times 40 = 80 \text{ g}\]

This means

80 g of NaOH produce 286 g of Na\(_2\)CO\(_3\) . 10H\(_2\)O crystals

5 g of NaOH would produce x g of the crystals

Cross multiply and solve for x
\[ \therefore x = \frac{5 \text{ g} \times 286 \text{ g}}{80 \text{ g}} \]

\[ = 17.88 \text{ g} \]

Solve the same example, by finding the number of moles in 5 g sodium hydroxide first.

**Exercise 2.20**

1. A student made a solution containing 0.95 g of magnesium chloride, then added a solution of silver nitrate. Calculate the amount of silver chloride that would be obtained. (Ag = 108; Mg = 24; Cl = 35.5; N = 14; O = 16)

2. If you added 5.3 g of sodium carbonate to excess dilute sulphuric acid, what mass of sodium sulphate would be formed. (Na = 23.0; C = 12.0; O=16 S = 32.0)

3. Iron is obtained by reduction of iron(III) oxide by carbon(II) oxide in a blast furnace. Calculate the amount of iron(III) oxide that you would need to produce 100 tonnes of iron. (Fe = 56; O= 16; 1 tonne = 1000 kg)

**Acid-base titration**

We have learnt that when an acid react with a base or alkali; neutralisation reaction takes place. Neutralisation reactions can be done more accurately by a process known as titration. Titration is the process of adding a solution from a burette into a standard solution usually in a conical flask until the reaction between the two solutions is complete. An indicator is used to show the point at which the reaction is complete. The point at which the reaction is complete is called the end-point. The volume of the solution added from the burette is known as the titre. The volume transferred by the pipette into the conical flask is known as an aliquot.

**Apparatus used when titrating**

When we titrate, we use pipettes and burettes. This is because we deal with accurate volumes and also involves finding exact amounts. These type of experiment is also known as volumetric analysis. A pipette delivers fixed and exact volumes of liquid or solution, while a burette is used to add accurate amount of liquid or solution for complete neutralisation. Table 2.11 shows the apparatus commonly used in volumetric analysis and their uses.
Table 2.11: The apparatus commonly used in volumetric analysis and their uses

<table>
<thead>
<tr>
<th>Apparatus</th>
<th>Use</th>
</tr>
</thead>
<tbody>
<tr>
<td>• Balance</td>
<td>To weigh solids to be dissolved.</td>
</tr>
<tr>
<td>• Measuring cylinder</td>
<td>Measure volume of solution to be diluted.</td>
</tr>
<tr>
<td>• Volumetric flask</td>
<td>To make solution of known concentration.</td>
</tr>
<tr>
<td>• Pipette</td>
<td>To deliver fixed volumes of liquid or solution into conical flasks.</td>
</tr>
<tr>
<td>• Burette</td>
<td>To release small amounts of solutions into the conical flasks.</td>
</tr>
<tr>
<td>• Beaker</td>
<td>To collect and pour solutions into the burette.</td>
</tr>
<tr>
<td>• Funnel</td>
<td>To facilitate pouring solutions into the burette.</td>
</tr>
<tr>
<td>• White tile or paper</td>
<td>To help in the observation of colour change in the flask more clearly.</td>
</tr>
<tr>
<td>• Dropper</td>
<td>To add 2-3 drops of an indicator into the conical flask.</td>
</tr>
</tbody>
</table>

Handling titration apparatus

Burettes and pipettes are expensive apparatus and must be handled with a lot of care to avoid breakage.

Filling a pipette

- Wash the pipette thoroughly with water and then rinse it with the solution given i.e. the one you are supposed to fill the pipette with.
- Suck the solution up the pipette well above the calibration mark. It is recommended to suck the solution from a tilted conical flask to avoid air bubbles entering into the pipette as shown in Fig. 2.21 and to use a pipette filler to fill the pipette.
Note that when the solution falls below the mark, you have to re-fill all over again. This wastes time. To avoid this hold the pipette up right and read below the meniscus. See Fig. 2.22.

Allow the solution to run out into the conical flask. Do not force out the little amount of solution left in the pipette as the pipette is graduated with this allowance.

Add 2 - 3 drops of the indicator as shown in Fig. 2.23.
Fig. 2.23: Adding an indicator to a solution

**Filling a burette**

A burette is calibrated from top downwards i.e. 0 cm$^3$ at the top and 50 cm$^3$ at the bottom near the tap. Examine the burette and note that the divisions are 1 cm$^3$.

- Wash the burette with water, then rinse it with the given acid solution or the solution to be added to the burette.
- Clamp the burette to the stand in an upright position
- Using a filter funnel and a beaker fill the burette with the acid or solution to above the 0 cm$^3$ graduation mark. When filling the burette, hold the funnel with one hand so that air can escape as you fill the burette otherwise air will hold the acid in the funnel and when you lift the funnel, the acid quickly fills the burette and spills over. See Fig. 2.24.

Fig 2.24: Correct method of filling a burette
- Place a white tile or paper under the conical flask to clearly see the colour changes.
• Run the acid into the conical flask, 1 or 2 cm$^3$ at a time, fairly quickly. After every addition, swirl the flask, keeping the eye on the solution in the flask to see any new colour change. When you swirl the flask and the new colour developing persists for a while, the end point is near. Add one drop of the acid at a time until you get a permanent colour. This is the end-point of the titration.
• Read and record the volume of the acid used immediately, to the nearest 0.1 cm$^3$. Record your results as in Table 2.12.
• Repeat the titration two more times using a different clean conical flask. Add the same number of drops of the indicator as used in the first titration. However, if you use the same flask, you must wash it thoroughly with distilled water and then rinse with a little of the solution you intend to suck into the pipette.

**Note:** The first volume obtained should guide you to get the second volume fairly quickly as follows:
If your first titre volume was 25.0 cm$^3$, you may add 21 cm$^3$ quickly while shaking the flask. Then add the acid dropwise while swirling the conical flask until you get the same permanent colour change as in previous titration.
• Repeat the titration using the second volume as a guide to the volume required. You can add the acid until you are about 1 cm$^3$ from the end point, then you add the acid drop by drop. These steps have been illustrated in Fig. 2.25 below.
• Average any two values for the volume of the acid added, which differ by not more than $\pm$ 0.2 cm$^3$. In other words, the three titrations should be fairly consistent. Accuracy refers to how close measurements (i.e. volumes) are close to the “correct or teachers’ value.

*Fig. 2.25: Titration process*
Table 2.12: Record of volume of titration results

<table>
<thead>
<tr>
<th>Burette readings</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading</td>
<td>21.5</td>
<td>43.2</td>
<td>21.1</td>
</tr>
<tr>
<td>Initial burette reading</td>
<td>0.0</td>
<td>22.0</td>
<td>0.0</td>
</tr>
<tr>
<td>Volume of acid used (cm³)</td>
<td>21.5</td>
<td>21.2</td>
<td>21.1</td>
</tr>
</tbody>
</table>

• Calculate the average volume of the acid used and show your working.

The average volume of acid used = \( \frac{21.2 + 21.1}{2} \) = 21.15 cm³

Note: The value 21.5 is omitted because it is off the limit of ± 0.2 from the other readings. If you got values such as 24.2 cm³, 24.5 cm³ and 24.6 cm³, which two values would you use for getting the averages? 24.5 and 24.6 because the difference between them is not more than 0.2 cm³.

Many of us can perform the experiment fairly well, up to and including getting the average of the acid (or any other solution) used from the burette. But, some may not know how to use the values obtained to proceed with further calculations. Now let us do the following experiment to show how we are supposed to answer typical examination questions on volumetric analysis.

Experiment 2.7
Example of a practical titration.

Example 1
You are provided with, Solution A, which was made by dissolving 5.3 g dm⁻³ of sodium carbonate Na₂CO₃.
Solution B, which is hydrochloric acid in 1 dm³ solution.
You are required to:
(i) calculate the molarity of the sodium carbonate solution.
(ii) standardise solution B using solution A, i.e find its concentration.
(Na = 23; O = 16; C = 12)

Procedure
• Fill the burette with dilute hydrochloric acid provided.
• Pipette 25.0 cm³ of solution A and transfer into the conical flask. Add 2 - 3 drops of methyl orange indicator. Observe the colour change.
• Titrate until you get the end-point. How would you tell that the end-point has been reached?
• Repeat the titration 2 to 3 times to get consistent values.
• Record your values as in Table 2.13 below

Note: The initial colour of indicator in the sodium carbonate and stop adding the acid when you get a different colour. It will change from yellow to pink.

Table 2.13: Burette reading

<table>
<thead>
<tr>
<th>Burette reading</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Initial burette reading</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Volume of solution B used (cm³)</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

• Find the volume of the acid that reacted.  
• Standardize solution B using solution A.  
• Find the molarity of the sodium carbonate.

The following is a sample practical titration data and detailed calculation that shows you how to use your results after titration.

Table 2.14: Specimen results

<table>
<thead>
<tr>
<th>Burette readings</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette readings</td>
<td>22.5</td>
<td>44.4</td>
<td>21.8</td>
</tr>
<tr>
<td>Initial burette reading</td>
<td>0.0</td>
<td>22.5</td>
<td>0.0</td>
</tr>
<tr>
<td>Volume of solution B used (cm³)</td>
<td>22.5</td>
<td>21.9</td>
<td>21.8</td>
</tr>
</tbody>
</table>

(i) Calculate the average volume of solution B used. Show how you get your average.

\[
\frac{21.9 + 21.8}{2} = 21.85 \text{ cm}^3
\]

(ii) Write an equation for the reaction.

\[
\text{Na}_2\text{CO}_3(aq) + 2\text{HCl}(aq) \rightarrow 2\text{NaCl}(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)
\]

The equation gives us the mole ratio of the reactants i.e. moles of the sodium carbonate (solution A) and acid (solution B).

We should start by calculating the moles of solution A because it is the one whose volume (25.0 cm³) and concentration (5.3 g dm⁻³) are known. This is done as follows:
(i) 1 mole of Na₂CO₃ = 2(23) + 12 + 3(16) = 106 g
x moles = 5.3 g
Cross multiply and solve for x
\[ \frac{5.3}{106} = 0.05 \text{ mole} \]

The molarity of solution A is 0.05 mol dm⁻³ (0.05 M)

(ii) From part (i) we can say that:
1000 cm³ of solution A contained 0.05 mole
∴ 25 cm³ that we took in the pipette contained x moles
\[ x = \frac{0.05 \text{ mole} \times 25 \text{ cm}^3}{1000 \text{ cm}^3} \text{ mole} \]
\[ = 0.00125 \text{ mole of Na}_2\text{CO}_3 \]

Since mol ratio of Na₂CO₃ : HCl is 1 : 2
moles of HCl = 0.00125 \times 2 = 0.0025 mole

This means, from our average volume
21.85 cm³ contained 0.0025 mole of HCl
∴ 1000 cm³ contained x moles
Cross multiply and solve for x
\[ x = \frac{0.0025 \text{ moles} \times 1000 \text{ cm}^3}{21.85 \text{ cm}^3} = 0.114 \text{ moles} \]
\[ = 0.114 \text{ moles} \]

The concentration of HCl = 0.114 mol dm⁻³

Example 2
You are provided with
• Solution M, which is 0.5 M hydrochloric acid.
• Solution N, containing 5.3 g of a metal carbonate, X₂CO₃ dissolved in 250 cm³ distilled water in a volumetric flask.
• Indicator - methyl orange.
• Titration apparatus.
  (Na = 23; C = 12; O = 16)
You are required to determine the,
  (a) concentration (molarity) of solution N in
You solve this question by titration.
Pipette 25 cm³ of solution N and transfer into a conical flask. Add 2-3 drops of the indicator. As in example 1, get 3 concordant values i.e values that are close to each other. Suppose you obtained values as in the following Table 2.15.

Table 2.15: Speciment results

<table>
<thead>
<tr>
<th>Burette readings</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading</td>
<td>22.0</td>
<td>43.9</td>
<td>21.8</td>
</tr>
<tr>
<td>Initial burette reading</td>
<td>0.0</td>
<td>22.0</td>
<td>0.0</td>
</tr>
<tr>
<td>Volume of acid reading (cm³)</td>
<td>22.0</td>
<td>21.9</td>
<td>21.8</td>
</tr>
</tbody>
</table>

Calculate the average volume of solution M used. Show how you get your answer.
\[
\frac{21.9 + 21.8}{2} = 21.85 \text{ cm}^3
\]

The equation for reaction that occurred is;

\[ X_2CO_3(aq) + 2HCl(aq) \rightarrow 2XCl(aq) + H_2O(l) + CO_2(g) \]

1 : 2

Ionic equation

\[ CO_2^-(aq) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g) \]

Mole ratio 1 : 2

Note, we can use either of the equations.

(i) From the instructions given, we can calculate the moles of HCl in the, 21.85 cm³ because we are given the concentration of HCl, as 0.5 M.

1000 cm³ contain 0.5 moles of HCl
21.85 cm³ contain x moles of HCl
Cross multiply and solve for x

\[
\therefore x = \frac{0.5 \text{ moles} \times 21.85 \text{ cm}^3}{1000 \text{ cm}^3} = 0.01 \text{ moles HCl}
\]
Since the mole ratio of \( X_2CO_3 : HCl \) is
\[
\begin{align*}
1 & : 2 \\
y & : 0.01
\end{align*}
\]
Let \( y \) be moles of \( X_2CO_3 \) \( y : 0.01 \)
Cross multiply and solve for \( y \)
\[
\begin{align*}
2y &= 0.01 \\
y &= \frac{0.01}{2} = 0.005 \text{ moles of } X_2CO_3
\end{align*}
\]
This means the 25 cm\(^3\) of \( X_2CO_3 \) solution in the pipette contained 0.005 moles.
How many moles are in 1 dm\(^3\)?
\[
\begin{align*}
25 \text{ cm}^3 & \text{ contained } 0.005 \text{ moles} \\
1000 \text{ cm}^3 & \text{ contain } x \text{ mol.}
\end{align*}
\]
Cross multiply and solve for \( x \)
\[
\begin{align*}
\therefore x &= \frac{0.005 \times 1000}{25} = 0.2 \text{ moles} \\
&= 0.2 \text{ moles}
\end{align*}
\]
The molarity (moles per dm\(^3\)) of solution N = 0.2 mol dm\(^{-3}\) or 0.2 M
(ii) 250 cm\(^3\) of solution N contained 5.3 g
1000 cm\(^3\) of solution N contain \( y \) g
Cross multiply and solve for \( y \)
\[
\begin{align*}
250 \text{ cm}^3 \times y &= 5.3 \text{ g} \times 1000 \text{ cm}^3 \\
\therefore y &= \frac{5.3 \text{ g} \times 1000 \text{ cm}^3}{250 \text{ cm}^3} \\
&= 21.2 \text{ g}
\end{align*}
\]
The concentration in g dm\(^{-3}\) is 21.2 g dm\(^{-3}\)

**Note:** We have expressed the concentration of \( X_2CO_3 \) solution N, in mol dm\(^{-3}\) and g dm\(^{-3}\). Let us now proceed and solve part (b)

(b) 0.2 mol dm\(^{-3}\) of N = 21.2 g dm\(^{-3}\)
\[
\therefore 1 \text{ mol dm}^{-3} = x \text{ g dm}^{-3}
\]
Cross multiply and solve for \( x \)
\[
\begin{align*}
x &= \frac{21.2 \text{ g dm}^{-3}}{0.2 \text{ mol dm}^{-3}} = 106 \text{ g mol}^{-1}
\end{align*}
\]
The mass of 1 mol of \( X_2CO_3 \) = 106 g

OR Molar mass is 106 g
This means $X_2CO_3 = 106$
\[2X + 12 + (16 \times 3) = 106\]
\[2X + 60 = 106\]
\[2X = 106 - 60 = 56\]
\[X = 23\]

∴ R.A.M of $X = 23$

Exercise 2.21
1. 24.8 cm$^3$ of a solution containing 9.8 g dm$^{-3}$ sulphuric acid required 25 cm$^3$ of a solution containing 13.8 g dm$^{-3}$ of a metal carbonate ($M_2CO_3$), for complete neutralisation. Calculate the
   (i) Molarity of the acid.
   (ii) Molarity of the carbonate
   (iii) Formula mass of the carbonate
   (iv) The relative atomic mass of M.
   (H = 1; S = 32; O = 16; C = 12)

2. 25 cm$^3$ of a solution of an acid $H_xX$ containing 0.1 moles in 1 dm$^3$ of solution reacts with 75 cm$^3$ of 0.1 M KOH. What is the value of $y$?

Redox titration
A titration process where oxidation and reduction takes place at the same time is an example of a redox reaction.
Oxidation–reduction (redox) reactions involve two processes.
   (a) A process involving the loss of electrons from a reducing agent — oxidation.
      e.g. $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^{-}$
      $Na(g) \rightarrow Na^+(g) + e^{-}$
   (b) A process involving gain of electrons by an oxidising agent — reduction
      e.g. $Cl(g) + e^{-} \rightarrow Cl^-(g)$; $O(g) + 2e^{-} \rightarrow O^{2-}(g)$

In redox titration the commonly used oxidising agents are acidified solutions of the following compounds.
   (i) Potassium manganate(VII), $K$MnO$_4$.
   (ii) Potassium dichromate(VI), $K_2Cr_2O_7$.

They are referred to as acidified $K$MnO$_4$ and $K_2Cr_2O_7$, respectively. The acid used is sulphuric acid. Hydrochloric acid is not used because it could be oxidised to chlorine.
In these titrations, an indicator is unnecessary. The oxidising agents are used as *self – indicators* because they change colour as follows:

- Potassium manganate(VII) which is purple in colour turns colourless. Why does it turn colourless? The MnO$_4^-$ ions are reduced to Mn$^{2+}$ ions i.e.

\[
\text{MnO}_4^- (aq) \rightarrow \text{Mn}^{2+}(aq)
\]

(purple) colourless

If you titrate acidified potassium manganate(VII) solution against a solution containing iron(II) ions e.g. ammonium iron(II) sulphate solution or iron(II) sulphate, iron(II) ions are oxidised to iron(III) ions i.e. Fe(II) $\rightarrow$ Fe(III).

The chemical equation for the reactions as follows:

\[
2\text{KMnO}_4(aq) + 8\text{H}_2\text{SO}_4(aq) + 10\text{FeSO}_4(aq) \rightarrow \text{K}_2\text{SO}_4(aq) + 2\text{MnSO}_4(aq) + 5\text{Fe}_2(\text{SO}_4)_3 + 8\text{H}_2\text{O}(l)
\]

Ionic equation

\[
\text{MnO}_4^-(aq) + 18\text{H}^+(aq) + 15\text{Fe}^{2+}(aq) \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) + 5\text{Fe}^{3+}(aq)
\]

In the above ionic equation, manganese changes its charge from +7 in MnO$_4^-$ ions to +2 in Mn$^{2+}$ ions i.e. it is reduced.

\[
\text{MnO}_4^-(aq) + 8\text{H}^+(aq) + 5\text{e} \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) \quad \text{(reduction)}
\]

\[\text{Mn}^7+(aq) + 5\text{e} \rightarrow \text{Mn}^{2+}(aq)\]

i.e Mn(VII) $\rightarrow$ Mn(II)

While the charge of Fe$^{2+}$ ions change to Fe$^{3+}$ by loss of electrons i.e. they are oxidised as shown below.

\[
\text{Fe}^{2+}(aq) \rightarrow \text{Fe}^{3+}(aq) + \text{e} \quad \text{(oxidation)}
\]

iron(II) ions $\rightarrow$ Iron(III) ions

**Note:** You may not be required to write the equations in the box. However they will help you to get the mole ratio of the oxidising and reducing agents.

Let us standardise acidified potassium manganate(VII) solution using ammonium iron(II) sulphate solution by performing the following experiment.

**Experiment 2.8**

To standardise potassium manganate(VII) solution using ammonium iron(II) sulphate solution.
**Apparatus and chemicals**

- burette
- pipette, 25 cm³ (or 20 cm³)
- 3 conical flasks
- stand and clamp
- white tile/paper
- funnel
- acidified potassium manganate(VII) solution
- ammonium iron(II) sulphate solution.

**Procedure**

1. Pipette 25.0 cm³ of ammonium iron(II) sulphate solution and transfer into the conical flask.
2. Clamp your burette and fill it with acidified potassium manganate(VII) solution.
3. Titrate until a permanent colour just appears in the conical flask. What colour do you observe? What can you conclude from your observation?
4. Record your results in a table like the one shown in Table 2.16.
5. Repeat steps 1-4 to complete table.

<table>
<thead>
<tr>
<th>Table 2.16: Results of oxidation of Fe²⁺ ions with acidified KMnO₄ solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>Burette readings</td>
</tr>
<tr>
<td>Final burette reading</td>
</tr>
<tr>
<td>Initial burette reading</td>
</tr>
<tr>
<td>Volume of potassium manganate(VII) solution used (cm³)</td>
</tr>
</tbody>
</table>

**Exercises 2.22**

(a) Calculate the average volume of potassium manganate(VII) used. Show how you arrive at your answer.

(b) Calculate the molarity of ammonium iron(II) sulphate, (NH₄)₂Fe(SO₄)₂·6H₂O, given that 23.5 g were dissolved in 1 dm³.
(N = 14;  H = 1;  Fe = 56;  S = 32;  O = 16)

(c) Find the number of moles of Fe²⁺ ions that you pipetted. The ionic equation of the redox reaction that occurred is given below. MnO₄⁻(aq) + 5Fe²⁺(aq) + 8H⁺(aq) → Mn²⁺(aq) + 5Fe³⁺(aq) + 4H₂O(l)

(d) What is the mole ratio of manganate(VII) ions to iron(II) ions? Calculate the concentration of manganate(VII) ions in acidified potassium manganate(VII) solution.
**Hint:** Treat this redox titration reaction just as we did in acid-base titration, by using mole ratios. Check the worked example below and see how you are supposed to treat results from your experiment. Do not force your values to be the same as those in Table 2.17.

**Experimental example**
You are provided with solutions P and Q.
Solution P is acidified potassium manganate(VII)
Solution Q was prepared by dissolving 23.5 g of ammonium iron(II) sulphate, 
\((\text{NH}_4)_2 \text{Fe(SO}_4)_2 \cdot 6\text{H}_2\text{O}\), per dm³.

You are required to standardize i.e find the concentration of solution P.

**Procedure**
1. Fill the burette with solution P.
2. Adjust the volume to exactly zero mark.
3. Pipette 25.0 cm³ of solution Q, and transfer into a 250 cm³ conical flask.
4. Titrate solution P against solution Q until a permanent pink colour just appears.
5. Record your results in a table as shown below.
6. Repeat the procedure two more times.

**Note:** The values in Table 2.17 below are just sample results. You may get different values when you perform the experiment. Make sure you are accurate.

**Table 2.17: Specimen results**

<table>
<thead>
<tr>
<th>Burette readings</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading</td>
<td>16.9</td>
<td>32.9</td>
<td>16.1</td>
</tr>
<tr>
<td>Initial burette reading</td>
<td>0.0</td>
<td>16.9</td>
<td>0.0</td>
</tr>
<tr>
<td>Volume of solution P added (cm³)</td>
<td>16.9</td>
<td>16.0</td>
<td>16.1</td>
</tr>
</tbody>
</table>

(a) Average any two readings that are within ± 0.2 of each other. Show your working.

\[
\frac{16.0 + 16.1}{2} = \frac{32.1}{2} = 16.05 \text{ cm}^3
\]

(b) 1 mol of \((\text{NH}_4)_2 \text{Fe(SO}_4)_2 \cdot 6\text{H}_2\text{O}\) = (14 + 4)2 + 56 + (32 + 16 × 4)2 + 6(18) = 392

1 mole of \((\text{NH}_4)_2 \text{Fe(SO}_4)_2 \cdot 6\text{H}_2\text{O}\) = 392 g

\[
x \text{ moles of } (\text{NH}_4)_2 \text{Fe(SO}_4)_2 \cdot 6\text{H}_2\text{O} = 23.5 \text{ g}
\]

Cross multiply and solve for x
∴ \( x = \frac{23.5 \ g}{392} \times \) moles = 0.0599 moles

0.0599 moles was dissolved in 1 dm\(^3\)

∴ The molarity is 0.0599 mol dm\(^{-3}\) (or 0.0599 M)

(c) When 1 mole of ammonium iron(II) sulphate was dissolved, it formed the following ions:

\[
(NH_4)_2 \ Fe(SO_4)_2 \cdot 6H_2O(aq) \rightarrow 2NH_4^+(aq) + Fe^{2+}(aq) + 2SO_4^{2-}(aq)
\]

1 mole of ammonium iron(II) sulphate produces 1 mole of iron(II) ion

But from answer (b)

1000 cm\(^3\) of \((NH_4)_2 \ Fe(SO_4)_2 \cdot 6H_2O\) contained 0.0599 moles

25 cm\(^3\) (pipette volume) contained \(x\) moles

Cross multiply and solve for \(x\)

∴ \( x = \frac{0.0599 \text{ moles} \times 25 \text{ cm}^3}{1000 \text{ cm}^3} = 0.00149 \text{ moles} \)

∴ 25 cm\(^3\) that was pipetted contained about 0.0015 moles.

(d) \(MnO_4^-\)(aq) + 5Fe\(^{2+}\)(aq) + 8H\(^+\)(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)\)

Mole ratio of \(MnO_4^-\) ions to \(Fe^{2+}\) ions is 1 : 5

∴ We require about 0.0003 mole of \(MnO_4^-\) ions (i.e 0.0015 ÷ 5).

From the average volume obtained in (a), it implies that

16.05 cm\(^3\) contained 0.0003 moles of \(MnO_4^-\) ions

∴ 1000 cm\(^3\) would contain \(x\) moles of \(MnO_4^-\) ions

Cross multiply and solve for \(x\)

\[
x = \frac{0.0003 \text{ mol} \times 1000 \text{ cm}^3}{16.05 \text{ cm}^3} = 0.01869
\]

\[
x = 0.019 \text{ mol dm}^{-3}
\]

∴ The concentration \(MnO_4^-\) ions is 0.019 mol dm\(^{-3}\) or 0.019 M.
We can also use potassium dichromate(VI), \( \text{K}_2\text{Cr}_2\text{O}_7 \) to oxidise \( \text{Fe}^{2+} \) ions to \( \text{Fe}^{3+} \) ions. It is orange in colour but during oxidation turns green. However it is a less powerful oxidising agent than potassium manganate(VII).

The chemical equations for the reactions are as follows

\[
\text{K}_2\text{Cr}_2\text{O}_7\text{(aq)} + 7\text{H}_2\text{SO}_4\text{(aq)} + 6\text{FeSO}_4\text{(aq)} \rightarrow \text{K}_2\text{SO}_4\text{(aq)} + \text{Cr}_2(\text{SO}_4)_3\text{(aq)} + 3\text{Fe}_2(\text{SO}_4)_2\text{(aq)} + 7\text{H}_2\text{O}(l).
\]

**Ionic equation:**

\[
\text{Cr}_2\text{O}_7^{2–}\text{(aq)} + 6\text{Fe}^{2+}\text{(aq)} + 14\text{H}^+\text{(aq)} \rightarrow 2\text{Cr}^{3+}\text{(aq)} + 6\text{Fe}^{3+}\text{(aq)} + 7\text{H}_2\text{O}(l)
\]

(Redox reaction)

In the above ionic equation; chromium (Cr) changes its charge from +6 in \( \text{Cr}_2\text{O}_7^{2–} \) ions to + 3 in \( \text{Cr}^{3+} \) ions i.e. it is reduced

\[
\text{Cr}_2\text{O}_7^{2–}\text{(aq)} + 14\text{H}^+(aq) + 6e– \rightarrow 2\text{Cr}^{3+}\text{(aq)} + 7\text{H}_2\text{O} \quad \text{(reduction)}
\]

(orange)  \quad (green)

i.e \( \text{Cr}^{6–} \) ion  \quad \rightarrow \text{Cr}^{3+} \) ions

The change of \( \text{Fe}^{2+} \) ions change to \( \text{Fe}^{3+} \) ions is by loss of electrons i.e they are oxidized as shown below.

\[
\text{Fe}^{2+}\text{(aq)} \rightarrow \text{Fe}^{3+}\text{(aq)} + e^– \quad \text{(oxidation)}
\]

iron(II)  \quad \rightarrow  \quad iron(III)

**NB:** You are not required to write the equations in this box. However, they will help you to get the mole ratios of the oxidising and reducing agents.

**Exercise 2.23**

1. During a titration experiment 25.0 cm\(^3\) of a solution of hydrogen peroxide, \( \text{H}_2\text{O}_2 \), required 44.2 cm\(^3\) of 0.02 mol dm\(^{-3}\) of acidified potassium manganate(VII), \( \text{KMnO}_4 \), solution for complete reaction.

Calculate the concentration (molarity) of hydrogen peroxide solution. The overall ionic reaction is given below.

\[
2\text{MnO}_4^{–}\text{(aq)} + 5\text{H}_2\text{O}_2\text{(aq)} + 6\text{H}^+(aq) \rightarrow 2\text{Mn}^{2+}\text{(aq)} + 8\text{H}_2\text{O}(l) + 5\text{O}_2(g)
\]

**2.6: Molar gas volume**

The molar volume of a gas is the volume occupied by one mole of the gas at standard temperature and pressure. Gases are very light and can be difficult to work with. It
is difficult to weigh them. When heated they expand rapidly. When under pressure they contract easily. See Fig. 2.26.

*Fig. 2.26: Weighing a gas*

However, we can quite easily measure the volume of a gas using a syringe, a measuring cylinder or a graduated gas jar shown in Fig. 2.27.

*Fig 2.27: Apparatus for measuring and collecting gas*

But how does the gas volume relate to the number of atoms or molecules present? Equal volume of gases contain the same number of particles i.e. $6.023 \times 10^{23}$ particles (at the same temperature and pressure). It is the same for all gases under the same conditions of temperature and pressure.

**How much is the molar volume of gas at s.t.p?**

We are familiar with the following relationship, of the density, the mass and the volume.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$
Experiments have shown that the density of hydrogen is 0.09 g per 1 dm³. Can we calculate the relative molecular mass of H₂, given R.A.M of H = 1.008? Using the above formula we can then find its volume at s.t.p.

Let us consider the following example.

**Example 1**

Find the molar volume of hydrogen at s.t.p.
(H = 1.008; density of hydrogen is 0.09 g/dm³)

**Solution**

\[
1 \text{ mol of } H_2 = 1.008 \times 2 = 2.016 \text{ g}
\]

\[
D = \frac{M}{V} \quad D = \text{density} \\
M = \text{mass} \\
V = \text{volume}
\]

Substitute the values 0.09 = \frac{2.016g}{V} 1 \text{ mol of } H_2 = 2 \times 1.008 = 2.016 \text{ g} 0.09 g = 1000 dm³

\[
\therefore V = \frac{2.016}{0.09 \text{ g dm}^{-3}}
\]

OR

\[
\therefore x = \frac{2.016 \times 1 \text{ dm}^3}{0.09 \text{ g}}
\]

Molar volume of hydrogen = 22.4 dm³ = 22.4 dm³

**Example 2**

Given the density of oxygen is 1.43 g/dm³ at s.t.p, and its R.A.M is 16. Find the molar volume of oxygen at s.t.p.

\[
D = \frac{M}{V}
\]

1 mole of oxygen gas, O₂ = 16 × 2 = 32 g

\[
\therefore V = \frac{M}{D}
\]

\[
= \frac{32}{1.43} = 22.4 \text{ dm}^3
\]

Copy Table 2.18. Calculate and fill the molar volumes of the named gases.
Table 2.18: Molar gas volumes

<table>
<thead>
<tr>
<th>Gas</th>
<th>density (g/dm$^3$) at s.t.p</th>
<th>Relative atomic mass</th>
<th>Molar gas volume at s.t.p</th>
</tr>
</thead>
<tbody>
<tr>
<td>N$_2$</td>
<td>1.25</td>
<td>28</td>
<td></td>
</tr>
<tr>
<td>CO$_2$</td>
<td>1.81</td>
<td>44</td>
<td></td>
</tr>
<tr>
<td>CO</td>
<td>1.15</td>
<td>28</td>
<td></td>
</tr>
</tbody>
</table>

Note that scientists use cubic decimeters (1 dm$^3$) as the unit for volume. One litre is equal to 1000 cm$^3$ i.e. 1 dm$^3$ = 1L = 1000 cm$^3$.

In fact, results from experiments show that one mole of any gas at 25°C (298 K) and 1 atmosphere pressure, occupy 24 dm$^3$. These conditions are known as room temperature and pressure (r.t.p). One mole of gas occupies 22.4 dm$^3$ at standard temperature and pressure (s.t.p) i.e at 0°C (273K) and 1 atmosphere pressure. Because of this, the volume occupied by 1 mole of a gas is known as the molar volume i.e molar gas volume.

(a) At r.t.p = 24 dm$^3$
(b) At s.t.p = 22.4 dm$^3$

Calculations involving molar volume
Whenever gases are involved in a reaction, we need to remember that:
- they have mass and volume.
- 1 mole occupies 22.4 dm$^3$ at s.t.p and 24 dm$^3$ at r.t.p.

Example 1
How many moles of hydrogen gas are there in 24 cm$^3$ of hydrogen at r.t.p? (molar gas volume at r.t.p = 24 dm$^3$).

We must use same units, either dm$^3$ or cm$^3$. Do not mix units.

1 mole of hydrogen gas occupies 24 dm$^3$ at r.t.p.

\[ x \text{ moles of hydrogen gas occupy } 0.024 \text{ dm}^3 \text{ at r.t.p} \]

Cross multiply and solve for \( x \)

\[ \therefore x = \frac{0.024 \text{ dm}^3}{24 \text{ dm}^3} = 0.001 \text{ moles.} \]

OR

1 mole of hydrogen gas occupies 24,000 cm$^3$

\[ x \text{ moles of hydrogen gas occupy } 24 \text{ cm}^3 \]
x \quad = \quad \frac{24 \text{ cm}^3}{24000 \text{ cm}^3} \quad = \quad 0.001 \text{ moles}

**Exercise 2.24**

How many moles of molecules are there in the following volumes at r.t.p?

(a) 2.4 dm\(^3\) of oxygen gas
(b) 2.4 cm\(^3\) of nitrogen gas
(c) 24 dm\(^3\) of chlorine gas.

**Example 2**

What volume in cm\(^3\) does 0.01 moles of oxygen occupy at s.t.p?

(molar gas volume at s.t.p = 22.4 dm\(^3\)).

1 mole of oxygen occupies 22.4 dm\(^3\)
0.01 moles of oxygen will occupy x dm\(^3\)

\[ \therefore x = \frac{22.4 \text{ dm}^3 \times 0.01 \text{ mole}}{1 \text{ mole}} = 0.224 \text{ dm}^3 \]

= 0.224 dm\(^3\)

But 1 dm\(^3\) = 1000 cm\(^3\)

0.224 dm\(^3\) = 0.224 \times 1000 cm\(^3\) = 224 cm\(^3\)

**Exercise 2.25**

What volume do these gases occupy at room temperature and pressure?

(a) 4 moles of hydrogen
(b) 0.1 moles of gas Y.
(c) 0.005 moles of carbon(IV) oxide.

(molar volume of any gas at r.t.p = 24 dm\(^3\))

**Example 3**

Calculate the volume that 8 g of oxygen gas would occupy at standard temperature and pressure.

(O = 16, molar volume any gas s.t.p = 22.4 dm\(^3\))

Remember that oxygen is a diatomic gas. The formula of oxygen is O\(_2\).

R.F.M of oxygen is \(16 \times 2 = 32\)

1 mole of oxygen = 32 g

x mole of oxygen = 8 g

Cross multiply and solve for x

\[ \therefore x = \frac{8 \text{ g} \times 1 \text{ mole}}{32 \text{ g}} = 0.25 \text{ moles} \]

= 0.25 moles
1 mole of any gas occupies 22.4 dm$^3$ at s.t.p.

0.25 moles of oxygen gas will occupy $y$ dm$^3$ at s.t.p.

Cross multiply and solve for $y$

$$y = \frac{22.4 \text{ dm}^3 \times 0.25 \text{ mole}}{1 \text{ mole}}$$

$$y = 5.6 \text{ dm}^3$$

OR

32 g of oxygen occupy 22.4 dm$^3$ at s.t.p.

1 g of oxygen occupies $\frac{22.4}{32}$ dm$^3$ at s.t.p.

∴ 8 g of oxygen would occupy $\frac{22.4}{32} \times 8$ dm$^3$ at s.t.p

$$= 5.6 \text{ dm}^3$$

**Exercise 2.26**

What volume at r.t.p, would each of the following gases occupy.

The R.A.Ms are given below.

(a) 4 g of hydrogen
(b) 71.0 g of chlorine
(c) 17.6 g of carbon(IV) oxide.
(d) 0.004 g of helium,

(H = 1; Cl = 35.5; C = 12; O = 16; He = 2)

**Example 4**

How many moles of atoms are there in each of the following?

(a) 4.8 dm$^3$ of nitrogen at r.t.p
(b) 11.2 dm$^3$ of chlorine at s.t.p.

(molar gas volume at s.t.p = 22.4 dm$^3$ and 24 dm$^3$ at r.t.p)

24 dm$^3$ at r.t.p contain 1 mole of nitrogen molecules

1 dm$^3$ at r.t.p contain $\frac{1}{24}$ mole of nitrogen molecules

∴ 4.8 dm$^3$ at r.t.p $\frac{1}{24} \times 4.8 = 0.2$ mol.

Remember nitrogen(N$_2$) is diatomic

1 mole of N$_2$ consist 2 moles of N atoms

∴ 0.2 moles of N$_2$ contains x mole of N atoms

Cross multiply to get x

$$x = 2 \times 0.2 = 0.4$$ moles
(b) 22.4 dm$^3$ contain 1 mole of chlorine molecules at s.t.p.
11.2 dm$^3$ contain x moles of chlorine molecules
Cross multiply and solve for x.

\[
x = \frac{11.2 \text{ dm}^3 \times 1 \text{ mole}}{22.4 \text{ dm}^3} = 0.5 \text{ moles Cl}_2 \text{ molecules}
\]

1 mole of Cl$_2$ consists of 2 moles of Cl atoms
∴ 0.5 moles of Cl$_2$ contain x moles of Cl atoms
Cross multiply and solve for x

\[
x = \frac{0.5 \text{ mole} \times 2 \text{ moles}}{1 \text{ mole}} = 1.0 \text{ mole}
\]

**Calculation using relative molecular mass (RMM) and molar gas volume**

**Example 1**
In an experiment, a student collected 510 cm$^3$ of a gas X at s.t.p. The mass of the gas was 1g. Calculate the relative molecular mass of the gas.

Molar volume at s.t.p = 22.4 dm$^3$.
510 cm$^3$ of gas X at s.t.p weigh 1 g
1 cm$^3$ of gas X at s.t.p weigh $\frac{1}{510}$ g.

∴ 22400 cm$^3$ of gas Y would weigh $\frac{22400}{510} = 43.9$ g

The R.M.M of gas is about 44

**Example 2**
Calculate the relative molecular mass of a gas if whose 100 cm$^3$ of the gas volume has a mass of 0.067 g at r.t.p.

(Molar gas volume at r.t.p = 24 dm$^3$)
100 cm$^3$ of the gas has mass = 0.067 g

1 cm$^3$ of the gas has a mass = $\frac{0.067 \text{ g}}{100 \text{ cm}^3}$

∴ 24,000 cm$^3$ must have a mass = $\frac{0.067 \text{ g} \times 24000 \text{ cm}^3}{100 \text{ cm}^3} = 16.08$ g

The relative molecular mass is 16.1
Exercise 2.27
Calculate the relative molecular mass of the following gases
(a) 2.2 g of a gas A that occupied a volume of 1120 cm$^3$ at s.t.p.
(b) The mass of 500 cm$^3$ of a gas at s.t.p is 1.5 g

Atomicity of gases
An atom is the smallest particle of an element. The atoms take part in chemical changes in whole numbers. A molecule of an element or compound contains at least two atoms. An exception to this are noble gases. Atoms of noble gases are separate.

Definition
Atomicity of an element is the number of atoms in one molecule.

Examples
The atomicity of the noble gases is 1. They are called monatomic gases

- Noble gases include helium, neon, argon and krypton. They are represented as single atoms He, Ne, Ar, Kr etc. See Fig. 2.28(a).

- Other gases, for example, hydrogen, oxygen, nitrogen and the halogens exist as molecules containing 2 atoms, in each molecule. They are called diatomic gases and are represented in equations as H$_2$, O$_2$, N$_2$, Cl$_2$, Br$_2$ and I$_2$. See Fig. 2.28(b).
- Ozone (O$_3$) is triatomic and phosphorus vapour (P$_4$) is tetratomic.

Gay – Lussac’s Law and volumes of gases in reactions
In 1808, a scientist called Joseph Gay-Lussac did a series of experiments with volumes of reacting gases. The results showed simple interesting relationships as shown in the following examples.
(a) 1 volume of hydrogen reacts with 1 volume of chlorine to produce 2 volumes of hydrogen chloride.
How do gas volumes combine in reactions?

Students in a certain school performed the following experiment to find out how gas volumes combine in reactions. They set up the apparatus as illustrated Fig.2.29.

They flushed out syringe Y with dry ammonia and collected 20 cm³ of the gas in it. 20 cm³ of hydrogen chloride were collected in syringe Q in a similar manner. The three
way tap was finally turned such that the two gases were mixed by alternately pushing in each plunger. The pushing was repeated several times until no further change in the volume was detected. All the residual gas was then pushed into one syringe and the apparatus was allowed to cool. The readings were noted and recorded.

The results of the experiments were as follows.

Initial volume of dry ammonia = 20 cm$^3$
Initial volume of hydrogen chloride = 60 cm$^3$
Final reading of the syringe = 40 cm$^3$.

**Conclusion**

Volume of ammonia that reacted is 20 cm$^3$.
Volume of hydrogen chloride that reacted is 20 cm$^3$.

- The results from the above experiment show that equal volumes of ammonia and hydrogen chloride react to produce a white solid, ammonium chloride.
- The equation of the reaction is
  \[ \text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4 \text{Cl}(s) \]
- These two gases react in the ratio 1 : 1 which is a simple whole number ratio. If the product of the reaction had been a gas, its volume would have been in a simple ratio with the volume of the reactants.

The results of the experiments also show that the volumes of other gases which react do so in simple whole number ratios. Let us consider the following examples.

**Example 1**

A mixture of 30 cm$^3$ of carbon(II) oxide and 15 cm$^3$ of oxygen were ignited by a spark in an enclosed container. What volume of the product was formed?. (All volumes are at the same temperature and pressure).

\[ 2\text{CO}(g) + \text{O}_2(g) \rightarrow 2\text{CO}_2(g) \]

\[ \begin{array}{c}
2\text{vols} \\
30 \text{ cm}^3
\end{array} + \begin{array}{c}
1 \text{ vol} \\
15 \text{ cm}^3
\end{array} \rightarrow \begin{array}{c}
2 \text{ vols} \\
30 \text{ cm}^3
\end{array} \]

\[ \therefore \text{30 cm}^3 \text{ of CO}_2 \text{ were formed.} \]
Example 2

300 cm$^3$ of ethene were mixed with 120 cm$^3$ of oxygen and the mixture exploded.

$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$$

After the reaction was complete, what was the volume of the remaining gas. (volumes are at room temperature and pressure).

$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$$

<table>
<thead>
<tr>
<th>1 vol</th>
<th>3 vols</th>
<th>2 vols</th>
</tr>
</thead>
<tbody>
<tr>
<td>300 cm$^3$</td>
<td>120 cm$^3$</td>
<td></td>
</tr>
</tbody>
</table>

Note

If we had 3 cm$^3$ of O$_2$(g) it could only support burning of 1 cm$^3$ of C$_2$H$_4$(g). Therefore 120 cm$^3$ oxygen requires x cm$^3$ ethene.

$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$$

<table>
<thead>
<tr>
<th>1 vol</th>
<th>3 vols</th>
<th>2 vols</th>
</tr>
</thead>
<tbody>
<tr>
<td>x cm$^3$</td>
<td>120 cm$^3$</td>
<td></td>
</tr>
</tbody>
</table>

Cross multiply and solve for x

\[3x = 120\]

\[x = \frac{120}{3} = 40 \text{ cm}^3\] of ethene will be burnt

This means the amount of ethane that will remain is \((300 - 60) = 240 \text{ cm}^3\).

From the equation 1 vol of C$_2$H$_4$(g) forms 2 vols of CO$_2$(g)

\[\therefore 40 \text{ cm}^3\] of C$_2$H$_4$(g) will form 80 cm$^3$ of CO$_2$(g)

Note: Water is excluded because it will be a liquid at room temperature.

In these and other experiments, the gas reactions where explosions were triggered by a spark in an enclosed container, Gay-Lussac found that gases react in simple whole number units of volumes provided that after the explosion the products are cooled back to the temperature and pressure of the initial gases. The simple relationship shown by these results led to Gay-Lussac’s law of gaseous volumes which states that

When gases react, they do so in volumes which bear a simple ratio to one another and to the volume of the products, if gaseous, provided all the volumes are measured at the same temperature and pressure.

Exercise 2.28

1. (a) What volume of ammonia would be formed if 10 cm$^3$ of nitrogen were reacted with 20 cm$^3$ of hydrogen?
(b) What will be the volume of any of the remaining gas?

2. Sulphur(VI) oxide is produced when sulphur(IV) oxide is reacted with oxygen.
   (a) What volume of sulphur(VI) oxide would be formed when 100 cm$^3$ of sulphur(IV) oxide is completely reacted with oxygen?
   (b) What volume of oxygen would be needed to react with the sulphur(IV) oxide.
   (Hint: write balanced chemical equations first)

**Avogadro’s law and molar gas volume**

In 1811 based on experiments on gases, an Italian scientist, Amedeo Avogadro made one of the most important hypothesis in the development of atomic theory. It is now called **Avogadro’s law**.

Avogadro’s law states that equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

We have already seen that the molar volume of any gas at standard temperature and pressure (s.t.p) is 22.4 dm$^3$ and will always contain $6.023 \times 10^{23}$ molecules as shown in Fig. 2.30. The molar masses of the gases will be different but the number of particles will be $6.023 \times 10^{23}$. This number is called the Avogadro’s constant and is usually represented by letter ‘L’ where $L = 6.023 \times 10^{23}$ particles.

<table>
<thead>
<tr>
<th></th>
<th>Molar mass</th>
<th>Molecular Mass</th>
<th>He</th>
<th>$6.023 \times 10^{23}$</th>
<th>H$_2$</th>
<th>$6.023 \times 10^{23}$</th>
<th>O$_2$</th>
<th>$6.023 \times 10^{23}$</th>
<th>CO$_2$</th>
<th>$6.023 \times 10^{23}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Molar mass</td>
<td>4 g</td>
<td>2 g</td>
<td>32 g</td>
<td>44 g</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Fig. 2.30: Equal volumes of gases though with different molar masses, contain same number of molecules

We can use Avogadro’s constant to calculate the number of molecules in given volumes as shown in the following examples.

**Example 1**

How many molecules are there in

(a) 1.2 dm$^3$ of hydrogen at s.t.p?
(b) 1.2 dm$^3$ of carbon(IV) oxide at r.t.p?

$L = 6.0 \times 10^{23};$ molar volume of any gas at s.t.p = 22.4 dm$^3$. at r.t.p. = 24 dm$^3$)

**Solution**

(a) Molar volume of any gas at s.t.p = 22.4 dm$^3$
This implies that:
22.4 dm$^3$ of any gas contain $6.0 \times 10^{23}$ molecules
1.2 dm$^3$ of hydrogen contain $x$ molecules
Cross multiply and solve for $x$

$$\therefore x = \frac{6.0 \times 10^{23} \text{ molecules} \times 1.2 \text{ dm}^3}{22.4 \text{ dm}^3}$$

$$= 3.2 \times 10^{22} \text{ molecules}$$

(b) 24 dm$^3$ of any gas contain $6.0 \times 10^{23}$ molecules
1.2 dm$^3$ of carbon(IV) oxide contain $x$ molecules
Cross multiply and solve for $x$

$$\therefore x = \frac{6.0 \times 10^{23} \text{ molecules} \times 1.2 \text{ dm}^3}{24 \text{ dm}^3}$$

$$= 3.0 \times 10^{22} \text{ molecules.}$$

Example 2
A student collected 48000 cm$^3$ of a diatomic gas Z at r.t.p. Find the
(a) moles of gas Z collected
(b) mass of the gas collected.
(c) Total number of atoms collected.

(d) (molar volume at r.t.p = 24 dm$^3$; R.A.M of gas Z = 16.0; L = $6.0 \times 10^{23}$)

(a) 1 mole of any gas occupies 24 dm$^3$ (24 000 cm$^3$) at r.t.p
$x$ moles of gas occupy 48000 cm$^3$
Cross multiply and solve for $x$

$$\therefore x = \frac{48000 \text{ cm}^3}{24000 \text{ cm}^3} = 2 \text{ moles}$$

The student collected 2 moles.

(b) Since gas Z is diatomic, its formula is $Z_2$.
1 mole of $Z_2$ has a mass $16 \times 2 = 32$ g
$$\therefore 2 \text{ mole of gas collected has mass } = 32 \times 2 = 64 \text{ g.}$$

(c) From (a) the moles collected
$$= 2 \text{ mols}$$
The molecules collected
$$= 2 \times 6.0 \times 10^{23}$$
$$= 1.2 \times 10^{24} \text{ molecules}$$

Since gas Z is diatomic, it implies
1 mole of gas Z contains 2 moles of Z atoms
$$\therefore \text{ The total number of atoms } = 1.2 \times 10^{24} \times 2$$
$$= 2.4 \times 10^{24} \text{ atoms}$$
Exercise 2.29

1. Calculate the number of:
   (a) moles
   (b) molecules
   (c) atoms in 9 dm$^3$ of fluorine at s.t.p.

2. The molar mass of a gaseous compound XO$_2$ is 44 g. If this gas occupied 5.6 dm$^3$ at s.t.p. Find:
   (a) the number of moles
   (b) the mass in grams of the gas collected and
   (c) the number of molecules in 5.6 dm$^3$.

Summary

1. The mole is a unit used by chemists to measure quantities of chemical substances.
2. The mole is a number. It is like “a dozen” but a lot bigger
3. The mole is the number of particles in 12,000 g of carbon-12 isotope. This number is called Avogadro’s number or constant and is equal to $6.0 \times 10^{23}$. Its symbol is L.
4. One mole of any gas occupies 22.4 cm$^3$ at s.t.p and 24 dm$^3$ at r.t.p
5. One mole of solute dissolved in one cubic decimetre of solution has a concentration (molarity) of one mole per cubic decimetre i.e. 1 mol dm$^{-3}$ (M).
6. Before you write an equation you should know the reacting substances and products formed in the reaction.
7. To convert a word equation to formulae equation remember the correct symbols and their valencies.
8. The relative atomic mass (R.A.M) and relative formula mass (R.F.M) have no units.
9. A balanced chemical equation is called a stoichiometric equation.
10. We use mole ratio in balanced equations to calculate both reacting quantities and quantities produced.
11. When we write ionic equations we choose the ions that form what we see e.g. a precipitate, gas or liquid and leave out spectator ions.
12. During titration, the solution pipetted is always transferred into the conical flask.

Revision Exercise 2

1. Write the formulae of the following compounds from the name given and calculate formula masses.
   (a) ammonium phosphate
   (b) Iron(II) sulphate
(c) Lead(IV) chloride
(d) copper(II) nitrate
(e) copper(II) sulphate
(f) calcium hydrogencarbonate

2. Calculate the volume of ammonia in dm³ at r.t.p needed to produce 40 g of ammonium nitrate when it is reacted with nitric acid.

3. Calculate the mass of
   (a) 0.5 mol sodium hydroxide
   (b) 0.1 mol magnesium nitrate
   (c) 0.2 mol lead(IV) oxide
   (d) 2 mols zinc oxide.

4. Calculate the percentage by mass of nitrogen in each of the following compounds.
   (a) \((NH_4)_2SO_4\)
   (b) \(NaNO_3\)
   (c) \(Ca(NO_3)_2\)
   (d) \(KNO_3\)

5. Calculate the percentage of each element in calcium phosphate.

6. Calculate the empirical formulae for compounds containing:
   (a) \(Na = 32.4\%, \ S = 22.5\%\); \(O = 45.1\%\)
   (b) \(Fe = 20.14\%; \ S = 11.51\%\); \(O = 63.31\%, \ H = 5.04\%\)
   (c) \(C = 75\% \ H = 25\%\)

7. A hydrocarbon has a relative molecular mass of 42. It has the following composition. Carbon 85.7% and hydrogen 14.3%. Find its molecular formula.

8. On analysis of a compound of iron and oxygen, it was found to contain 3.4 g of iron and 1.5 g oxygen. Its relative molecular mass is 160. Find the molecular mass of the oxide.

9. 1.34 g of aluminium chloride was added to silver nitrate solution, 4.31 g of silver chloride precipitate were deposited. Calculate the empirical formula of the aluminium chloride.

10. A student decomposed 1 g of silver nitrate solid. Silver metal, nitrogen(IV) oxide gas and oxygen were produced. Calculate the total volume of gases produced at r.t.p.
11. Write ionic equations for the following reactions.
   (a) Magnesium with dilute sulphuric acid.
   (b) Barium chloride solution with sodium sulphate solution.
   (c) Zinc with silver nitrate solution.
   (d) Copper(II) oxide with dilute hydrochloric acid.
   (e) Sodium carbonate solution with nitric acid.

12. Acidified potassium manganate(VII) solution oxidizes iron(II) ions to iron (III)
    ions as shown in the ionic equation below. If 0.2 mole dm⁻³ KMnO₄ solution
    is needed to react with 25 cm³ of 0.1 mol dm⁻³ of iron(II) ammonium sulphate.
    Calculate the volume of KMnO₄ solution required.
    \[
    \text{MnO}_4^-(aq) + 5\text{Fe}^{2+}(aq) + 8\text{H}^+(aq) \rightarrow \text{Mn}^{2+}(aq) + 5\text{Fe}^{3+}(aq) + 4\text{H}_2\text{O}(l)
    \]

13. 25 cm³ of a solution containing 4 g dm⁻³ of NaOH reacts with 50 cm³ of a
    solution of hydrochloric acid. Calculate the molarity of the acid.

14. In a titration experiment, a student pipetted 25.0 cm³ of ammonia solution and
    titrated with 0.1 M HCl. The calculated average volume of acid was 33.30 cm³
    What was the concentration of ammonia solution in
    (a) mol dm⁻³
    (b) g dm⁻³?

15. Iron(II) ammonium sulphate solution was reacted with acidified potassium
    manganate(VII) solution. 25.0 cm³ of the Fe²⁺ ion solution required 24.8 cm³ of
    0.02 M KMnO₄ solution. Calculate the molarity of iron(II) ammonium sulphate
    solution.

16. 25.0 cm³ of a solution of an acid, H₂X, containing 0.1 mol dm⁻³ reacted with 75
    cm³ of a solution of 0.1 mol dm⁻³ NaOH. What is the basicity of the acid?
    \( \text{Hint: Find the value of } a. \)

17. Ethane dioic (oxalic acid) acid crystals have the formula H₂C₂O₄ . nH₂O. A
    solution of this acid was made by dissolving 10.1 g and making it to 1 dm³.
    During titration it was found that 25.0 cm³ of the acid solution required 20 cm³
    of 0.2 mol dm⁻³ sodium hydroxide solution. Find the:
    (a) Concentration of the acid in moles per dm³.
    (b) value of n, which gives the number of moles of water of crystallisation.

18. 15 cm³ of 0.02 M silver nitrate solution were added to 10 cm³ of a sodium
    chloride solution.
    (a) write an ionic equation for the reaction
    (b) calculate the concentration of sodium chloride solution in g dm⁻³
What is organic Chemistry?

Organic chemistry refers to the chemistry of compounds of carbon. Compounds such as carbon(II) oxide, CO, carbon(IV) oxide, CO₂, and those of carbonates are studied under the branch of chemistry known as inorganic chemistry. Although such compounds contain carbon atoms, the atoms do not link to form chains.

Other carbon-containing compounds such as bottled liquid petroleum gas (LPG), kerosene, petrol among others are studied under organic chemistry. Such compounds are known as hydrocarbons.

3.1: Hydrocarbons

What are ‘hydrocarbons’? As the name suggests, hydrocarbons are compounds which contain hydrogen and carbon only. A hydrocarbon has the molecular formula, CₓHᵧ where x and y represent whole numbers.

One of the world’s most important raw materials is crude oil or petroleum. Crude oil is a mixture of hydrocarbons. The hydrocarbons in crude oil are important fuels and raw materials for many important new products such as plastics and fabrics.

Hydrocarbons can be classified according to their structure. In this unit we shall study the main hydrocarbons which are classified according to the number of covalent bonds between two carbon atoms. Three classes of hydrocarbons are:

1. Alkanes - contain single bonds.
2. Alkenes - contain double bonds.
3. Alkynes - contain triple bonds.

3.2: Alkanes

There is a variety of hydrocarbons in crude oil. Most of them are alkanes. Examples of alkanes which are found in crude oil are methane, ethane, propane and butane. These four alkanes are gases. Crude oil also contains petrol and kerosene.

General formula, nomenclature and structural formula of alkanes

Alkanes are hydrocarbons with the general formula CₓH₂ₓ₊₂ where n denotes the number of carbon atoms in each molecule of an alkane. The number, n = 1,2,3 ..... The simplest alkane has only one carbon atom in each molecule. From the general formula shown above, we can get the formula and structure of the alkane. For example; where number of carbon atoms is one i.e n = 1. The formula is obtained as follows:
The first member of the alkane family has the formula CH\(_4\). It is called **methane**.

**Note:** When a subscript is 1, it is not shown in the formula.

The four hydrogen atoms of methane are arranged around the carbon atom as shown below. This is called a **structural formula**.

\[
\begin{align*}
\text{H} \\
\text{H} - \text{C} - \text{H} \\
\text{H}
\end{align*}
\]

methane

The second alkane has two carbon atoms, that is n = 2. Substituting n = 2 in the general formula gives the formula for the second alkane as follows:-

\[
\begin{align*}
\text{C}_n \text{H}_{2n+2} &= \text{C}_2 \text{H}_{(2\times 2 + 2)} \\
&= \text{C}_2 \text{H}_6 \\
&= \text{C}_2 \text{H}_6
\end{align*}
\]

Therefore the formula for the second alkane is C\(_2\)H\(_6\). It is called **ethane**. The structural formula of ethane is as shown below.

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} - \text{C} - \text{C} - \text{H} \\
\text{H} & \quad \text{H}
\end{align*}
\]

ethane

The third alkane has three carbon atoms that is n = 3. We then substitute for n in the general formula:

\[
\begin{align*}
\text{C}_n \text{H}_{2n+2} &= \text{C}_3 \text{H}_{(2\times 3 + 2)} \\
&= \text{C}_3 \text{H}_8 \\
&= \text{C}_3 \text{H}_8
\end{align*}
\]

The formula of the third alkane is therefore C\(_3\)H\(_8\) and it is called **propane**. The structural formula of propane is shown below;
Similarly, if we derive the fourth alkane, where \( n = 4 \), we get \( \text{C}_4\text{H}_{10} \). This alkane is called \textbf{butane}.

**Exercise 3.1**

- Write down the structural formula of butane

We have derived the first four straight chain alkanes which are:

- Methane
- Ethane
- Propane
- Butane

Now let us consider the following. Probably, you still remember in your Form 1 Mathematics, names given to figures with different number of sides e.g.

- Five
- Six
- Seven
- Eight
- Nine
- Ten

Compare the names of these figures with the prefixes of the alkanes listed below. This will help you to remember the names of the fifth alkane onwards easily. Make sure that you learn the names of the first four alkanes.

- Meth-for 1 carbon atom
- Eth-for 2 carbon atoms
- Prop-for 3 carbon atoms
- But-for 4 carbon atoms
- Pent-for 5 carbon atoms
- Hex-for 6 carbon atoms
- Hept-for 7 carbon atoms
- Oct-for 8 carbon atoms
- Non-for 9 carbon atoms
- Dec-for 10 carbon atoms
Naming alkanes

The names of alkanes are derived by first writing a prefix and then ending with “ane”. For example an alkane with:

• one carbon atom is *Methane*
• two carbon atoms is *Ethane*
• three carbon atoms is *Propane*
• four carbon atoms is *Butane*

Following the same procedure, write the names of straight chain alkanes with five, six, seven, eight, nine and ten carbon atoms.

You may have noted that the last three letters, *ane*, form the last part of the names listed. This indicates that the compounds belong to the class of hydrocarbons called alkanes.

Structural formula shows the bonds between the atoms that make the molecule as shown below.

(a) The molecular formula for methane is $\text{CH}_4$.
(b) The structural formula for methane is represented as:

\[
\begin{array}{cc}
\text{H} \\
\text{H} \quad \text{C} \quad \text{H} \\
\text{H} \\
\text{H}
\end{array}
\]

(a) The molecular formula for ethane is $\text{C}_2\text{H}_6$.
(b) The structural formula for ethane is represented as:

\[
\begin{array}{cc}
\text{H} & \text{H} \\
\text{H} \quad \text{C} \quad \text{C} \quad \text{H} \\
\text{H} & \text{H}
\end{array}
\]

Exercise 3.2

• Write the molecular formula and draw the structural formula of straight chain alkanes with five, six, seven, eight, nine and ten carbon atoms.
• Construct models of methane, ethane, propane and butane.
### Table 3.1: The first ten straight chain alkanes

<table>
<thead>
<tr>
<th>Number of carbon atoms</th>
<th>Name of alkane</th>
<th>Molecular formula</th>
<th>Structural formula</th>
<th>Condensed form of the structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Methane</td>
<td>CH₄</td>
<td>H [H[H[H]H[H]</td>
<td>CH₄</td>
</tr>
<tr>
<td>2</td>
<td>Ethane</td>
<td>C₂H₆</td>
<td>H H [H[H[H]H|CH₃CH₃</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>Propane</td>
<td>C₃H₈</td>
<td>H H H [H[H[H]</td>
<td>CH₃CH₂CH₃</td>
</tr>
<tr>
<td>4</td>
<td>Butane</td>
<td>C₄H₁₀</td>
<td>H H H H [H[H|CH₃CH₂CH₂CH₃</td>
<td></td>
</tr>
<tr>
<td>5</td>
<td>Pentane</td>
<td>C₅H₁₂</td>
<td>H H H H H [H|CH₃CH₂CH₂CH₂CH₃</td>
<td></td>
</tr>
<tr>
<td>6</td>
<td>Hexane</td>
<td>C₆H₁₄</td>
<td>H H H H H H [H|CH₃CH₂CH₂CH₂CH₂CH₃</td>
<td></td>
</tr>
<tr>
<td>7</td>
<td>Heptane</td>
<td>C₇H₁₆</td>
<td>H H H H H H H [H|CH₃CH₂CH₂CH₂CH₂CH₂CH₃</td>
<td></td>
</tr>
<tr>
<td>8</td>
<td>Octane</td>
<td>C₈H₁₈</td>
<td>H H H H H H H H [H|CH₃CH₂CH₂CH₂CH₂CH₂CH₂CH₃</td>
<td></td>
</tr>
<tr>
<td>9</td>
<td>Nonane</td>
<td>C₉H₂₀</td>
<td>H H H H H H H H H [H|CH₃CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₃</td>
<td></td>
</tr>
<tr>
<td>10</td>
<td>Decane</td>
<td>C₁₀H₂₂</td>
<td>H H H H H H H H H H H [H|CH₃CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₂CH₃</td>
<td></td>
</tr>
</tbody>
</table>
Homologous series

Study the molecular and structural formulae of alkanes shown in Table 3.1. How do they differ from each neighbouring member? The alkanes form a family called a homologous series.

A homologous series is a group of compounds in which all the members:
- have the same general formula e.g. alkanes have the formula \( C_nH_{2n+2} \)
- differ from each other by a \(-\text{CH}_2-\) group.
- they have similar chemical properties.
- can be prepared using a general method.
- exhibit a gradual change along the series in physical properties such as melting point, boiling point and solubility.

Fractional distillation of crude oil

We have seen that alkanes are a major component of crude oil. Crude oil is a thick black liquid with a strong smell. Unless its components are separated, crude oil does not have much use. During fractional distillation at a refinery, the mixture of hydrocarbons is sorted out into groups or individual hydrocarbons called fractions. A fractionating column is used for separating the mixture.

![Fractional distillation of crude oil diagram](image)

*Fig. 3.1 Fractional distillation of crude oil*
**Fractionating column**

The crude oil is first heated up in a furnace. As it is heated, the small molecules boil off first. They enter the column as a gas. The fractionating column is hot at the bottom and cooler at the top. The gas molecules then condense. The larger hydrocarbons have higher boiling points. This means that the larger hydrocarbons, with the high boiling points, turn back to liquids easily nearer the bottom. At high temperatures, the hydrocarbons are in form of gases. They rise up the column. The different fractions condense and are collected at different levels as shown in Fig 3.1.

**Isomerism**

Earlier in this unit, we discussed straight chain alkanes. However, there are many alkanes that are not straight chains. Some of the carbon atoms are branched off the continuous chain of carbon atoms in the molecule. These branched alkanes have the same molecular formula as straight chain alkanes but the way the carbon atoms are connected to one another differs. Compounds which have the same molecular formula but different in structural formulae are called isomers. Alkanes have one type of isomerism known as chain isomerism.

Let us consider an example butane, $\text{C}_4\text{H}_{10}$. The carbon and hydrogen atoms in butane can be arranged in a number of different ways. Let us start with the straight chain butane.

```
   H   H   H
   |   |   |
   H--C--C--C--H
   |   |   |
   H   H   H
```

Butane

Study the following molecular structures of butane in Table 3.2.

**Table 3.2: Molecular structures of butane**

<table>
<thead>
<tr>
<th>Molecular structure</th>
<th>Number of C atoms</th>
<th>Number of H atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a)</td>
<td>$\text{H \ H \ H}$</td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{H--C--C--C--H}$</td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{H \ H \ H}$</td>
<td></td>
</tr>
<tr>
<td>(b)</td>
<td>$\text{H \ H \ H}$</td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{H--C--C--C--H}$</td>
<td></td>
</tr>
<tr>
<td></td>
<td>$\text{H \ H \ H}$</td>
<td></td>
</tr>
</tbody>
</table>
• Count the number of carbon and hydrogen atoms for each structure in Table 3.2.

• Copy Table 3.2 in your note book and fill in the empty spaces.

• Use your index finger to follow the carbon chain without moving it backwards. Which molecular structure, a, b or c do you have to lift your finger to pass over all the four carbon atoms?

• Record the results in your note book.

The structural formulae (a) and (b) in Table 3.2 are said to be straight chain alkanes. However, the carbon atoms are not in a straight line although they are in a continuous chain.

The structural formula in (c) is the only one that is branched. The carbon atoms are not in a straight line and neither are they in a continuous chain.

The structures in (a), (b) and (c) have identical molecular formulae. However, they have different arrangements of carbon atoms and this gives them different structural formulae.

Such forms of a compound such as butane which are different in structure but have the same molecular formulae are called isomers.

The existence of two or more compounds with the same molecular formula but different structures is called isomerism.

• Now, let us see how we name the isomers.

**Systematic (IUPAC) naming of branched alkanes**

The following are rules used in naming branched alkanes

1. Determine the longest continuous chain of carbon atoms. This chain does not have to be a straight line. It can be zigzag. This will give us the parent or ending name.

2. Identify the branching(side) groups attached to this continuous chain of carbon atoms.

3. Number the carbon atoms in the longest chain so that the groups would be attached to the carbon assigned the lowest number possible. This means that you have to decide whether to number from the right or left end of the chain or from top to bottom and vice versa.

4. The names of the side groups and the numbers of the carbon atoms to which they are attached will form prefixes of the parent or ending name.
5. If more than one of the same kind of branched groups are attached to the chain, identify the numbers of the carbon atoms bearing each group and write this as a series of numbers separated by commas between each number, then a hyphen and finally use a Greek prefix attached to the branch name as discussed under rule No. 6 below.

Substituent groups
We have discussed branched groups. These groups are named according to the number of carbon atoms. A branch with one carbon atom (–CH₃) is known as a methyl group, while that with 2 carbon atoms is known an ethyl group (–CH₂CH₃). A branch with three carbon atoms is known as a propyl group (CH₃CH₂CH₂–) and so on.

In alkanes and many organic compounds, hydrogen atoms which are bonded to the continuous carbon chain are substituted by other groups known as substituent groups. The commonest substituent groups or branches are obtained from alkanes, whose molecules have one hydrogen atom less. The general name of such a substituent group is an alkyl group. The name is obtained from the name of the alkane by replacing -ane with -yl.

Naming the alkyl groups
Alkyl groups are named by adding suffix -yl to the prefix of the parent name as given in Table 3.3

Table 3.3: Naming alkyl groups

<table>
<thead>
<tr>
<th>Number of carbon atoms in the longest continuous carbon chain</th>
<th>Name of alkane</th>
<th>Prefix</th>
<th>Name of alkyl group</th>
<th>Formula of the alkyl group</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>methane</td>
<td>meth-</td>
<td>methyl</td>
<td>CH₃– CH₃CH₂–</td>
</tr>
<tr>
<td>2</td>
<td>ethane</td>
<td>eth-</td>
<td>ethyl</td>
<td>CH₃CH₂–</td>
</tr>
<tr>
<td>3</td>
<td>propane</td>
<td>prop-</td>
<td>propyl</td>
<td>CH₃CH₂CH₂–</td>
</tr>
</tbody>
</table>

In case there is more than one member of the same group for example, a methyl group attached to different carbon atoms in the chain, we use Greek prefixes e.g.

- di-for two, hence the alkyl group becomes dimethyl
- tri-for three, hence the alkyl group becomes trimethyl.

Let us now apply the rules listed above to name a branched alkane which has the following structure.
Rule No. 1 — for parent name
Identify the longest continuous chain of carbon atoms

- When you trace the carbon chain from all sides, the longest continuous carbon chain has three carbon atoms. Therefore the ending or parent name will be **propane**.

Rule No. 2 — for branches
Identify the branches (substituent groups).

There is **one methyl group** branched off the continuous carbon chain. The prefix will therefore be methyl, hence the name is methylpropane.

Rule No. 3 — for position of branches on the carbon chain
Number the carbons in the continuous chain so that the methyl group will be attached to the carbon atom assigned the lowest number on the continuous carbon chain. In this case, it is important to number carbon atoms from all sides of the chain to be able to identify the lowest possible numbering of the position(s) of the branch(es).

- From all sides the lowest number on the carbon atom within the chain that is bearing the methyl group is number two (2).
Rule No. 4 — for inserting a hyphen
Write carbon number bearing the methyl group followed by a hyphen then the name of the alkyl group e.g. 2–methyl.

Rule No. 5 — for ending name
Use the parent alkane name corresponding to the number of carbon atoms on the longest continuous chain and write this next to the alkyl group as shown below.

2–methylpropane

Note
Sometimes more than one hydrogen atoms in an alkane may be substituted by two identical groups. Consider the example of the structure shown below.

![Structure](image)

To name this isomer correctly we use rule No. 6.

Rule No. 6 — for identical branches on the same carbon
When the substituent groups on the same carbon atom are identical, we write the number on that carbon twice before the name of the compound.

If you use rules No. 1 to 5, the name of the compound is 2–dimethylpropane. This is not correct! We must apply rule No. 6. Hence the correct name is 2,2–dimethylpropane.

Exercise 3.3
- Write down the structural formulae of the isomers of pentane,
- Following rules No. 1 to 6 name the isomers.

3.4 Preparation and properties of methane and ethane

Experiment 3.1
Laboratory preparation of methane

Apparatus and chemicals
- hard glass test tube/ round bottomed flask
- source of heat
• delivery tube
• trough
• beehive shelf
• mortar and pestle
• gas jar
• sodium ethanoate
• soda lime (a mixture of sodium hydroxide and calcium oxide). It is easier to handle soda lime than the deliquescent sodium hydroxide. It does not dissolve easily.

Procedure
1. Place about 4-5g of sodium ethanoate and an equal amount of soda lime in a mortar and grind well with a pestle.
3. Transfer the mixture into a hard glass test tube.
4. Set up apparatus as in Fig. 3.2
5. Heat the test tube carefully. Make sure that the water is not sucked back by removing the delivery tube from water immediately after heating is stopped.

Fig.3.2: Preparation of methane

- What do you observe?
- What is the smell of the gas collected?
- What physical property of the gas is tested by the method of collection used?
- What can you conclude from your observations?
- Record your observations and conclusions in your notebook.

6. Collect three test tubes full of this gas for carrying out the tests shown in Table 3.4. Copy Table 3.4 and record your observations in the spaces provided.
Table 3.4: Observations on chemical reactions of methane

<table>
<thead>
<tr>
<th>Test-tube</th>
<th>Test</th>
<th>Observations</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Introduce a lighted splint at the mouth of the test tube containing the gas</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>Invert a test tube full of methane in a beaker containing methylbenzene</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>Add about 4-5 drops of bromine water</td>
<td></td>
</tr>
</tbody>
</table>

- What do you conclude about methane?
- Write equations for reactions that take place.

Methane is prepared in the laboratory by heating a mixture of sodium ethanoate and soda lime. The sodium hydroxide in the mixture reacts with sodium ethanoate to form methane gas and sodium carbonate. The equation for the reaction is:

$$\text{Sodium ethanoate} + \text{soda lime} \rightarrow \text{methane} + \text{sodium carbonate}$$

$$\text{CH}_3\text{COONa}(s) + \text{NaOH}(s) \rightarrow \text{CH}_4(g) + \text{Na}_2\text{CO}_3(s)$$

Generally any alkane can be prepared by similar reactions represented by the following equation:

$$\text{Sodium alkanoate} + \text{soda lime} \rightarrow \text{alkane} + \text{sodium carbonate}$$

$$\text{C}_n\text{H}_{2n+1}\text{COONa}(s) + \text{NaOH}(s) \rightarrow \text{C}_n\text{H}_{2n+2} + \text{Na}_2\text{CO}_3(s)$$

where $n$ is a whole number 1, 2, 3,...

Physical properties of methane
Methane:
- is a colourless gas.
- is an odourless gas.
- is less dense than air.
- is insoluble in water but soluble in organic solvents.

Preparation of ethane
Study carefully the procedure for preparation of methane.
• Predict the chemicals required for preparation of ethane.
• List the apparatus used in preparation of ethane.
• Write the procedure for the preparation of ethane.
• Write an equation for the preparation of ethane.
• What are the physical properties of ethane?
• What are the chemical properties of ethane?

Trends in physical properties of alkanes
Table 3.5 shows a summary of the physical properties of the first ten alkanes.

• What are the trends in physical properties of alkanes?

Table 3.5: Trends in physical properties of alkanes

<table>
<thead>
<tr>
<th>No. of carbon atoms</th>
<th>Name of alkane</th>
<th>Melting point (°C)</th>
<th>Boiling point (°C)</th>
<th>Density g/cm³</th>
<th>Physical state of alkanes at room temperature</th>
<th>Solubility in water</th>
<th>Solubility in organic solvent</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>methane</td>
<td>–182</td>
<td>–161</td>
<td>0.424</td>
<td>gas</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>2</td>
<td>ethane</td>
<td>–183</td>
<td>–88</td>
<td>0.546</td>
<td>gas</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>3</td>
<td>propane</td>
<td>–188</td>
<td>–42</td>
<td>0.501</td>
<td>gas</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>4</td>
<td>butane</td>
<td>–138</td>
<td>0</td>
<td>0.579</td>
<td>liquid</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>5</td>
<td>pentane</td>
<td>–130</td>
<td>36</td>
<td>0.626</td>
<td>liquid</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>6</td>
<td>hexane</td>
<td>–95</td>
<td>69</td>
<td>0.657</td>
<td>liquid</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>7</td>
<td>heptane</td>
<td>–90</td>
<td>99</td>
<td>0.684</td>
<td>liquid</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>8</td>
<td>octane</td>
<td>–57</td>
<td>126</td>
<td>0.703</td>
<td>liquid</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>9</td>
<td>nonane</td>
<td>–53</td>
<td>151</td>
<td>0.718</td>
<td>liquid</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>10</td>
<td>decane</td>
<td>–29</td>
<td>174</td>
<td>0.730</td>
<td>liquid</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
</tbody>
</table>

The trends in physical properties of alkanes is mainly determined by the carbon chain length. As the length of the carbon chain increases the molecular masses also increase hence the melting and boiling points increase. This is evident by the fact that the first four alkanes are gases and the next six are liquids. The increase in melting and boiling points of alkanes with increasing carbon atoms is also caused by the increase in intermolecular forces of attraction. The density also increases with increasing number of carbon atoms. This is due to increase in molecular mass. Alkanes are insoluble in water but highly soluble in non-polar solvents such as methylbenzene and hexane among other organic solvents.

Chemical properties of methane
The strong Carbon—Carbon and Carbon—Hydrogen covalent bonds make the alkanes relatively unreactive.
1. **Combustion of methane**

Methane burns in air or sufficient oxygen to produce carbon(IV) oxide gas and water.

\[
\text{Methane} + \text{oxygen} \rightarrow \text{carbon(IV) oxide} + \text{water}
\]

\[
\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l)
\]

This reaction is exothermic hence methane is a good fuel. If there is insufficient oxygen for complete burning, methane produces carbon(II) oxide and soot. This soot is carbon.

\[
\text{Methane} + \text{oxygen} \rightarrow \text{carbon(II) oxide} + \text{water}
\]

\[
2\text{CH}_4(g) + 3\text{O}_2(g) \rightarrow 2\text{CO}(g) + 4\text{H}_2\text{O}(l)
\]

\[
\text{Methane} + \text{oxygen} \rightarrow \text{carbon} + \text{water}
\]

\[
\text{CH}_4(g) + \text{O}_2(g) \rightarrow \text{C}(s) + 2\text{H}_2\text{O}(l)
\]

2. **Substitution reactions of alkanes**

Alkanes, react with halogens in the presence of sunlight. They do not react with halogens in the dark, because there is no enough energy which is necessary to start the reaction. Energy is needed to break the halogen bond in the reactants first.

**Substitution reactions of methane with chlorine and bromine**

Methane and chlorine react in sunlight or ultra-violet (u.v) light explosively. They form chloromethane and hydrogen chloride gas. The reaction is exothermic. Sunlight supplies the energy necessary to start the reaction by splitting the bond between the chlorine atoms (Cl – Cl). The free atoms formed attack methane and products are formed.

\[
\text{Methane} + \text{chlorine} \xrightarrow{\text{u.v}} \text{chloromethane} + \text{hydrogen chloride}
\]

\[
\text{CH}_4(g) + \text{Cl}_2(g) \xrightarrow{\text{u.v}} \text{CH}_3\text{Cl}(g) + \text{HCl}(g)
\]

In this reaction, one of the chlorine atoms reacts with methane and replaces (substitutes) a hydrogen atom in the methane molecule. This kind of reaction is called a substitution reaction. Hence we say alkanes undergo substitution reactions.

In the presence of excess chlorine, further substitution of hydrogen can occur. Finally all the hydrogen atoms are substituted by chlorine atoms forming tetrachloromethane.
Example 1

(a) Chloromethane + chlorine $\xrightarrow{\text{u.v}}$ dichloromethane + hydrogen chloride

$$\text{CH}_3\text{Cl}(g) + \text{Cl}_2(g) \xrightarrow{\text{u.v}} \text{CH}_2\text{Cl}_2(g) + \text{HCl}(g)$$

or

$$\text{CH}_3\text{Cl} + \text{Cl} - \text{Cl} \xrightarrow{\text{u.v}} \text{CH}_2\text{Cl}_2 + \text{HCl}$$

(b) Dichloromethane + chlorine $\xrightarrow{\text{u.v}}$ trichloromethane + hydrogen chloride

$$\text{CH}_2\text{Cl}_2(g) + \text{Cl}_2(g) \xrightarrow{\text{u.v}} \text{CHCl}_3(l) + \text{HCl}(g)$$

or

$$\text{CH}_2\text{Cl}_2 + \text{Cl} - \text{Cl} \xrightarrow{\text{u.v}} \text{CHCl}_3 + \text{HCl}$$

(c) Trichloromethane + chlorine $\xrightarrow{\text{u.v}}$ tetrachloromethane + hydrogen Chloride

$$\text{CHCl}_3(l) + \text{Cl}_2(g) \xrightarrow{\text{u.v}} \text{CCl}_4(l) + \text{HCl}(g)$$

or

$$\text{CHCl}_3 + \text{Cl} - \text{Cl} \xrightarrow{\text{u.v}} \text{CCl}_4 + \text{HCl}$$

Bromine substitutes the hydrogen atoms in methane in a similar way as chlorine in presence of sunlight. The reactivity of alkanes with halogens increase down the halogens’ group.

Example 2

(a) Methane + bromine $\xrightarrow{\text{u.v}}$ bromomethane + hydrogen bromide

$$\text{CH}_4(g) + \text{Br}_2(g) \xrightarrow{\text{u.v}} \text{CH}_3\text{Br}(g) + \text{HBr}(g)$$

or

$$\text{CH}_4 + \text{Br} - \text{Br} \xrightarrow{\text{u.v}} \text{CH}_3\text{Br} + \text{HBr}$$

Further substitution reactions can occur. Follow the procedure in example 1 and write equations for the remaining reactions.
Preparation of ethane

Ethane is prepared in the laboratory by heating a mixture of sodium propanoate and soda lime as shown in Fig 3.3.

What is soda lime and why do we use it to prepare ethane gas?

\[
\begin{align*}
\text{Sodium propanoate} & \quad + \quad \text{sodium hydroxide} \quad \longrightarrow \quad \text{ethane} \quad + \quad \text{sodium carbonate} \\
\text{CH}_3\text{CH}_2\text{COONa(s)} & \quad + \quad \text{NaOH(s)} \quad \longrightarrow \quad \text{C}_2\text{H}_6(g) \quad + \quad \text{Na}_2\text{CO}_3(s)
\end{align*}
\]

Physical properties of ethane

- Ethane is a colourless gas.
- It has no smell.
- It is insoluble in water but soluble in organic solvents such as benzene, ether, carbon tetrachloride, etc.

Chemical properties of ethane

1. Combustion
   Ethane burns in plentiful supply of oxygen with a clean luminous flame to form carbon(IV) oxide and water. The reaction is exothermic.
   \[
   \text{Ethane} \quad + \quad \text{oxygen} \quad \longrightarrow \quad \text{carbon(IV)oxide} \quad + \quad \text{water} \\
   2\text{C}_2\text{H}_6(g) \quad + \quad 7\text{O}_2(g) \quad \longrightarrow \quad 4\text{CO}_2(g) \quad + \quad 6\text{H}_2\text{O}(g)
   \]

2. Substitution reaction with chlorine
   Ethane reacts with chlorine in the presence of sunlight or ultraviolet radiation. Energy from sunlight provides energy for the chlorine molecules to split into
separate free atoms which are very reactive. One of the chlorine atoms react with ethane to substitute the hydrogen atom. The other chlorine atom combines with the substituted hydrogen atom to form hydrogen chloride gas.

\[
\text{Ethane} + \text{chlorine} \xrightarrow{\text{u.v light}} \text{chloroethane} + \text{hydrogen chloride}
\]

\[
\text{CH}_3\text{CH}_3(g) + \text{Cl}_2(g) \xrightarrow{\text{u.v light}} \text{CH}_3\text{CH}_2\text{Cl}(g) + \text{HCl}(g)
\]

Further substitution can occur. The second substitution can occur in two different positions; that is, on the carbon with a substituted hydrogen atom or on the carbon bearing three hydrogen atoms.

(a) Chloroethane + chlorine \(\xrightarrow{\text{u.v light}}\) 1,1-dichloroethane + hydrogen chloride

(i) \[
\text{CH}_3\text{CH}_2\text{Cl}(g) + \text{Cl}_2(g) \xrightarrow{\text{u.v light}} \text{CH}_3\text{CHCl}_2(l) + \text{HCl}(g)
\]

(ii) \[
\text{Cl} - \text{Cl} \xrightarrow{\text{u.v light}} \text{HCl}(g)
\]

(b) Chloroethane + chlorine \(\xrightarrow{\text{u.v light}}\) 1,2-dichloroethane + hydrogen chloride

(i) \[
\text{CH}_3\text{CH}_2\text{Cl}(g) + \text{Cl}_2(g) \xrightarrow{\text{u.v light}} \text{CH}_2\text{ClCH}_2\text{Cl}(g) + \text{HCl}(g)
\]

(ii) \[
\text{Cl} - \text{Cl} \xrightarrow{\text{u.v light}} \text{HCl}
\]

**Exercise 3.4**

1. Write down equations for further substitutions of 1,1-dichloroethane and 1,2-dichloroethane with chlorine and name the products.
2. When reacted with ethane, bromine substitutes in the same way as chlorine but
the reaction is slower. Write both word and structural formulae equations for the substitution reaction of ethane with bromine. Explain why sunlight is required for this reaction.

3.8: Uses of alkanes

1. Methane \((\text{CH}_4)\) which has been compressed in cylinders is used as fuel in vehicles in some countries as a substitute for petrol. Fig 3.4 (a).

2. Propane \((\text{C}_3\text{H}_8)\) and butane \((\text{C}_4\text{H}_{10})\) which are liquid components of liquified petroleum gas (LPG) are used in heaters, cars, school laboratories and machinery. Fig 3.4 (a), (c) and (d)

3. Butane \((\text{C}_4\text{H}_{10})\) mixed with small proportions of propane \((\text{C}_3\text{H}_8)\) is used in household lighting, water heating and cooking, Fig 3.4(e) (f).

4. Petrol is used in automobiles all over the world. Petrol is mainly heptane \((\text{C}_7\text{H}_{16})\) but other compounds like tetraethyllead(IV), \(\text{Pb}(\text{CH}_3\text{CH}_2)_4\) are added to create more smoothly burning fuel. Another additive is 1,2-dibromoethane.
(CH₂BrCH₂Br). This prevents lead from accumulating in the engine. Unfortunately this results in formation of lead bromide which is emitted in car exhaust system. It adds lead to the atmosphere, where it can be inhaled. Lead is a poison. It can cause damage to the brain and nervous system especially in young children. This is the reason why it is recommended to use unleaded petrol. See Fig. 3.4(b).

5. Higher alkanes are used as solvents in the manufacture of other chemicals.

6. Solid alkanes find other uses. Petroleum jelly for example vaseline, is a mixture of paraffin wax (a solid alkane) and oil.

3.3: Alkenes

General formula, nomenclature and structural formulae of alkenes

Alkenes are also hydrocarbons. They form a homologous series. The general formula of alkenes is \( \text{C}_n\text{H}_{2n} \), where \( n = 2,3,4... \) Alkenes differ from alkanes in that they have at least one double bond between two adjacent carbon atoms. The first alkene is ethene with value of \( n \) being 2.

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\mid & \quad \mid \\
\text{C} & \quad \text{C} \\
\mid & \quad \mid \\
\text{H} & \quad \text{H}
\end{align*}
\]

Ethene

OR

\[
\begin{align*}
\text{H} & \quad \text{C} \quad \text{C} \quad \text{H} \\
\text{H} & \quad \text{H}
\end{align*}
\]

NB. There is no alkene where \( n = 1 \). Why do you think this is so?

Nomenclature

This term means ‘systematic naming’. The names of all alkenes end with \textit{ene}. We have already seen how alkanes are named. An easy method of naming alkenes may be done by replacing the ‘a’ in alkanes with an “e”as follows:

- ethane gives us \textit{ethene}
- propane gives us \textit{propene}
- butane gives us \textit{butene}
- pentane gives us \textit{what}?
- hexane gives us \textit{what}?

Table 3.6 shows the first six members of alkenes. Study the information given in this table carefully.
Table 3.6: The first six alkenes

<table>
<thead>
<tr>
<th>Number of carbon atoms (n)</th>
<th>Name</th>
<th>Formula CnH2n</th>
<th>Structural formula</th>
<th>Condensed formula of the structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>ethene</td>
<td>C2H4</td>
<td>H H</td>
<td>CH2=CH2</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>C=C</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>H H</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>propene</td>
<td>C2H4</td>
<td>H H H</td>
<td>CH2=CHCH3</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>C=C – C – H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>H H</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>but-1-ene</td>
<td>C4H8</td>
<td>H H H H</td>
<td>CH2=CH2CH3</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>C=C – C – C – H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>H H H</td>
<td></td>
</tr>
<tr>
<td>5</td>
<td>pent-1-ene</td>
<td>C5H10</td>
<td>H H H H H</td>
<td>CH2=CHCH2CH3</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>C=C – C – C – C – H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>H H H H H</td>
<td></td>
</tr>
<tr>
<td>6</td>
<td>hex-1-ene</td>
<td>C6H12</td>
<td>H H H H H H</td>
<td>CH2=CH2CH2CH2CH3</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>C=C – C – C – C – C – H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>H H H H H H</td>
<td></td>
</tr>
</tbody>
</table>

Look at Table 3.6. Check on the structural formulae of the alkenes. What difference do you notice from one formula to the next? They differ by a –CH2– group, hence form a homologous series.

**IUPAC naming of alkenes**

The IUPAC rules used in naming alkanes should be used to derive systematic names for straight chain alkenes. The rules demand that the position of the double bond must be included in the name. This is why number 1 is included in some alkenes in Table 3.6. It tells us where the double bond is placed in that particular alkene as shown in the examples below.

![Structural formula of but-1-ene and but-2-ene](image)

The number shows the position of the double bond, and we write the lower number of the two carbon atoms bearing the double bond. In case of an alkene like ethene or
propene the number is not necessary because if you number the carbon chain from left or right hand side the double bond will be on the 1st carbon.

Note: In alkenes, you must count the number of bonds around each carbon atom properly to ensure each carbon atom is surrounded by maximum of 4 bonds.

Now let us discuss the other IUPAC rules of naming alkenes and see how we apply them.

**IUPAC naming of alkenes**

**Rules for naming alkenes**

1. Determine the longest continuous chain of carbon atoms that has a double bond between two carbon-atoms. This will give us the parent or ending name.
2. Number the carbon atoms in the chain so that the double bond is between the two carbon atoms assigned the lowest numbers. This means that you have to decide whether to number beginning on the right or left end of the chain. Always start from the side that gives the double bond the lowest number.
3. Identify the various branching groups or substituent groups attached to this continuous chain of carbon atoms.
4. If more than one member of a substituents groups is attached to the same carbon atom, the number of times we write the position number should correspond to the number of times the member appears separated by a comma then a hyphen then a prefix denoting the number of times the group is attached e.g di, tri etc. followed by the group name and then the name of the parent alkene e.g 3,3-dimethylbutene etc.
5. Write the lowest number of the carbon atoms bearing the double bond.

Now let us follow the above rules in the following example and give the systematic name of the compound whose structure is shown below.

1. Identify the longest continuous chain of carbon atoms with the double bond.
The longest carbon chain has **four** carbon atoms. Therefore the parent name is *butene*.

2. **Number the carbon chain so that the double bond is between the lowest numbered carbon atoms.**

3. **Identify and locate all branched groups, attaching a prefix number to which the branch is attached.**

4. **Write the number of the carbon atoms bearing alkyl group first followed by a hyphen then the name of the alkyl group. Then, end with the parent name changing “ane” to “ene”.**

   **Note:** “a” changes to “e”. We must also include the lowest number in the alkene name to indicate where the double bond is placed. In this case, we insert 1 in between the parent name as follows:

   *2 –methylbut –1– ene*
Therefore \( \text{2} \) tells us the position of the substituent group, i.e. the methyl group is on the second carbon. \( \text{1} \) tells us that the double bond starts on the 1st carbon, so it must be between the first and second carbon.

**Exercise 3.5**

1. Name the following alkenes:
   (a) \( \text{CH}_3\text{CH}_2\text{CH}≡\text{CH}_2 \)
   (b) \( \text{CH}_2≡\text{C}—\text{CH}_3 \)

2. Write the structural formulae for the following alkenes:
   (a) hex – 2 – ene
   (b) 2,3 – dimethylbut – 2 –ene

**Isomerism**

Earlier we learnt that alkanes exhibit only chain isomerism. Unlike alkanes, alkenes show two types of structural isomerism.

These are:
- chain isomerism.
- position isomerism.

Let us consider the structure of pentene to study the two types of isomerism.

\[
\begin{align*}
\text{H} & \quad \text{H} & \quad \text{H} & \quad \text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} & \equiv & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{H} \\
\text{H} & \quad \text{H} & \quad \text{H} & \quad \text{H} & \\
\text{Pent-1-ene} & 
\end{align*}
\]

**1. Chain isomerism**

It is possible to draw other different structural formulae of pent-1-ene. The atoms of carbon are arranged in a different way in each structure as shown below.

\[
\begin{align*}
\text{H} \\
\text{H} & \quad \text{C} & \equiv & \quad \text{C} & \quad \text{H} \\
\text{H} & \quad \text{C} & \equiv & \quad \text{C} & \quad \text{C} & \quad \text{C} & \quad \text{H} \\
\text{H} & \quad \text{H} & \quad \text{H} & \quad \text{H} & \\
\text{3 – methylbut – 1 –ene} & 
\end{align*}
\]
The above compounds (a) and (b) have the same molecular formula, but they have different structural formulae. The general name of such process as above is called isomerism. In 3–methylbut–1–ene and 2–methylbut–1–ene, the position of the double bond remains the same but the methyl group has been moved from position 3 to position 2. Such isomers are known as chain structural isomers.

2. Position isomerism
Sometimes the position of the double bond in the alkene can be shifted to a different position within the molecule. This type of isomerism is known as position isomerism. This is shown in the following examples:-

Example 1

\[ \text{Pent – 1 –ene} \]

\[ \text{Pent – 2 –ene} \]

Example 2

\[ \text{3 – methylbut – 1 –ene} \]
Note that in the above isomers, the position of the **double bond** shifts while the position of the carbon atoms remains the same.

**Exercise 3.6**
1. What are isomers?
2. Is there any difference between structural isomers and position isomers?
3. Draw and name all structural and position isomers of hexene.
4. Name the following compounds.
   (a) 
   (b) \( \text{CH}_3\text{CH} = \text{CHCH}_3 \)
   (c) \( \text{CH}_3(\text{CH}_2)_2\text{CH} = \text{CH}_2 \)
5. What is the main feature in the bonding of an alkene? Why is the first member of this series not methene?
6. Draw the structural formulae of the following compounds.
   (a) But – 2 –ene
   (b) 2 – methylprop-1-ene
   (c) 3,3 – dimethylbut –1 –ene
7. What is the difference between saturated and unsaturated hydrocarbons.

**Sources of alkenes**
(a) Dehydration of alcohols
(b) Cracking hydrocarbons
(a) Preparation and properties of ethene

Experiment 3.2
Laboratory preparation of ethene by dehydration of ethanol

Apparatus and chemicals
- thermometer
- round-bottomed flask
- sand bath
- boiling tube
- cork rubber stopper
- bunsen burner
- delivery tubes
- trough
- beehive shelf
- gas jars
- concentrated sulphuric acid
- ethanol
- bromine water
- acidified potassium manganate(VII)
- broken porcelain / dry sand.

Caution: Concentrated sulphuric acid is corrosive. Ethanol and ethene are flammable. There is substantial production of sulphur(IV) oxide in this reaction which can cause irritation and even trigger an asthmatic attack. This experiment should therefore be done in a fume chamber or in open space.

Procedure
1. Place about 20 cm³ of ethanol in a round-bottomed flask and slowly add about 40 cm³ concentrated sulphuric acid, while cooling and shaking the flask.

![Fig. 3.5: Laboratory preparation of ethene gas](image-url)
2. To the round-bottomed flask, add 2-3 g of clean dry broken pieces of porcelain or sand and assemble the apparatus as shown in Fig 3.5.

3. Heat the flask gently over a sand bath. What do you observe?

4. Allow the gas produced to escape from the delivery tube for a few minutes then collect samples of this gas in a boiling tubes. Close the tube with a cork, then remove the cork and light the gas. If the tube contains ethene, it burns with a blue flame; if ethene is mixed with air, the mixture in the test tube ignites with an explosion. Allow more gas to escape from the delivery tube until you prove that it is pure ethene. Collect several gas jars of ethene.
   • The first portion of the gas evolved is simply displaced hot air from apparatus.
   • Ethene is flammable and it is important not to ignite the gas directly from the delivery tube.

5. Disconnect the delivery tube connecting the round-bottomed flask to the trough before you stop heating.
   • Why must the tube be disconnected before you stop heating?

6. Bubble the gas through the test tubes containing:
   (a) bromine water
   (b) acidified potassium manganate(VII). Record your observations.
   • What do you conclude?

7. The residue in the flask should be disposed off by first allowing it to cool, and then pouring it into a large volume of cold water.
   • Why is sulphuric acid added into the flask while cooling and shaking the glass?
   • Why should clean, dry broken porcelain or sand be added into the mixture of conc. sulphuric acid and ethanol?
   • Why is it advisable to heat the acid gently?
   • Why does the gas ignite with an explosion if it is not pure ethene?
   • Why must the delivery tube be disconnected from the round-bottomed flask before the heat is turned off?
   • Why should the residue in the flask be first poured into a large volume of cold water when disposing it?

When a mixture of ethanol and concentrated sulphuric acid is heated gently an oily liquid called ethyl hydrogensulphate and water are formed.

\[
\text{Ethanol} + \text{conc. Sulphuric acid} \rightarrow \text{ethyl hydrogensulphate} + \text{water} \\
C_2H_5OH(l) + H_2SO_4(aq) \rightarrow C_2H_5HSO_4(l) + H_2O(l)
\]

When the mixture is heated to 170°C, a steady stream of ethene is produced. The ethyl hydrogensulphate formed is unstable. It decomposes to sulphuric acid and ethene gas.
Ethanol is produced by the dehydration of ethanol by concentrated sulphuric acid and can be shown below.

\[ \text{Ethanol} \quad \xrightarrow{\text{dehydration by conc. } \text{H}_2\text{SO}_4} \quad \text{ethyl hydrogensulphate} \]

\[ \text{C}_2\text{H}_5\text{OH} \quad \xrightarrow{\text{conc. } \text{H}_2\text{SO}_4} \quad \text{C}_2\text{H}_5\text{HSO}_4 \]

The above chemical reactions are equivalent to the dehydration of ethanol by concentrated sulphuric acid and can be as shown below.

\[ \text{CH}_3\text{CH}_2\text{OH} \quad \xrightarrow{\text{conc. } \text{H}_2\text{SO}_4 - \text{H}_2\text{O}} \quad \text{CH}_2 = \text{CH}_2 \]

Ethanol contains elements that make up water i.e. hydrogen and oxygen hence water is removed in the above reaction. When concentrated sulphuric acid reacts with this water, a lot of heat is produced i.e. the reaction is extremely exothermic. It is therefore necessary to cool the flask while this reaction is being carried out. Ethanol and concentrated sulphuric acid are miscible although they have different densities. Shaking the flask helps to mix the two liquids properly. Dry broken porcelain increase the surface area on which gas bubbles can form. They also ensure smooth boiling of the mixture.

The mixture of ethanol and the acid is heated gently to prevent a violent reaction which can cause large amounts of sulphuric acid to jump off the flask into the delivery tube. Pure ethene burns with a blue flame, otherwise if mixed with air, it burns with an explosion. That is why the first portion of the gas collected is allowed to escape.

The delivery tube should first be disconnected from the round bottomed flask before we stop heating to prevent water from being sucked back into the flask. The sulphuric acid remaining in the flask as a residue should be diluted by pouring it into a large volume of cold water. This is to dilute the acid so as to make it less corrosive.

**Note:** By dehydrating different alcohols, different alkenes are produced. The general equation for the production of alkenes through dehydration of alcohols is:

\[ \text{Alcohol} \quad \xrightarrow{\text{Conc. sulphuric acid} - \text{water}} \quad \text{Alkene} \]

\[ \text{C}_n\text{H}_{2n+1}\text{OH} \quad \xrightarrow{\text{conc. } \text{H}_2\text{SO}_4 - \text{H}_2\text{O}} \quad \text{C}_n\text{H}_{2n} \]

(b) **Cracking hydrocarbons in industry**

Alkenes can be obtained from alkanes by heating oil after refining. This process is called **cracking**. Cracking can be done in two ways i.e. thermal cracking (heating) or catalytic cracking (heating with a catalyst). Cracking involves breaking long-chain alkane molecules into a shorter alkane and an alkene. For example, when hexane is cracked butane and ethene are obtained.
After distilling the crude oil, too many large hydrocarbons are left behind. The smaller hydrocarbons like petrol, are in great demand. Scientists have found a way to change the larger less useful molecules into smaller, more useful ones.

Cracking helps produce more petrol. The petrol obtained this way is of a better quality than obtained by distillation of crude oil. Cracked petrol is used to blend other petrol to improve quality. When a long carbon chain alkane is cracked (split) at high temperature it produces a shorter alkane and an alkene.

\[
\text{long alkane } \xrightarrow{700^\circ C} \text{ shorter alkane + alkene}
\]

Example
When a long molecule e.g. decane, C_{10}H_{22}, is cracked it produces octane and ethene.

\[
\text{C}_{10}\text{H}_{22} \xrightarrow{} \text{C}_8\text{H}_{18} + \text{C}_2\text{H}_4
\]

When a catalyst is used, cracking can be made to occur at fairly low temperatures. This is known as **catalytic cracking**.

**Note:** Cracking produces extra petrol and is a source of alkenes.

![Cracking large hydrocarbons](image)

**Trends in physical and chemical properties of alkenes**
Table 3.7 shows a summary of physical properties of the first five alkenes. What can you observe about the trends in the physical properties of alkenes?
Alkenes have similar chemical properties and show a steady gradation of physical properties.

- There is increased melting and boiling points as the number of carbon atoms increase. This is because with increasing carbon atoms the molecular mass increases. This also causes an increase in the intermolecular forces of attraction. To break these intermolecular forces, more energy is required hence the increase in melting and boiling points.
- The first three alkenes, i.e., ethene, propene and but-1-ene are gases at room temperature while pent-1-ene and hex-1-ene are liquids. This also has to do with increase in intermolecular forces of attractions.
- Alkenes are organic compounds hence they are insoluble in water but soluble in organic solvents.

### Physical properties of ethene

**Ethene:** Is a colourless gas, has no smell, and is insoluble in water.

**Chemical properties of ethene**

A double bond between carbon atoms is a stronger link than a single bond. However, the second bond is weaker than a single bond i.e., in terms of energy needed to break bonds.

As such, the second bond is easier to break. This bond easily opens up and takes additional atoms for example, those of chlorine and bromine, making the unsaturated carbon atoms become saturated.

Let us now consider the individual addition reactions of ethene.

#### 1. Combustion of ethene

Ethene is highly flammable and burns readily in air to form carbon(IV) oxide and water.

\[
\text{Ethene} + \text{oxygen} \rightarrow \text{carbon(IV) oxide} + \text{water}
\]

\[
\text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O(l)}
\]
2. Addition reactions

(i) Addition of hydrogen to ethene
Ethene reacts with hydrogen in the presence of a nickel catalyst to form ethane. A temperature of 150°C is required.

\[
\text{Ethene} + \text{hydrogen} \xrightarrow{\text{nickel catalyst}} \text{ethane}
\]

\[
C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)
\]

This process is also called **hydrogenation**.

(ii) Addition of hydrogen halides (HX) to ethene
The general formula of hydrogen halides is HX. X represents any atom of a halogen such as chlorine, bromine or iodine. The reaction of ethene with a hydrogen halide is as follows:

\[
\text{Ethene} + \text{HX} \rightarrow \text{alkyl halides}
\]

Hydrogen halides readily add across the double bond of alkenes to form alkyl halides. The reactivity of ethene with the hydrogen halides is as follows: HI > HBr > HCl. Ethene reacts readily with hydrogen iodide and with hydrogen bromide at room temperature to form iodoethane and bromoethane respectively.

\[
\text{Ethene} + \text{hydrogen iodide} \rightarrow \text{iodoethane}
\]

\[
C_2H_4(g) + HI(aq) \rightarrow C_2H_5I(g)
\]

\[
\text{Ethene} + \text{hydrogen bromide} \rightarrow \text{bromoethane}
\]

\[
C_2H_4(g) + HBr(aq) \rightarrow C_2H_5Br(g)
\]
Note: With ethene, the hydrogen atom of the hydrogen iodide can add to either carbon atom to yield iodoethane. Follow the above procedure to write the equations and structures of the reaction between ethene with hydrogen chloride.

(ii) Addition of halogens to alkenes
Halogens readily add across a double bond of alkenes to form compounds called **dihalides**. This process is called **halogenation**. We may represent a halogen molecule with \( X_2 \) or \( X - X \), for example \( Cl_2 \) or \( Br_2 \).

A general reaction of ethene with a halogen is shown below.

\[
\ce{H2 + C==C + H-Br -> H-\ce{C=C-Br}}
\]

**Note:** When naming the product formed, in which a hydrogen atom is substituted with a halogen atom, the prefix is borrowed from the name of the halogen where \( ine \) part is replaced with "o" as follows:

- from chlorine we get **chloro**
- bromine we get **bromo**
- iodine we get **iodo**.

Chlorine and bromine react with ethene to form 1,2-dichloroethane and 1,2-dibromoethane respectively as shown below. Both the products are colourless. Ethene instantly decolourises both bromine and chlorine. The decolourization of bromine is a test for unsaturated hydrocarbons.

\[
\ce{C2H4(g) + Cl2(g) -> C2H4Cl2(g)} \quad \text{(green-yellow)}
\]

\[
\ce{C2H4(g) + Br2(l) -> C2H4Br2(l)} \quad \text{(brown)}
\]
Then, the monomers add up in a self-reaction. Many monomers join to form a long chain consisting of thousands of carbon atoms. The process by which many small molecules called monomers join to form a large molecule is called polymerisation. The large molecule formed is called a polymer. The process can be represented by the general equation shown below.

\[ n \left( \begin{array}{c} H \\ H \\ C = C \\ H \end{array} \right) \quad \Rightarrow \quad \left( \begin{array}{c} H \\ H \\ C - C \\ H \end{array} \right)_n \]

A monomer in this case is an individual unit of ethene.

Follow the above procedure and write the equation and structures of the reactants and products formed in the reaction ethene with iodine.

(c) Polymerisation reactions (self addition reactions)

When ethene is heated under high pressure and high temperatures, 200°C, in the process low density polymer (polythene) forms. First, one of the double bond in each molecule breaks open, forming a monomer with free electrons which form bonds with next neighbours.

Many monomers join to form a long chain consisting of thousands of carbon atoms. The process by which many small molecules called monomers join to form a large molecule is called polymerisation. The large molecule formed is called a polymer. The process can be represented by the general equation shown below.

It can also be represented as:

\[ n(CH_2 = CH_2 + CH_2 = CH_2 + CH_2 = CH_2) \rightarrow \left[ CH_2 - CH_2 - CH_2 - CH_2 - CH_2 \right]_n \]

As shown above, one bond of each ethene molecule undergoes addition reaction. The process is therefore called addition polymerisation.
Test for unsaturation
Ethene as we have already mentioned, is the first alkene. Ethene has a double bond between its two carbon atoms. Hence it is said to be unsaturated. Now let us see how we can show the presence of a hydrocarbon with a double bond.

Experiment 3.3
Test for unsaturation

Apparatus and chemicals
• test-tubes
• teat pipette
• acidified potassium manganate(VII) solution.
• bromine water

Procedure
1. Bubble ethene through acidified potassium manganate(VII) solution.
   • What do you observe?
   • Record your observation in your notebook.
2. Bubble ethene through bromine water.
   • What do you observe?
   • What can you conclude from the above observations?
3. Record your observations and conclusions in your notebook.
4. What would happen if we bubbled ethane through bromine water?

Ethene decolourises potassium manganate(VII) solution as well as bromine water. The decolouration of these solutions is the test for a multiple bond i.e double or triple bond, hence the test for unsaturation.

(vi) Reaction of ethene with acidified potassium manganate(VII) solution
When alkenes are reacted with acidified potassium manganate(VII) solution, they decolourise the solution. The alkenes are oxidised to compounds with two –OH groups called diols.

\[
\begin{align*}
\text{Ethene} & \quad \text{MnO}_4^-/H^+ & \quad \text{Ethane–1,2 – diol} \\
H & \quad \text{C} & \quad \text{H} \\
\text{H} & \quad \text{C} & \quad \text{H} \\
\text{H} & \quad \text{H} & \quad \text{OH} & \quad \text{OH}
\end{align*}
\]

Acidified potassium manganate (VII) is represented by \(\text{MnO}_4^-/H^+\)
(vii) **Reaction of ethene with bromine water**

When ethene reacts with bromine water, it decolourises bromine water and 2-bromoethanol is formed. Bromine water contains these acids, HBr and HOBr.

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\big| & \quad \big| \\
\text{C} & \quad \text{C} \quad \xrightarrow{\text{HOBr}} \quad \text{H} - \text{C} - \text{C} - \text{H} \\
\big| & \quad \big| \\
\text{H} & \quad \text{H} \\
\end{align*}
\]

2 - bromoethanol

- Write an equation for the reaction of hydrobromic acid with ethene.

**Summary of reactions of ethene**

**Uses of alkenes**

1. Ethene and propene are used in the manufacture of plastics. These plastics are made through addition polymerisation. Polythene is a plastic used to make dustbins, bags, electrical insulators, clothing etc. Fig 3.7 (a), (b), (c), (e), (f) and (g).

2. Ethene is used in artificial ripening of fruits such as mangoes, bananas etc. Fig. 3.7 (d).

Some uses of alkenes are illustrated in Fig. 3.7.
3.4: Alkynes

General formula, nomenclature and structural formulae of alkynes

Alkynes are the third homologous series of hydrocarbons. The general formula is $C_nH_{2n-2}$. Like alkenes, alkynes undergo addition reactions and also form polymers.

Fig. 3.7: Uses of alkenes

Exercise 3.7

1. An alkene has a melting point of $-140^\circ C$ and a boiling point of $64^\circ C$. Would you expect this alkene to be a solid, liquid or gas at room temperature ($25^\circ C$)? Explain.

2. Draw and name the structural formulae of all the isomers of pentene.

3. Write an equation to show the complete combustion of ethene in oxygen.

4. Draw a structural formula to show how ethene reacts with chlorine to produce 1,2-dichloroethane.

5. Write an equation to show how ethene reacts with hydrogen to produce ethane.
The first alkyne is ethyne, $C_2H_2$. The structural formula of ethyne is as follows:

$$\text{H} \quad \text{C} \quad \text{≡} \quad \text{C} \quad \text{―} \quad \text{H}$$

**ethyne**

**Note:** Alkynes have a triple bond between two carbon atoms.

**Nomenclature**
The names of all alkynes end with-**yne**. The names are derived from respective alkanes where “a” is replaced by “y” as follows:

- From ethane we get **ethyne**.
- From propane we get **propyne**.
- From butane we get **butyne**.
- From hexane we get **hexyne**, and so on.

Table 3.8 shows the first five members of alkynes. Study the information given.

**Table 3.8: The first five members of alkynes**

<table>
<thead>
<tr>
<th>Number of carbon atoms</th>
<th>Name</th>
<th>Molecular formula</th>
<th>Structural formula</th>
<th>Condensed formula of the structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>ethyne</td>
<td>$C_2H_2$</td>
<td>H–C≡C–H</td>
<td>HC ≡ CH</td>
</tr>
<tr>
<td>3</td>
<td>propyne</td>
<td>$C_3H_4$</td>
<td>H–C≡C–C–H</td>
<td>HC ≡ CCH$_3$</td>
</tr>
<tr>
<td>4</td>
<td>but-1-yne</td>
<td>$C_4H_6$</td>
<td>H–C≡C–C–C–H</td>
<td>HC ≡ CCH$_2$CH$_3$</td>
</tr>
<tr>
<td>5</td>
<td>pent-1-yne</td>
<td>$C_5H_8$</td>
<td>H–C≡C–C–C–C–H</td>
<td>HC ≡ CCH$_2$CH$_2$CH$_3$</td>
</tr>
<tr>
<td>6</td>
<td>hex-1-yne</td>
<td>$C_6H_{10}$</td>
<td>H–C≡C–C–C–C–C–H</td>
<td>HC ≡ CCH$_2$CH$_2$CH$_2$CH$_3$</td>
</tr>
</tbody>
</table>

- Construct models for the above structures of alkynes.

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Systematic(IUPAC) naming of alkynes

The rules for naming alkynes are virtually the same as those of alkenes. The “ene” ending is replaced by “yne”. It is important to indicate the position of the triple bond when writing the name of an alkyne. When numbering the carbon chain, start from the end that will give the triple bond the lowest number. Let us consider the following example:

Copy the above structure in your notebook.

- Number the carbon atoms starting from left hand side to the right.
- Number the carbon atoms from right to left below carbon atoms.
- What number should the triple bond get?

This is a six-carbon chain alkyne. The triple bond would be found between 1 and 2 if you number the carbon atoms from left to right or 5 and 6 if you number the carbon atoms from right to left. Numbering the carbon atoms from left to right is the better choice because it gives the triple bond the lower number. Therefore the name should be hex–1–yne.

Exercise 3.8

Name the following molecules:

1. 

2. 

3. 

4. 

H — C ≡ C — C — C — C — C — H

H — C ≡ C — C — C — C — C — H

H — C ≡ C — C — C — C — C — H

H — C ≡ C — C — C — C — C — H
Rules for naming branched alkynes

1. Determine the largest continuous chain of carbon atoms that has the triple bond between any two carbon atoms. This will give us the parent name of the compound.
2. Number the carbons in the chain so that the triple bond would be between the carbon atoms with the lowest numbers.
3. Identify and name the branching groups attached to this continuous chain of carbons.
4. Write the number of the carbon bearing the branching group in front of the parent name followed by a hyphen.
5. Look at the two carbon atoms between which the triple bond is located. Identify the carbon atom which has the lowest number. For example if the triple bond is found between the second and third carbon atom, the lowest is 2.

Let us try the following example

\[
\begin{align*}
&\text{H} \\
&\text{H-C-H} \\
&\text{H-C=C-C-C-H} \\
&\text{H-H} \\
\end{align*}
\]

Step 1
Identify the longest continuous chain of carbon atoms

\[
\begin{align*}
&\text{H} \\
&\text{H-C-H} \\
&\text{H-C=C-C-C-H} \\
&\text{H-H} \\
\end{align*}
\]

The longest chain is four carbon atoms long.

Step 2
Identify the lowest carbon number on which the triple bond is located.
In the above example the lowest number is 1.
Replace “e” in ene with “y”. Hence, the parent name of the compound is but–1–yne.
Step 3  
Identify branching groups

The branching group is a methyl group  
It is located on carbon number 3.

Step 4  
Write the number of the carbon bearing the branching group before the name of the group, followed by a hyphen.  
• 3 – Methyl

Step 5  
Write the full name of the alkyne.  
To write the name of the alkyne, start with the branching group(s) then the parent name.  
• 3–methylbut-1-yne.

You can now use the above steps to name any branched alkyne.

Exercise 3.9  
1. Name the following alkynes.

(i)
2. Draw the structures of the following compounds
   (i) 4 - methylpent-2-yne
   (ii) 3 - chloro-3-methylbut-1-yne.

Isomerism
Isomerism in alkynes is the same as that of alkenes. There are two types of structural isomerism.
   • chain isomerism
   • position isomerism.

In the following example, we will consider both chain and position isomers of pentyne.

1. Chain isomerism
   (a) The first isomer is when the carbon atoms are in a continuous chain as shown below.

\[ \text{pent-1-yne} \]
The second isomer is formed when a methyl group is shifted to the third carbon.

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{C} \\
\text{H} & \quad \text{C} \quad \text{C} \\
\text{H} & \quad \text{H} \quad \text{H} \\
\text{H} & \quad \text{H} \quad \text{H} \\
\text{H} & \quad \text{H} \quad \text{H} \\
\end{align*}
\]

3-methylbut-1-yne

With the chain isomers, we have two isomers; the straight chain alkyne pent-1-yne and one branched 3-methylbut-1-yne.

2. Position isomerism
We have two position isomers of pentyne, depending on where the triple bond is placed.
- Pent-1-yne
- Pent-2-yne

Structures of the position isomers.

(a) 
\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{C} \\
\text{H} & \quad \text{C} \quad \text{C} \\
\text{H} & \quad \text{H} \quad \text{H} \\
\text{H} & \quad \text{H} \quad \text{H} \\
\end{align*}
\]

Pent-1-yne

(b) 
\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} \quad \text{C} \\
\text{H} & \quad \text{C} \quad \text{C} \\
\text{H} & \quad \text{H} \quad \text{H} \\
\text{H} & \quad \text{H} \quad \text{H} \\
\end{align*}
\]

Pent-2-yne

Exercise 3.10
1. Draw and name all chain and position isomers of hex-1-yne.
2. Draw the structural formula of:
   (a) but-2-yne
   (b) 4-methylpent-2-yne
3. Name the following
   (a) CH≡CCH₂CH₃
   (b) CH₃CH₂CH₂C≡CCH₃
Preparation and properties of ethyne

Experiment 3.4
To prepare and investigate the properties of ethyne.

Apparatus and chemicals
- flat bottomed flask
- thistle funnel or dropping funnel
- delivery tube
- trough
- beehive shelf
- gas jars
- calcium dicarbide
- water

Procedure
1. Cover the bottom of the flask with a thin layer of sand. Ensure that the flask is dry. The reaction between calcium carbide and water produce a lot of heat. Sand is placed in the flask in order to absorb some of the heat produced to prevent the flask from cracking.
2. Add 2-3 lumps of calcium carbide on the sand.
3. Add water dropwise to produce a steady flow of the gas as illustrated in Fig. 3.8.

![Diagram of laboratory preparation of ethyne]

Fig. 3.8: Laboratory preparation of ethyne

- What is the colour of ethyne?
- Does it have a smell?
- Why is ethyne collected over water?
4. For the following tests collect two test tubes full of ethyne.
   (i) Place a burning splint at the mouth of one of the test tubes. What do you observe?
   (ii) To the 2nd test tube of ethyne add about 2-3 cm³ of bromine water and shake well. What do you observe?

5. Record all your observations in your notebook. What do you conclude?
   The reaction that occurs is as in the equation below.
   
   \[
   \text{Calcium dicarbide} + \text{water} \rightarrow \text{calcium hydroxide} + \text{ethyne}
   \]
   \[
   \text{CaC}_2(s) + \text{H}_2\text{O(l)} \rightarrow \text{Ca(OH)}_2(aq) + \text{C}_2\text{H}_2(g)
   \]

**Physical properties**

Ethyne is:
- colourless.
- odourless.
- less dense than air.
- insoluble in water but readily dissolves in organic solvents.

**Trends in physical properties of alkynes**

The table below shows a summary of physical properties of alkynes

*Table 3.9 Trends in physical properties of the first five alkynes*

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
<th>Structural formula</th>
<th>Melting point °C</th>
<th>Boiling point °C</th>
<th>Density g/cm³</th>
<th>Solubility in water</th>
<th>Solubility in organic solvents</th>
</tr>
</thead>
<tbody>
<tr>
<td>C₂H₂</td>
<td>Ethyne</td>
<td>H–C≡C–H</td>
<td>−82</td>
<td>−75</td>
<td>—</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>C₃H₄</td>
<td>Propyne</td>
<td>HCC≡CH₃</td>
<td>−101.5</td>
<td>−23</td>
<td>—</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>C₄H₆</td>
<td>But-1-yne</td>
<td>HC≡CCH₂CH₃</td>
<td>−122</td>
<td>8.1</td>
<td>0.695</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>C₅H₈</td>
<td>Pent-1-yne</td>
<td>HC≡CCH₂CH₂CH₃</td>
<td>−98</td>
<td>40</td>
<td>0.695</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
<tr>
<td>C₆H₁₀</td>
<td>Hex-1-yne</td>
<td>HC≡CCH₂CH₂CH₂CH₃</td>
<td>−124</td>
<td>72</td>
<td>0.719</td>
<td>insoluble</td>
<td>soluble</td>
</tr>
</tbody>
</table>

Alkynes are compounds that have physical properties that are similar to those of alkanes and alkenes.
- They are insoluble in water.
- They are soluble in organic solvents like benzene and tetrachloromethane among others. Their densities increase with increase in number of carbon atoms due to increase in molecular mass.
The boiling points of alkynes increase with increase in number of carbon atoms due to increase in molecular mass and also increase in intermolecular forces of attraction.

Chemical properties of ethyne

1. **Combustion of ethyne**
   Ethyne burns in air with a luminous, smoky flame forming carbon(IV) oxide and water.

   \[
   \text{Ethyne} + \text{oxygen} \rightarrow \text{carbon(IV)oxide} + \text{water}
   \]

   \[
   2\text{C}_2\text{H}_2(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})
   \]

   Generally alkynes are highly explosive when mixed with air and oxygen. In a limited supply of air, ethyne undergoes incomplete combustion to form a mixture of carbon and carbon(II) oxide.

   \[
   \text{Ethyne} + \text{oxygen} \rightarrow \text{carbon} + \text{carbon(II) oxide} + \text{water}
   \]

   \[
   \text{C}_2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{C}(\text{s}) + \text{CO}(\text{g}) + \text{H}_2\text{O}(\text{l})
   \]

2. **Addition reactions of alkynes**
   Because of the unsaturated nature of ethyne, addition reactions occur across the triple bond. Note that two moles of the substance that is reacting are required for complete saturation.

   **(a) Addition of hydrogen**
   When ethyne and hydrogen are passed over nickel catalyst at 150°C or over platinum catalyst at room temperature, ethene is first formed and then it is further reduced to ethane.

   \[
   \text{(i) Ethyne} + \text{hydrogen} \rightarrow \text{ethene}
   \]

   \[
   \text{C}_2\text{H}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_4(\text{g})
   \]

   OR

   \[
   \text{H} - \text{C} \equiv \text{C} - \text{H} + \text{H} - \text{H} \rightarrow \text{H} - \text{C} \equiv \text{C} - \text{H}
   \]

   \[
   \text{H} \quad \text{H}
   \]

   \[
   \text{ethene}
   \]

   \[
   \text{(ii) Ethene} + \text{hydrogen} \rightarrow \text{ethane}
   \]

   \[
   \text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})
   \]
We may represent a halogen molecule with $X_2$ or $X - X$. The general reaction of triple bonds with a halogen is as follows:

(i) $H - C \equiv C - H + X - X \rightarrow H - C \equiv C - H$

(ii) $H - C \equiv C - H + X - X \rightarrow H - C - C - H$

Ethyne reacts explosively with chlorine at room temperature to form hydrogen chloride and carbon.

$$C_2H_2(g) + Cl_2(g) \rightarrow 2C(s) + 2HCl(g)$$

In controlled situation, ethyne reacts with chlorine and bromine as follows;

(a) Ethyne + chlorine $\rightarrow$ 1,2-dichloroethene

$$C_2H_2(g) + Cl_2(g) \rightarrow CHClCHCl(g)$$

$$CHBrCHBr + Br_2(g) \rightarrow CHBr2CHBr_2$$

(i) $H - C \equiv C - H + Cl - Cl \rightarrow H - C \equiv C - H$

$$Cl \quad Cl$$

(ii) $H - C \equiv C - H + Cl - Cl \rightarrow H - C - C - H$

$$Cl \quad Cl \quad Cl \quad Cl$$

1,2-dichloroethene 1,1,2,2-tetrachloroethane

(b) (i) Ethyne + bromine $\rightarrow$ 1,2-dibromoethene

$$C_2H_2(g) + Br_2(g) \rightarrow CHBrCHBr(g)$$

(ii) 1,2-dibromoethene + bromine $\rightarrow$ 1,1,2,2-dibromoethene

$$CHBrCHBr + Br_2 \rightarrow CHBr_2CHBr_2$$
(c) Addition of hydrogen halides

Ethyne reacts with hydrogen halides. The vigour of reaction of ethyne with hydrogen halides increases as follows; HI>HBr>HCl>HF so that hydrogen iodide add most readily to ethyne at room temperature. A similar reaction occurs with hydrogen bromide at 100°C. The reaction with hydrogen chloride occurs very slowly.

(i) Ethyne + hydrogen iodide → iodoethene

\[
\begin{align*}
\text{C}_2\text{H}_2(\text{g}) & + \text{HI(aq)} \rightarrow \text{C}_2\text{H}_3\text{I} \\
\text{H} & \text{C} \equiv \text{C} \text{H} + \text{H} & \text{I} \rightarrow \text{H} \text{C} \equiv \text{C} \text{H} \\
\text{H} & \text{I} & \text{H}
\end{align*}
\]

OR

(ii) Iodoethene + hydrogen iodide → 1,1-diiodoethane

\[
\begin{align*}
\text{C}_2\text{H}_3\text{I} & + \text{HI} \rightarrow \text{C}_2\text{H}_4\text{I}_2 \\
\text{H} & \text{C} \equiv \text{C} \text{H} + \text{H} & \text{I} \rightarrow \text{H} \text{C} \equiv \text{C} \text{H} \\
\text{H} & \text{I} & \text{H}
\end{align*}
\]

When bromine reacts with ethyne, the red-brown colour of bromine disappears. Bromine finally adds to the carbon-carbon triple bond to form 1,1,2,2-tetrabromoethane. The decolourisation of bromine as mentioned earlier is the test used to detect for unsaturation in a molecule.
Follow the above procedure to predict the word equation and chemical equation of addition reaction of hydrogen bromide or hydrogen fluoride with ethyne.

**Uses of alkynes**
1. Ethyne burns in oxygen to give a very luminous flame. This is why it is used to produce light in lamps Fig. 3.9(g).
2. Ethyne is used in oxy-acetylene flame for metal cutting and welding Fig. 3.9(a).
3. Some alkynes are used for manufacture of plastics, synthetic rubbers, synthetic fibres, adhesives and other things Fig. 3.9(b),(c),(d), (e), (f).

![Fig. 3.9: Uses of alkynes](image)

**Exercise 3.11**
1. What is the molecular formula of but-1-yne? Draw its structural formula.
2. Draw and name the structural formula for all the isomers of but-1-yne.
3. Why do you think that alkynes are more reactive than the alkanes and alkenes?
4. Write an equation to show the complete combustion of ethyne in oxygen.
5. List three uses of ethyne.
Summary
1. Organic chemistry is the study of chemistry of the carbon compounds.
2. A homonogous series is a series of compounds, which have the same general formula and similar chemical properties.
3. Alkanes have a general formula of $C_nH_{2n+2}$ where ‘n’ is a whole number.
4. Alkenes have a general formula of $C_nH_{2n}$.
5. Alkynes have a general formula of $C_nH_{2n-2}$.
6. All alkane names end with-ane, alkenes names end with –ene and alkynes end with -yne.
7. Isomerism is the occurrence of two or more compounds with the same molecular formula but different structural formula.
8. Alkenes and alkynes are unsaturated compounds but alkanes are saturated compounds
9. Alkenes contains double bond between the two carbon atoms in the molecule, while alkynes contain a triple bond.

Revision Exercise 3
1. What is the difference between a structural formula and molecular formula? Show this difference using methane and ethane.
2. Study the following formulas: $C_4H_{10}$, $CH_3CH_2CH_2CH_3$, $C_3H_4$, $C_9H_{20}$ and $C_{15}H_{30}$.
   (a) Which of these compounds are:
      (i) alkanes?
      (ii) alkenes?
      (iii) alkynes?
   (b) Name each compound listed above.
3. Draw a structural formula for $CH_3Br$. From which alkane is this derived from?
4. What is the difference between saturated and unsaturated hydrocarbon? Give an example of each.
5. Name the following compounds.
   (a) $\begin{array}{c}
   CH_3---CH_2---CH---CH_2---CH_3 \\
   \end{array}$
   \[ \begin{array}{c}
   \text{CH}_3 \\
   \end{array} \]
   (b) $\begin{array}{c}
   CH_3---CH_2---CH_2 \\
   \end{array}$
   \[ \begin{array}{c}
   \text{CH}_3 \\
   \end{array} \]
6. Write a balanced equation for each of the following reactions.
   (a) Burning of hexane.
   (b) The reaction between hex-2-ene and hydrogen.
   (c) The addition of bromine to but-1-yne.
   (d) The reaction between ethane and chlorine.
   (e) Addition polymerisation of ethene.

7. What is the purpose of the fractional distillation of crude oil? What is the basic principle that is used in fractional distillation? How can the process be demonstrated in the laboratory?
Unit 4: Nitrogen and its compounds

Introduction
Nitrogen is the first element in Group V of the Periodic Table as shown in Fig. 4.1. It is located in Period 2. Using this information, work out the electron arrangement of an atom of nitrogen.

Fig 4.1: Showing the Position of nitrogen in the Periodic Table

Nitrogen has an atomic number 7 and an electron arrangement of 2.5. It occupies about 78% of air in the atmosphere. Nitrogen is one of the most essential nutrients required for plants growth. It is found in animal and plant proteins and also in compounds known as nitrates.

4.1: Laboratory isolation of nitrogen from air
Air is a mixture of gases which include nitrogen, oxygen, carbon(IV) oxide, noble gases and water vapour.

Experiment 4.1
Laboratory isolation of nitrogen from air.

Apparatus and chemicals
- aspirator
- wash bottle
- combustion tube
- copper turnings
- gas jars
- trough
• potassium hydroxide solution
• glass tubing

Procedure
1. Arrange the apparatus as shown in Fig. 4.2

![Diagram of laboratory isolation of nitrogen from air]

2. Heat the copper turnings strongly.
3. Force air out of the aspirator by passing a stream of tap water slowly into the aspirator. Answer these questions in your notebook.
   • What do you observe in the wash bottle containing potassium hydroxide solution?
   • What is the use of this potassium hydroxide solution?
   • Which other solution can we use in place of potassium hydroxide?
   • What do you observe inside the combustion tube?
   • Why should copper turnings be heated?
   • What conclusions can you draw from this experiment?

As air passes from the aspirator and into the potassium hydroxide solution, bubbles are seen. Potassium hydroxide solution absorbs the carbon(IV) oxide in the air. The equation for the reaction between the potassium hydroxide solution and carbon(IV) oxide is as follows:

$$\text{2KOH(aq)} + \text{CO}_2(g) \rightarrow \text{K}_2\text{CO}_3(aq) + \text{H}_2\text{O(l)}$$
Sodium hydroxide solution can be used instead of potassium hydroxide to absorb carbon(IV) oxide.

\[
\text{Sodium hydroxide + carbon(IV) oxide} \rightarrow \text{sodium carbonate + water}
\]

\[
2\text{NaOH(aq) + CO}_2\text{(g)} \rightarrow \text{Na}_2\text{CO}_3\text{(aq)} + \text{H}_2\text{O(l)}
\]

Hot copper turnings combine with oxygen in the air when heated to form black copper(II) oxide.

\[
\text{Copper + oxygen} \rightarrow \text{copper(II) oxide}
\]

\[
2\text{Cu(s) + O}_2\text{(g)} \rightarrow 2\text{CuO(s)}
\]

As nitrogen is almost insoluble in water, it is collected over water. However it will have some impurities such as noble gases, water vapour and some unreacted carbon(IV) oxide and oxygen.

If required dry, it may be passed through concentrated sulphuric acid after leaving the combustion tube. It is then collected in a syringe or by upward delivery.

When oxygen and carbon(IV) oxide are removed from air, about 99% pure nitrogen can be obtained.

Why is it difficult to remove impurities of noble gases from air by chemical means?

**4.2: Industrial isolation of nitrogen from air**

The best way to obtain nitrogen on large scale is to separate it from liquid air. This is done through the process called **fractional distillation** of liquid air.

Apart from nitrogen, air also contains water vapour, oxygen, carbonate(IV) oxide, noble gases, dust particles, among other things. The isolation process therefore aims at eliminating these and remaining with only nitrogen. The process starts by first removing water vapour, dust particles and carbon(IV) oxide. This ensures that carbon(IV) oxide and water vapour do not solidify because if this happened, they would block the pipes.

After carbon(IV) oxide and water vapour have been removed, the air is compressed to about 200 atmospheres. This makes the air hot (just like air gets hot when we pump it into a bicycle tube). It is then allowed to expand through a jet. This makes it very cold and some of it turns into liquid.

The compression and expansion are repeated several times and each time the air gets colder. By the time it reaches −200°C, the gases, namely nitrogen, oxygen and argon have become liquid except neon and helium. These two noble gases are then removed. Study Fig. 4.3.
Since nitrogen and oxygen have different boiling points, they are separated by fractional distillation. When the liquid air is warmed slowly, liquid nitrogen which has a lower boiling point (-196°C) distills first. The nitrogen obtained is then stored under pressure in steel cylinders. The remaining liquid which contains oxygen and argon is heated further. Argon whose boiling point is -186°C distils leaving oxygen which has a higher boiling point of -183°C. Once separated, the two gases are also stored and sold commercially in steel cylinders.

Experiment 4.2
Laboratory preparation of nitrogen from ammonium nitrite

Apparatus and chemicals
- round-bottomed flask
- bunsen burner/or other source of heat
- trough
- gas jars
- sodium nitrite
- ammonium chloride
**Procedure**

1. Make a solution of a mixture of sodium nitrite and ammonium chloride in water.
2. Put the solution in a round-bottomed flask.
3. Arrange the apparatus as shown in Fig. 4.4.

![Diagram](image)

**Fig. 4.4: Laboratory preparation of nitrogen**

4. Warm the flask slightly and then remove the heating apparatus. What do you observe in the flask and gas jar? Record your observations in your notebook.
5. Collect five gas jars of nitrogen for experiment 4.3

**Note:** Solid ammonium nitrite can be heated to produce nitrogen. However, obtaining nitrogen this way is not advisable. Suggest a reason for this.

When a solution of sodium nitrite and ammonium chloride is warmed in a flask, effervescence is observed. The reaction increases rapidly even without further heating. Bubbles are seen in the gas jar as a colourless gas is collected over water.

The reaction between sodium nitrite and ammonium chloride first forms ammonium nitrite.
The ammonium nitrite is then decomposed to produce nitrogen.

\[
\text{NH}_4\text{NO}_2(aq) \rightarrow \text{N}_2(g) + 2\text{H}_2\text{O}(l)
\]

The overall equation for the reaction is as follows:

\[
\text{NaNO}_2(aq) + \text{NH}_4\text{Cl}(aq) \rightarrow \text{NaCl}(aq) + 2\text{H}_2\text{O}(l) + \text{N}_2(g)
\]

If required dry, nitrogen gas can be passed through concentrated sulphuric acid to remove water vapour.

The gas is then collected in a syringe or by upward delivery. Nitrogen can also be prepared by direct heating of ammonium nitrite. However, when ammonium nitrite is in small quantities, the reaction can be explosive.

### 4.3: Properties of nitrogen

#### Experiment 4.3
To determine some of the properties of nitrogen

**Apparatus and chemicals**
- wooden splint
- blue and red litmus papers
- aqueous calcium hydroxide
- magnesium ribbon
- sulphur powder
- five gas jars full of dry nitrogen
- a pair of tongs

**Procedure**
Perform the experiments described in Table 4.1 using the nitrogen gas collected in Experiment 4.2. Draw Table 4.1 in your notebook and record your observations and conclusions.
Table 4.1: Investigating the properties of nitrogen

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Observations</th>
<th>Conclusions</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Observe the colour and smell of nitrogen in the first gas jar.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. Put a glowing splint into the first gas jar.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. Insert a burning splint into the first gas jar.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4. Insert wet, blue and red litmus papers into the second gas jar.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5. Add a few drops of aqueous calcium hydroxide into the third gas jar. Shake gently.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6. Lower a burning magnesium ribbon in the fourth jar. Add water to the products. Smell the gas produced.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>7. Lower burning sulphur on a deflagrating spoon in the fifth gas jar.</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Physical properties of nitrogen**

Nitrogen
- has no colour.
- has no smell.
- is almost insoluble in water and hence can be collected over water.
- is lighter than air and therefore can be collected by upward delivery.

**Chemical properties of nitrogen**

- **Combustion**
Nitrogen does not burn nor support combustion. This is why a glowing or burning splint does not continue glowing or burning in nitrogen.
• **Effect on litmus indicator**
Nitrogen is neutral. Litmus paper show no colour change.

• **Effect on aqueous calcium hydroxide**
Nitrogen has no effect on calcium hydroxide.

The above three chemical properties are used to distinguish nitrogen from other gases like oxygen, carbon(II) oxide, carbon(IV) oxide, sulphur(IV) oxide, ammonia, hydrogen chloride, nitrogen(IV) oxide and other common gases.

• **Reaction with magnesium and sulphur**
Nitrogen reacts with burning magnesium to form magnesium nitride.

\[
3\text{Mg}(s) + \text{N}_2(g) \rightarrow \text{Mg}_3\text{N}_2(s)
\]

Magnesium nitride reacts with water to form magnesium hydroxide solution and ammonia gas.

\[
\text{Mg}_3\text{N}_2(s) + 6\text{H}_2\text{O}(l) \rightarrow 3\text{Mg(OH)}_2(aq) + 2\text{NH}_3(g)
\]

However, when burning sulphur is introduced into a gas jar of nitrogen, no reaction is observed. This is because the heat produced from burning sulphur is not sufficient to break the triple covalent bonds (N≡N) of the nitrogen molecules, N\(_2\). Nitrogen is inert at room temperature. It does not react with other elements at room temperature.

However, the high temperature produced by burning metals like magnesium break down the covalent bonds of the nitrogen molecules to form single atoms of nitrogen which react with magnesium.

\[
\text{N≡N} \xrightarrow{\text{heat}} 2\text{N} \text{ (very high temperature needed).}
\]

### 4.4: Uses of nitrogen

1. It is used in the Haber process to manufacture ammonia gas (Fig. 4.5a).
2. The boiling point of liquid nitrogen is almost -200°C. Therefore it is used to freeze substances quickly. It can be used to store body tissues that are required to last a long period of time. For instance,
• tissues used in hospitals can be stored for many years.
• bulls’ semen for artificial insemination can be stored for use in the future.

It can also be used to mend leaking pipes. When liquid nitrogen is poured on the pipe it freezes the liquid inside while repair is done.

3. It is used in food processing. When food is being packed, nitrogen is used to keep off oxygen and the food stays fresh for a long time because there is no oxidation. Most bacteria cannot survive if there is no oxygen. (Fig. 4.5 c).

4. Because of its inert nature at low temperature, it is pumped in the ships’ tanks that transport crude oil to remove any oxygen. This prevents dangerous explosion that could occur from crude oil vapours. (Fig. 4.5d).

Fig. 4.5: Uses of nitrogen
4.5 Oxides of nitrogen

There are three common oxides of nitrogen. These are:

(i) Nitrogen(I) oxide, N₂O.
(ii) Nitrogen(II) oxide, NO.
(iii) Nitrogen(IV) oxide, NO₂.

These oxides are gases at room temperature.

4.6: Nitrogen(I) oxide

Experiment 4.4

Laboratory preparation of nitrogen(I) oxide

Nitrogen(I) oxide is prepared by heating ammonium nitrate.

Apparatus and chemicals

• round-bottomed flask
• delivery tube
• gas jars
• beehive shelf
• ammonium nitrate.

Procedure

1. Place ammonium nitrate in a round-bottomed flask and arrange the apparatus as shown in Fig. 4.6

Fig. 4.6: Laboratory preparation of nitrogen(I) oxide
2. Heat ammonium nitrate. What do you observe in the flask and gas jar?
   • Collect about six gas jars full of this gas for Experiment 4.5.

   When ammonium nitrate is heated it melts and forms bubbles of nitrogen(I) oxide gas. Bubbles of this gas are seen in the gas jar.

   \[
   \text{Ammonium nitrate} \xrightarrow{\text{heat}} \text{nitrogen(I) oxide} + \text{water}
   \]

   \[
   \text{NH}_4\text{NO}_3(s) \xrightarrow{\text{heat}} \text{N}_2\text{O(g)} + 2\text{H}_2\text{O(l)}
   \]

   Caution
   Ammonium nitrate should not be heated further if the quantity remaining is small because it may explode. It is advisable to use a mixture of potassium nitrate and ammonium chloride instead of ammonium nitrate.

   If the gas is required dry, pass it through anhydrous calcium chloride in a U-tube to remove water vapour and then collect it by downward delivery.

4.7: Properties of nitrogen(I) oxide

Experiment 4.5
To determine some of the properties of nitrogen(I) oxide

Apparatus and chemicals
   • gas jars full of nitrogen(I) oxide
   • blue and red litmus papers
   • wooden splint
   • magnesium ribbon
   • phosphorus
   • sulphur powder
   • charcoal powder
   • copper turnings
   • a pair of tongs

Procedure
Perform the following experiments with the nitrogen(I) oxide collected in experiment 4.4. Draw Table 4.2 in your notebook and record your observation and conclusions.
### Table 4.2: Investigating properties of nitrogen(I) oxide

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Observations</th>
<th>Conclusion</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td>Check the colour and the smell of nitrogen(I) oxide in the first gas jar.</td>
<td></td>
</tr>
<tr>
<td>2.</td>
<td>Insert a glowing splint into the first gas jar</td>
<td></td>
</tr>
<tr>
<td>3.</td>
<td>Insert a burning splint into the second gas jar</td>
<td></td>
</tr>
<tr>
<td>4.</td>
<td>Insert wet, blue and red litmus papers into the third gas jar</td>
<td></td>
</tr>
<tr>
<td>5.</td>
<td>Lower a burning magnesium ribbon into the fourth gas jar</td>
<td></td>
</tr>
<tr>
<td>6.</td>
<td>Lower burning (i) Phosphorus into the fourth jar (ii) Sulphur into the sixth jar (iii) Carbon into the seventh jar</td>
<td></td>
</tr>
</tbody>
</table>

• How can we differentiate nitrogen(I) oxide from oxygen?

**Properties of nitrogen(I) oxide**

(a) **Physical properties**

Nitrogen(I) oxide:
• colourless.
• has a slight sweet smell.
• is fairly soluble in water and hence it can be collected over water.
• is denser than air and therefore can be collected by downward delivery.
(b) Chemical properties

• Combustion

Nitrogen(I) oxide relights a glowing splint just like oxygen because the heat of the glowing splint decomposes (splits) this gas into nitrogen and oxygen. The oxygen produced supports combustion. However nitrogen(I) oxide does not burn.

\[
\text{Nitrogen(I) oxide } \xrightarrow{\text{heat}} \text{nitrogen + oxygen}
\]

\[
2\text{N}_2\text{O}(g) \xrightarrow{\text{heat}} 2\text{N}_2(g) + \text{O}_2(g)
\]

**Effect on litmus**

Nitrogen(I) oxide is neutral. Litmus paper do not change colour.

**Effect on burning elements**

Burning magnesium continues to burn with brilliant flame in nitrogen(I) oxide because heat from burning magnesium splits the gas into oxygen and nitrogen. Magnesium burns in this oxygen forming magnesium oxide.

\[
\text{Mg(s)} + \text{N}_2\text{O}(g) \rightarrow \text{MgO(s)} + \text{N}_2(g)
\]

Similarly, phosphorus burns with a bright yellow flame, forming phosphorus(V) oxide and nitrogen.

\[
2\text{P}(s) + 5\text{N}_2\text{O}(g) \rightarrow \text{P}_2\text{O}_5(g) + 5\text{N}_2(g)
\]

**Feebly burning** sulphur does not produce enough heat to decompose nitrogen(I) oxide to nitrogen and oxygen. However, if it is **vigorously burning** sulphur(IV) oxide and nitrogen will be formed.

\[
\text{S(s)} + 2\text{N}_2\text{O}(g) \rightarrow \text{SO}_2(g) + 2\text{N}_2(g)
\]

**Vigorously burning** carbon forms carbon(IV) oxide and nitrogen.

\[
\text{C(s)} + 2\text{N}_2\text{O}(g) \rightarrow \text{CO}_2(g) + 2\text{N}_2(g)
\]

In the above reactions, nitrogen(I) oxide is reduced to nitrogen gas while the other elements are oxidised.
Experiment 4.6
Reaction of nitrogen(I) oxide with copper

Apparatus and chemicals
- copper turnings
- combustion tube
- delivery tube
- gas jar
- trough
- beehive shelf

Procedure
1. Arrange the apparatus as shown in Fig. 4.7.

Fig. 4.7: Reaction of nitrogen(I) oxide with copper

2. Heat the combustion tube as shown.
3. Pass dry nitrogen(I) oxide over the hot copper turnings for a few minutes.
4. Record your observations in your notebook.
   - What was the original colour of copper?
   - What is the final colour of the product in the combustion tube?
   - Which other metal can we use in place of copper? Write a chemical equation for the reaction.

When nitrogen(I) oxide is passed over hot copper turnings, the copper which is red-brown in colour turns black.
Nitrogen(I) oxide is reduced by the hot copper to nitrogen gas. Black copper(II) oxide is also formed. The equation for the reaction is as follows:
Copper + nitrogen(I) oxide $\rightarrow$ copper(II) oxide + nitrogen
Cu(s) + $N_2O(g)$ $\rightarrow$ CuO(s) + $N_2(g)$

What do we call this type of reaction? What do you conclude from this reaction?
Since nitrogen(I) oxide is an oxidising agent, if we use a metal like zinc it will be oxidised to zinc oxide and the gas reduced to nitrogen. See also table 4.3.

Zinc + nitrogen(I) oxide $\rightarrow$ zinc oxide + nitrogen
Zn(s) + $N_2O(g)$ $\rightarrow$ ZnO(s) + $N_2(g)$

4.8: Test for nitrogen (I) oxide
Nitrogen(I) oxide can be confused with oxygen because they both rekindle a glowing splint. However, for nitrogen(I) oxide, the splint should be brightly glowing. Oxygen has no smell and also reacts with nitrogen(I) oxide at room temperature to form red-brown nitrogen(IV) oxide as shown in the following equation.

$\text{Nitrogen(I) oxide} + \text{oxygen} \rightarrow \text{nitrogen(IV) oxide}$
$2\text{NO(g)} + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$

Table 4.3: Differences between oxygen and nitrogen (I) oxide

<table>
<thead>
<tr>
<th>Properties</th>
<th>Oxygen</th>
<th>Nitrogen(I) oxide</th>
</tr>
</thead>
<tbody>
<tr>
<td>Smell</td>
<td>None</td>
<td>Slightly sweet</td>
</tr>
<tr>
<td>Solubility in water</td>
<td>Slightly solube</td>
<td>Fairly soluble</td>
</tr>
<tr>
<td>Effect on nitrogen(II)</td>
<td>Forms a red-brown gas (NO$_2$)</td>
<td>No reaction</td>
</tr>
<tr>
<td>oxide gas (NO$_2$)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Effect on hot copper</td>
<td>No gas formed</td>
<td>Nitrogen gas(N$_2$) formed</td>
</tr>
</tbody>
</table>

4.9: Uses of nitrogen(I) oxide
The temperatures of flames of hydrogen-oxygen (2000°C) or ethyne-oxygen (3000°C) mixtures are very high, for analytical work where lower temperatures (1000°C – 1500°C) are needed nitrogen(I) oxide-oxygen mixture is used. For example in Atomic Absorption Spectrophotometer used for measuring intensity of light.

200
4.10: Nitrogen(II) oxide
Nitrogen(II) oxide is prepared in the laboratory by reacting copper turnings with 50% concentrated nitric acid.

Experiment 4.7
Laboratory preparation and properties of nitrogen(II) oxide

Apparatus and chemicals
- thistle funnel or dropper
- flat-bottomed flask
- delivery tubes
- gas jar
- beehive shelf
- trough
- copper turnings
- 50% concentrated nitric acid
- distilled water
- measuring cylinder
- beaker

Procedure
1. Put some copper turnings in a flat-bottomed flask.
2. Arrange the apparatus as shown in Fig. 4.8.
3. Add 50 cm$^3$ of concentrated nitric acid into 50 cm$^3$ of distilled water in a beaker.
   Stir the mixture.
4. Add this solution to the copper turnings in the flask.
• What do you observe in the flask and the gas jar?
• What is the colour of the solution left in the flask?

5. Expose the gas collected to air.
• What do you observe?
• Explain your observation.

When the 50 per cent concentrated nitric acid is added to the copper turnings, vigorous effervescence occurs and the flask is filled with red-brown fumes of nitrogen(IV) oxide. Bubbles are observed in the gas jar as a colourless gas fills it. A green solution of copper(II) nitrate is formed in the flask. The red-brown fumes are partly from the reaction of nitric acid and copper and partly from the oxidation of nitrogen(II) oxide by the atmospheric oxygen in the flask and delivery tube.

Copper + nitric acid $\rightarrow$ copper(II) + water + nitrogen(II) nitrate
3Cu(s) + 8HNO$_3$(aq) $\rightarrow$ 3Cu(NO$_3$)$_2$(aq) + 4H$_2$O(l) + 2NO(g) (colourless gas)

Nitrogen(II) oxide + oxygen $\rightarrow$ nitrogen(IV) oxide
2NO(g) + O$_2$(g) $\rightarrow$ 2NO$_2$(g) (red-brown fumes)

What happens to the red-brown fumes in the flask? The fumes dissolve in the water in the trough leaving a colourless gas of nitrogen(II) oxide which is collected in the gas jar.

**Physical properties of nitrogen(II) oxide**

*Nitrogen(II) oxide is:*
• a colourless gas.
• almost insoluble in water.
• slightly denser than air.

It is difficult to know the smell of nitrogen(II) oxide because before one can detect its smell, it is quickly oxidised to nitrogen(IV) oxide hence the smell of nitrogen (IV) oxide.

**Chemical properties of nitrogen(II) oxide**

• **Reaction with air**
When exposed to air, nitrogen(II) oxide reacts with oxygen in the air to form nitrogen(IV) oxide. This is a simple test for nitrogen(II) oxide.

• **Effect on litmus indicator**
Nitrogen(II) oxide is neutral to litmus indicator.
**Effect on burning elements**

Nitrogen(II) oxide is more stable than nitrogen(I) oxide. Hence it extinguishes all burning substances except very hot flames of magnesium and phosphorus. These elements reduce nitrogen(II) oxide to nitrogen.

\[
\begin{align*}
\text{Magnesium + nitrogen(II)} & \rightarrow \text{magnesium + nitrogen} \\
2\text{Mg(s)} + 2\text{NO(g)} & \rightarrow 2\text{MgO(s)} + \text{N}_2(g)
\end{align*}
\]

\[
\begin{align*}
\text{Phosphorus + Nitrogen(II)} & \rightarrow \text{Phosphorus(V) + nitrogen} \\
4\text{P(s)} + 10\text{NO(g)} & \rightarrow 2\text{P}_2\text{O}_5 + 5\text{N}_2(g) \\
& \text{(white fumes)}
\end{align*}
\]

Red hot copper can also reduce it to nitrogen.

\[
\begin{align*}
\text{Copper + nitrogen(II)} & \rightarrow \text{copper(II) oxide + nitrogen} \\
2\text{Cu(s)} + 2\text{NO(g)} & \rightarrow 2\text{CuO(s)} + \text{N}_2(g) \\
& \text{black}
\end{align*}
\]

**Reaction with iron(II) sulphate**

When nitrogen(II) oxide reacts with cold freshly prepared iron(II) sulphate solution, a brown solution of nitroso-iron(II) sulphate is formed.

The equation of the reaction is as follows:

\[
\begin{align*}
\text{Iron(II) + nitrogen(II)} & \rightarrow \text{nitroso-iron(II)} \\
\text{FeSO}_4(aq) + \text{NO(g)} & \rightarrow \text{FeSO}_4 \cdot \text{NO(aq)}
\end{align*}
\]

**Reaction with hydrogen**

When an electrical spark is passed through a mixture of nitrogen(II) oxide and hydrogen it is reduced to nitrogen.

\[
\begin{align*}
\text{Nitrogen(II) + hydrogen} & \rightarrow \text{nitrogen + water} \\
2\text{NO(g)} + 2\text{H}_2(g) & \rightarrow \text{N}_2(g) + 2\text{H}_2\text{O(l)}
\end{align*}
\]

4.11: Uses of nitrogen(II) oxide

Nitrogen(II) oxide gas is not easy to handle. This is because it is readily oxidised by oxygen in the air to nitrogen(IV) oxide. It is however useful in the following ways:
(i) It is used as an intermediary in the industrial manufacture of nitric acid.
(ii) During lightning, nitrogen and oxygen in the atmosphere combine to form nitrogen(II) oxide. It therefore acts as an intermediary in the formation of very dilute nitric acid which is washed into the soil. This increases the amount of nitrates in soil.

4.12: Nitrogen(IV) oxide
Nitrogen(IV) oxide is prepared in the laboratory by reacting copper turnings with concentrated nitric acid.

Experiment 4.8:
Laboratory preparation of nitrogen(IV) oxide.

Apparatus and chemicals
- thistle funnel or dropper
- flat-bottomed flask
- delivery tube
- 3 gas jars
- cardboard cover
- copper turnings
- concentrated nitric acid

Caution: Nitrogen(IV) oxide is poisonous. This experiment should therefore be performed in a fume cupboard or in an open space.

Procedure
1. Put copper turnings in a flat-bottomed flask
2. Arrange the apparatus as shown in Fig. 4.9
3. Add concentrated nitric acid to the copper turnings in the flask.
4. Record in your notebook what you observe in the flask and the gas jar.
   • What is the colour of the resulting solution in the flask?
5. Collect seven jars full of the gas. Also collect some in a test tube and cork it.
   These will be used in Experiment 4.9.

Concentrated nitric acid reacts immediately with copper to form red-brown fumes of nitrogen(IV) oxide. These fumes fill the flask and the gas jar. A green solution of copper(II) nitrate is left in the flask.

\[
\text{Copper} + \text{nitric acid} \rightarrow \text{copper(II) nitrate} + \text{water} + \text{nitrogen(IV) oxide}
\]

\[
\text{Cu(s)} + 4\text{HNO}_3(\text{l}) \rightarrow \text{Cu(NO}_3)_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) + 2\text{NO}_2(\text{g})
\]

4.13: Properties of nitrogen(IV) oxide

Physical properties of nitrogen(IV) oxide
• it is a red-brown gas
• it has a irritating smell
• it is denser than air
• it is soluble in water

Chemical properties of nitrogen(IV) oxide

Experiment 4.9
To determine the chemical properties of nitrogen(IV) oxide.

Apparatus and chemicals
• six gas jars full of nitrogen(IV) oxide
• litmus papers (blue and red)
• magnesium ribbon
• phosphorus
• sulphur
• copper turnings
• a pair of tongs
• wooden splints

Procedure
Carry out the experiments contained in Table 4.4 using the nitrogen(IV) oxide collected in Experiment 4.8.
Draw Table 4.4 in your notebook and record your observations and conclusions.
Combustion

Nitrogen(IV) oxide does not burn nor support combustion. This is why it puts off glowing and burning splints.

Reaction with water

Nitrogen(IV) oxide reacts with water to form nitric acid and nitrous acid.

\[
\text{Nitrogen(IV)} + \text{water} \rightarrow \text{nitric acid} + \text{nitrous acid}
\]

\[
2\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HNO}_3(aq) + \text{HNO}_2(aq)
\]

The gas is acidic. The solution formed turns wet blue litmus paper red.
• **Effects on burning elements**

Burning magnesium continues to burn in nitrogen(IV) oxide forming magnesium oxide and nitrogen.

\[
\text{Magnesium + nitrogen(IV) } \rightarrow \text{ magnesium + nitrogen oxide}
\]

\[
4\text{Mg(s)} + 2\text{NO}_2(g) \rightarrow 4\text{MgO(s)} + \text{N}_2(g)
\]

The heat from burning magnesium decomposes this gas into nitrogen and oxygen. This oxygen supports the burning of magnesium.

The same type of reaction takes place if **vigorously** burning phosphorus and sulphur burn in nitrogen(IV) oxide separately.

\[
\text{Phosphorus + nitrogen(IV) } \rightarrow \text{ phosphorus(V) + nitrogen oxide}
\]

\[
8\text{P(s)} + 10\text{NO}_2(g) \rightarrow 4\text{P}_2\text{O}_5(s) + 5\text{N}_2(g)
\]

\[
\text{Sulphur + nitrogen(IV) } \rightarrow \text{ sulphur(IV) + nitrogen oxide}
\]

\[
2\text{S(s)} + 2\text{NO}_2(g) \rightarrow 2\text{SO}_2(g) + \text{N}_2(g)
\]

Nitrogen(IV) oxide oxidises red hot copper while itself is reduced to nitrogen.

\[
\text{Copper + nitrogen(IV) } \rightarrow \text{ copper(II) + nitrogen oxide}
\]

\[
4\text{Cu(s)} + 2\text{NO}_2(g) \rightarrow 4\text{CuO(s)} + \text{N}_2(g)
\]

Heated nitrogen(IV) oxide easily reacts with burning substances because heat decomposes it to nitrogen(II) oxide and oxygen.

\[
\text{Nitrogen(IV) } \stackrel{\text{heat}}{\rightarrow} \text{nitrogen(II) + oxygen oxide}
\]

\[
2\text{NO}_2(g) \stackrel{\text{heat}}{\rightarrow} 2\text{NO(g)} + \text{O}_2(g)
\]

Then nitrogen(II) oxide is decomposed to nitrogen and oxygen.

\[
\text{Nitrogen(II) } \rightarrow \text{ nitrogen + oxygen oxide}
\]

\[
2\text{NO(g)} \rightarrow \text{N}_2(g) + \text{O}_2(g)
\]
The oxygen produced in the two reactions support combustion of burning substances.

- **Reaction with sodium hydroxide solution**
  Nitrogen(IV) oxide reacts with sodium hydroxide solution to form two salts i.e. sodium nitrate (NaNO₃) and sodium nitrite (NaNO₂). The two salts are formed because nitrogen(IV) oxide forms two acids, nitric acid (HNO₃) and nitrous acid (HNO₂) when dissolved in water.

\[
\text{Nitrogen(IV) + water} \rightarrow \text{nitric acid + nitrous acid oxide}
\]

\[
2\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HNO}_3(\text{aq}) + \text{HNO}_2(\text{aq})
\]

Then:

\[
\text{Sodium hydroxide} + \text{nitric acid} \rightarrow \text{sodium nitrate} + \text{water}
\]

\[
\text{NaOH}(\text{aq}) + \text{HNO}_3(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + \text{H}_2\text{O}(l)
\]

\[
\text{Sodium hydroxide} + \text{nitrous acid} \rightarrow \text{sodium nitrite} + \text{water}
\]

\[
\text{NaOH}(\text{aq}) + \text{HNO}_2(\text{aq}) \rightarrow \text{NaNO}_2(\text{aq}) + \text{H}_2\text{O}(l)
\]

The overall equation for the reaction between nitrogen(IV) oxide and sodium hydroxide is as shown below.

\[
\text{Sodium} + \text{nitrogen(IV) oxide} \rightarrow \text{sodium nitrate + sodium nitrite + water}
\]

\[
2\text{NaOH}(\text{aq}) + 2\text{NO}_2(g) \rightarrow \text{NaNO}_3(\text{aq}) + \text{NaNO}_2(\text{aq}) + \text{H}_2\text{O}(l)
\]

**4.14: Alternative laboratory method of preparation of nitrogen(IV) oxide**

Nitrogen(IV) oxide can also be prepared by heating the nitrates of calcium, magnesium, zinc, iron, lead and copper. They form nitrogen(IV) oxide, oxygen and their respective metal oxides.

Lead(II) nitrate is preferably used because it has no water of crystallisation which can dissolve some of the nitrogen(IV) oxide produced.

**Experiment 4.10**

Laboratory preparation of nitrogen(IV) oxide from lead(II) nitrate

**Apparatus and chemicals**
- hard glass test-tube
- delivery tube
- U-tube
- beaker
- lead(II) nitrate
- ice-salt mixture
Procedure
1. Put some crystals of lead(II) nitrate in a hard glass test-tube.
2. Arrange the apparatus as shown in Fig. 4.10.

![Fig. 4.10: Laboratory preparation of nitrogen(IV) oxide]

3. Heat the lead(II) nitrate. Note the cracking sound produced.
4. Test the gas coming off the U-tube with a glowing splint. Record your observations.
   - What is the colour of the hot residue and cold residue in the test-tube?
   - What is the colour of the liquid in the U–tube?
   - What is your conclusion from this experiment?

Lead(II) nitrate produces lead(II) oxide, nitrogen(IV) oxide and oxygen on heating. Lead(II) oxide is red-brown when hot and yellow when cold.

\[
\text{Lead(II) nitrate} \xrightarrow{\text{heat}} \text{lead(II) oxide} + \text{nitrogen(IV) oxide} + \text{oxygen}
\]

\[
2\text{Pb(NO}_3\text{)}_2(s) \xrightarrow{\text{heat}} 2\text{PbO(s)} + 4\text{NO}_2(g) + \text{O}_2(g)
\]

- red-brown hot
- yellow cold

Nitrogen(IV) oxide is easily liquified by a freezing mixture (ice-salt mixture) to form a green liquid which can also be yellow if pure. This liquid is called dinitrogen tetraoxide (N\(_2\)O\(_4\)). Therefore in liquid form nitrogen(IV) oxide exists as yellow dinitrogen tetraoxide (N\(_2\) O\(_4\)).

Dissociation of dinitrogen tetraoxide
Dinitrogen tetraoxide dissociates on heating at any temperature between 22°C and 150°C. At 150°C only NO\(_2\) is present and the colour darkens.
Dinitrogen tetraoxide $\xrightarrow{\text{heat}}$ nitrogen(IV) oxide $\xrightarrow{\text{cool}}$

\[ \text{N}_2\text{O}_4(l) \quad \longrightarrow \quad 2\text{NO}_2(g) \]

light yellow $\quad$ red brown

Further heating dissociates nitrogen(IV) oxide to nitrogen(II) oxide and oxygen. Thus:

Dinitrogen tetroxide $\xrightarrow{\text{heat}}$ nitrogen(IV) oxide $\xrightarrow{\text{cool}}$ nitrogen(II) oxide $\xrightarrow{\text{heat}}$ oxygen

\[ \text{N}_2\text{O}_4(l) \quad \xrightarrow{\text{heat}} \quad 2\text{NO}_2(g) \quad \xrightarrow{\text{cool}} \quad 2\text{NO}(g) \quad + \quad \text{O}_2(g) \]

**Table 4.5: Comparison of oxides of nitrogen**

<table>
<thead>
<tr>
<th>Nitrogen(I) oxide (N$_2$O)</th>
<th>Nitrogen(II) oxide (NO)</th>
<th>Nitrogen(IV) oxide (NO$_2$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Colourless gas; has slight sweet smell</td>
<td>Colourless gas which turns red-brown in air</td>
<td>Red-brown gas; irritating smell</td>
</tr>
<tr>
<td>2. Neutral to litmus</td>
<td>Neutral to litmus</td>
<td>Turns litmus red</td>
</tr>
<tr>
<td>3. Fairly soluble in water</td>
<td>Almost insoluble in water</td>
<td>Soluble in water</td>
</tr>
<tr>
<td>4. Denser than air</td>
<td>Slightly denser than air</td>
<td>Denser than air</td>
</tr>
<tr>
<td>5. Rekindles a glowing splint</td>
<td>Extinguishes a burning splint</td>
<td>Extinguishes a burning splint</td>
</tr>
</tbody>
</table>

**4.15: Uses of nitrogen(IV) oxide**

1. It is used as an intermediary in the manufacture of nitric acid (Ostwald Process).
2. It is an intermediary in the formation of very dilute nitric acid during lightning. This acid is washed into the soil, thereby increasing the amounts of nitrates in the soil.

**Exercises 4.1**

1. When air is compressed and cooled to $-200^\circ$C a blue liquid is formed.
   (i) What is the composition of this liquid?
   (ii) Name the gas from the mixture at $-196^\circ$C and the process used.
2. Study the flow diagram below for the preparation of nitrogen and answer the questions that follow.
State the use of each of the following:
(i) Aqueous NaOH
(ii) Copper turnings
(iii) Write the equation for the reaction in step 3.
(iv) Name gas X
(v) Name a major impurity in gas X.

3. A white salt produces a red-brown gas and a colourless gas that re-lights a glowing splint. The residue is red-brown when hot but yellow when cold.
   (i) Write the formulae of the ions in the white salt.
   (ii) Write the chemical equation for the reaction.
   (iii) Name the two products formed when the red-brown gas is dissolved in water.

4. If 50 per cent concentrated nitric acid is dropped into copper turnings, a colourless gas and later a red-brown gas is produced.
   (i) Name the colourless gas.
   (ii) Name the red-brown gas.
   (iii) Write a chemical equation for the reaction which leads to the formation of the red-brown gas.
   (iv) Why is it not possible to identify the smell of the colourless gas?

5. A concentrated acid A, is added to copper turnings and water in a flask. A red-brown gas B, is liberated.
   In a second experiment, air is removed by filling the flask with nitrogen before adding the acid. A colourless gas C, is produced.
   (i) Name the substances
       A
       B
       C
   (ii) State and explain what is observed when burning magnesium is lowered into a gas jar full of the colourless gas.
   (iii) State and explain what happens when a glowing splint is inserted in a test tube which had previously heated nitrogen(IV) oxide.

4.16: Ammonia

Introduction
Ammonia is a compound of nitrogen and hydrogen. At room temperature and pressure, it is gaseous. Very small traces of ammonia gas are found in the air. Bacteria in the soil
break down nitrogen compounds in manure to form ammonia. The smell of ammonia is always noticeable around barns and stables where farm animals are housed.

4.17: Laboratory preparation of ammonia gas
Ammonia gas is usually prepared in the laboratory using an ammonium salt and a base.

Experiment 4.11
Laboratory preparation of ammonia.

Apparatus and chemicals
- round-bottomed flask
- glass tubing
- drying tower
- several gas jars
- mortar and pestle
- cardboard cover
- calcium hydroxide
- ammonium chloride
- calcium oxide
- litmus papers

Procedure
1. Place small amounts of calcium hydroxide and ammonium chloride in a mortar.
2. Grind the mixture well using the pestle and put it in a round-bottomed flask.
   Arrange the apparatus as shown in Fig. 4.11.

Caution: Make sure that the flask is sloping before the experiment starts.
3. Heat the mixture.
   • What do you observe on the cool part of the flask and the delivery tube?
   • Record the observations in your notebook.

4. Check whether the gas jar is full, by putting a moist red litmus paper at the mouth of the jar. It will turn blue if the gas jar is full.

5. Collect four jars of this gas for experiment 4.12.
   • What is special about the drying of ammonia gas?
   • What property is indicated by the method of collection above?

   Ammonia gas is given off after heating the mixture and it is dried by calcium oxide.

   $$\text{Ca(OH)}_2(s) + 2\text{NH}_4\text{Cl}(s) \rightarrow \text{CaCl}_2(s) + 2\text{H}_2\text{O}(l) + 2\text{NH}_3(g)$$

   Ammonia reacts with the usual drying agents such as anhydrous calcium chloride or concentrated sulphuric acid. With anhydrous calcium chloride it forms a complex solid compound (CaCl$_2$.4NH$_3$).

   Ammonia reacts with concentrated sulphuric acid to form ammonium sulphate.

   $$2\text{NH}_3(g) + \text{H}_2\text{SO}_4(l) \rightarrow (\text{NH}_4)_2\text{SO}_4(aq)$$

   Water vapour is also formed in the reaction. Some of it condenses on the cool part of the flask and delivery tube. If the flask is not kept in sloping position, the water can run back into the hot part making it crack. Ammonia is collected by upward delivery because it is less dense than air.

   Collect several gas jars full of this gas for Experiment 4.12. Heat a mixture of a metal hydroxide such as sodium hydroxide and an ammonium salt. Test the gas given out. Ammonia gas is given out when a metal hydroxide is heated with any ammonium salt.

**4.18: Properties of Ammonia gas**

**Experiment 4.12**

Determination of some of the properties of ammonia

**Apparatus and chemicals**

- four gas jars of ammonia
- wooden splints
- trough
- glass rod
- concentrated hydrochloric acid.
**Procedure**

1. Carry out the following experiments/tests with ammonia collected in Experiment 4.11.
2. Draw Table 4.6 in your notebook and record your observations and conclusions.

**Table 4.6: Properties of ammonia gas**

<table>
<thead>
<tr>
<th>Experiment/Test</th>
<th>Observations</th>
<th>Conclusions</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td>Note the colour and smell of ammonia.</td>
<td></td>
</tr>
<tr>
<td>2.</td>
<td>Dip wet red and blue litmus papers into the first gas jar.</td>
<td></td>
</tr>
<tr>
<td>3.</td>
<td>Put a glowing splint and then a burning splint into the second gas jar.</td>
<td></td>
</tr>
<tr>
<td>4.</td>
<td>Invert the third gas jar in a trough of water. Note the level of the water inside the gas jar after 2 to 3 minutes.</td>
<td></td>
</tr>
<tr>
<td>5.</td>
<td>Dip a glass rod in concentrated hydrochloric acid and bring the dipped end to the mouth of the fourth gas jar Fig. 4.12.</td>
<td></td>
</tr>
</tbody>
</table>

**Fig 4.12: Reaction of ammonia and hydrogen chloride**

**Physical properties of ammonia gas**

Ammonia gas:
- has a pungent choking smell.
- is colourless.
- is very soluble in water. Water rises rapidly in a gas jar full of this gas and fills or almost fills it if inverted in a trough containing water.
Chemical properties of ammonia gas

- **Combustion**
  Ammonia does not burn in air or support combustion.

- **Effect on litmus paper**
  Ammonia turns moist red litmus paper blue. It is the only common alkaline gas.

- **Reaction with hydrogen chloride**
  It reacts with hydrogen chloride gas to form white fumes of ammonium chloride.

**Experiment 4.13**
To demonstrate the high solubility of ammonia in water *(the fountain experiment)*

**Apparatus and chemicals**
- a dry round-bottomed flask
- rubber bang
- glass tubing
- litmus indicator
- dilute sulphuric acid.

**Procedure**
1. Fill a dry round-bottomed flask with dry ammonia and connect the apparatus as shown in Fig. 4.13.

![Fig. 4.13: The fountain experiment](image)

1. Add litmus indicator into the water and stir.
   - What is the colour of the solution now?
2. Add a little dilute sulphuric acid into the solution in the trough.
   - What is the colour of the solution now?
3. Invert the set up in solution as in Fig. 4.13 for sometime to allow some little water to rise up the glass tube.
• What do you observe?
5. Close the tube with your finger and remove the whole flask from the trough and invert it.
6. Remove the finger to allow the little solution that rose into the tube to enter into the flask.
7. Close the tube with the finger again and quickly dip it into the solution as shown in Fig. 4.13. With the tube open, wait for a while.
8. Record your observations and conclusions in your notebook.

When the glass tube is placed into litmus solution, water slowly rises up the tube. When it enters the flask it dissolves so much ammonia that there is partial vacuum in the flask. Water is then forced rapidly up the tube and enters the flask as a fountain.

The litmus indicator turns blue because ammonia solution is alkaline. The fountain experiment shows that ammonia is a very soluble gas. In fact at room temperature 1 cm$^3$ of water dissolves 800 cm$^3$ of ammonia.

**Test for ammonia (Reaction of ammonia with hydrogen chloride)**
Ammonia turns red litmus paper blue. It forms white fumes of ammonium chloride when reacted with hydrogen chloride gas.

$$\text{Ammonia} + \text{hydrogen chloride} \rightarrow \text{ammonium chloride}$$

$$\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s)$$

Why is it not possible to pour ammonia gas into a gas jar of hydrogen chloride gas? Which of these gases would you detect first through smell when released from one location in the laboratory? Give a reason for your answer.

When ammonia gas and hydrogen chloride gas are placed in the corner of a room the smell of ammonia is detected first. This is because ammonia is less dense than air and has a lower density than hydrogen chloride. Ammonia diffuses faster.

**4.19: Aqueous ammonia**
Aqueous ammonia is prepared by dissolving ammonia gas in water

**Experiment 4.14**
Preparation of aqueous ammonia (ammonium hydroxide).

**Apparatus and chemicals**
- beaker
- delivery tube
- filter funnel
- source of ammonia
Procedure
1. From preparation flask pass ammonia gas into water via an inverted filter funnel as shown in Fig. 4.14 for sometime.

![Fig. 4.14: Preparation of aqueous ammonia](image)

The filter funnel should just dip into the water surface.
2. Smell the aqueous solution in the beaker.
3. Record your observations in your notebook. What conclusions can you make from this experiment?

1. Why don’t we dissolve ammonia in the beaker using delivery tube instead of a funnel?
2. Why do we not dry the ammonia gas used in this experiment?

After sometime, the water in the beaker acquires the smell of ammonia gas which has dissolved in it. The solution is known as aqueous ammonia or ammonia solution or ammonium hydroxide.

Dissolution of ammonia in water is a reversible reaction. This is why aqueous ammonia produces ammonia gas at any temperature. It decomposes on warming to produce a strong smelling gas which is ammonia.

\[
\text{Ammonia} + \text{water} \rightleftharpoons \text{ammonium hydroxide}
\]

\[
\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4\text{OH(aq)}
\]

Water rises up the funnel as the ammonia gas dissolves in it. Due to this rise, the water level in the beaker falls and the funnel loses contact with the water surface. Air enters into the funnel. The water in the funnel falls back into the beaker due to force of gravity. So, there is a repeated process of water entering the funnel and falling back into the beaker. Because of this, there is a bubbling noise and no “sucking back” of water can therefore occur.

If a delivery tube is used in place of the funnel, water would “suck back” into the hot preparation flask, thereby cracking it. The funnel also provides a large surface area for absorption of the gas by the water.

In this preparation of aqueous ammonia, it does not make any sense to first dry ammonia gas because we will eventually dissolve the ammonia gas produced in water.
4.20: Properties of aqueous ammonia

Experiment 4.15
Reaction of aqueous ammonia with cations.

Apparatus and chemicals
- test tubes
- droppers
- solutions containing ions of
  - calcium
  - magnesium
  - aluminium
  - zinc
  - lead (II)
  - iron (II)
  - iron (III)
  - copper (II)

Procedure
1. To each solution, add aqueous ammonia dropwise until in excess, shaking the test tube after every drop or two added. Draw Table 4.7 in your notebook and record your observations.

<table>
<thead>
<tr>
<th>Cation in solution</th>
<th>Observations on addition of aqueous ammonia</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Few drops</td>
</tr>
<tr>
<td>Ca^{2+}</td>
<td></td>
</tr>
<tr>
<td>Mg^{2+}</td>
<td></td>
</tr>
<tr>
<td>Al^{3+}</td>
<td></td>
</tr>
<tr>
<td>Zn^{2+}</td>
<td></td>
</tr>
<tr>
<td>Pb^{2+}</td>
<td></td>
</tr>
<tr>
<td>Fe^{3+}</td>
<td></td>
</tr>
<tr>
<td>Fe^{2+}</td>
<td></td>
</tr>
<tr>
<td>Cu^{2+}</td>
<td></td>
</tr>
</tbody>
</table>

The cations in the various solutions react with aqueous ammonia to form insoluble metal hydroxide precipitates.
- Is there any observable change/reaction for calcium cations?
- Which ions form coloured precipitates?
Which precipitates dissolve in excess ammonia solution?
The results of reaction of ammonia solution with cations are as shown in Table 4.8.

Table 4.8: Identifying cations using aqueous ammonia

<table>
<thead>
<tr>
<th>Cation in solution</th>
<th>Observations on addition of aqueous ammonia</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Few drops</td>
</tr>
<tr>
<td>Ca$^{2+}$</td>
<td>white precipitate</td>
</tr>
<tr>
<td>Mg$^{2+}$</td>
<td>white precipitate</td>
</tr>
<tr>
<td>Al$^{3+}$</td>
<td>white precipitate</td>
</tr>
<tr>
<td>Zn$^{2+}$</td>
<td>white precipitate</td>
</tr>
<tr>
<td>Pb$^{2+}$</td>
<td>white precipitate</td>
</tr>
<tr>
<td>Fe$^{3+}$</td>
<td>red-brown precipitate</td>
</tr>
<tr>
<td>Fe$^{2+}$</td>
<td>green precipitate</td>
</tr>
<tr>
<td>Cu$^{2+}$</td>
<td>blue precipitate</td>
</tr>
</tbody>
</table>

Ionic equations representing some of the reactions above are:

\[
\text{Mg}^{2+}(aq) + 2\text{OH}^-(aq) \rightarrow \text{Mg(OH)}_2(s) \quad \text{white precipitate}
\]

\[
\text{Al}^{3+}(aq) + 3\text{OH}^-(aq) \rightarrow \text{Al(OH)}_3(s) \quad \text{white precipitate}
\]

\[
\text{Zn}^{2+}(aq) + 2\text{OH}^-(aq) \rightarrow \text{Zn(OH)}_2(s) \quad \text{white precipitate soluble in excess}
\]

\[
\text{Pb}^{2+}(aq) + 2\text{OH}^-(aq) \rightarrow \text{Pb(OH)}_2(s) \quad \text{white precipitate}
\]

\[
\text{Fe}^{2+}(aq) + 2\text{OH}^-(aq) \rightarrow \text{Fe(OH)}_2(s) \quad \text{green precipitate}
\]
Fe\(^{3+}\)(aq) + 3OH\(^-\)(aq) \rightarrow Fe(OH)\(_3\)(s) 
red-brown precipitate

Cu\(^{2+}\)(aq) + 2OH\(^-\)(aq) \rightarrow Cu(OH)\(_2\)(s) 
blue precipitate soluble in excess forming a deep blue solution.

From the experiments we can conclude that some cations form insoluble precipitates while others form soluble precipitates when reacted with excess aqueous ammonia.

• Write the ionic equation for the reaction between Ca\(^{2+}\) ions and ammonia solution.

4.21: Reactions of ammonia with air/oxygen
Ammonia reacts differently with air or oxygen. The reaction is even more different when a catalyst is introduced.

Experiment 4.16
Reaction of ammonia with air/oxygen without catalyst.

Apparatus and chemicals
• open glass tube
• delivery tube
• glass or cotton wool
• sources of oxygen and ammonia
• wooden splint
• cork/rubber bang

Procedure
1. Arrange the apparatus as shown in Fig. 4.15

Fig. 4.15: Burning ammonia in oxygen

2. Ignite ammonia gas at the jet delivering ammonia before allowing oxygen in.
• What happens?
3. Now adjust the rate of flow of both gases until a flame can be lit at the jet of ammonia tube.

4. Record your observation in your notebook.
   - What is the use of glass or cotton wool?
   - What do you conclude from this experiment?

When a burning splint is inserted in a glass tube containing ammonia gas, it is extinguished. However, ammonia burns with a yellow-brown flame when oxygen is allowed to flow in. Nitrogen and water are formed.

\[
\text{Ammonia} + \text{oxygen} \rightarrow \text{nitrogen} + \text{water} \\
4\text{NH}_3(g) + 3\text{O}_2(g) \rightarrow 2\text{N}_2(g) + 6\text{H}_2\text{O}(l)
\]

The loosely packed glass or cotton wool helps to spread out the oxygen. Ammonia gas neither burns, nor supports burning. It can only burn in oxygen or air enriched with oxygen.

**Experiment 4.17**

Catalytic oxidation of ammonia.

**Apparatus and chemicals**
- platinum coil
- delivery tube
- beaker
- concentrated ammonia solution
- source of oxygen
- litmus papers (blue and red)

**Caution:** The reaction can sometimes be explosive.

**Procedure**
1. Bubble oxygen through the concentrated ammonia solution for a few minutes.
2. Put a red-hot platinum coil into the mixture of oxygen and ammonia solution (concentrated ammonia solution produces ammonia gas) as shown in Fig. 4.16.
3. Record your observations in your notebook.
   - What is the colour of the fumes that are seen coming out of the solution?
   - What is the effect of these fumes on moist blue litmus paper?
   - What is the effect of the resulting solution on blue litmus paper?
   - What conclusions can you make from your observations?

A reaction takes place when oxygen is bubbled through ammonia solution, and a red-hot platinum coil lowered into the gas mixture. The fumes produced turn moist blue litmus paper red. The solution formed in the beaker turns blue litmus paper red indicating that it is acidic.

The main reactions that take place are shown below:

\[
\begin{align*}
\text{Ammonia} + \text{oxygen} & \rightarrow \text{nitrogen(II) oxide} + \text{water} \\
4\text{NH}_3(g) + 5\text{O}_2(g) & \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g) \\
\text{Nitrogen(II) oxide} + \text{oxygen} & \rightarrow \text{nitrogen(IV) oxide} \\
2\text{NO}(g) + \text{O}_2(g) & \rightarrow 2\text{NO}_2(g) \\
& \text{(Red-brown fumes)} \\
\text{Nitrogen(IV) oxide} + \text{water} & \rightarrow \text{nitric acid} + \text{nitrous acid} \\
2\text{NO}_2(g) + \text{H}_2\text{O}(l) & \rightarrow \text{HNO}_3(aq) + \text{HNO}_2(aq) \\
\text{Nitrous acid} + \text{oxygen} & \rightarrow \text{nitric acid} \\
2\text{HNO}_2(aq) + \text{O}_2(g) & \rightarrow 2\text{HNO}_3(aq)
\end{align*}
\]

The solution formed is nitric acid. We can conclude that it is possible to convert ammonia to nitric acid using platinum as a catalyst.

4.22 Reaction of ammonia with copper(II) oxide

Experiment 4.18
Reaction of ammonia with copper(II) oxide

Apparatus and chemicals
- delivery tube
- source of dry ammonia
- copper(II) oxide
- beaker
- U-tube
- ice-cold water
- anhydrous copper(II) sulphate or cobalt(II) chloride paper
- gas-jar
- combustion tube
- source of heat
Procedure

1. Set up the apparatus as shown in Fig. 4.17.

   ![Diagram of the apparatus](Diagram.png)

   **Fig 4.17: Reaction of ammonia and copper(II) oxide**

2. Heat strongly copper(II) oxide in the combustion tube.
3. Pass dry ammonia over the hot oxide.
4. Record observation in your notebook.
5. Test the gas collected in the gas jar for colourless and odourless gases like:
   - oxygen
   - hydrogen
   - carbon(IV) oxide
6. Record your results in your notebook.

1. What is the colour of the residue in the combustion tube?
2. Test the liquid in the U-tube with anhydrous copper(II) sulphate or a dry cobalt(II) chloride paper.
3. What is your observation?
4. What conclusions do you draw from this experiment?

When copper(II) oxide is heated strongly and dry ammonia passed over it, the black oxide is reduced to red-brown copper. Ammonia is oxidised by the copper(II) oxide to a colourless liquid which collects in the U-tube.

\[
3\text{CuO(s)} + 2\text{NH}_3(g) \rightarrow 3\text{Cu(s)} + 3\text{H}_2\text{O(l)} + \text{N}_2(g)
\]

Copper(II) oxide is the **oxidizing agent** and ammonia is the **reducing agent**.
The colourless liquid turns white anhydrous copper(II) sulphate blue. It also turns blue cobalt(II) chloride paper pink. If this liquid is boiled and its boiling point found to be 100˚C at 760mm/Hg pressure, this liquid is therefore pure water.

The tests show that the gas in the jar is nitrogen because it:
- does not re-light a glowing splint.
- does not produce a “pop” sound when a burning splint is introduced.
- does not burn.
- has no effect on aqueous calcium hydroxide.

What happens to the unreacted ammonia gas?
Give another reaction where ammonia acts as a reducing agent.
The unreacted ammonia dissolves in the water in the trough. When ammonia gas burns in oxygen, it acts as a reducing agent.

4.23 Manufacture of ammonia by Haber process

Raw materials
The raw materials required for the manufacture of ammonia by Haber process are nitrogen and hydrogen. Nitrogen is obtained from air by fractional distillation of liquid air. Hydrogen is usually obtained from natural gas (methane) or crude oil. Hydrogen can also be obtained by electrolysis of acidified water or brine where electricity is cheap and readily available.

Optimum conditions needed during manufacture of ammonia
- A mixture of nitrogen and hydrogen in the ratio of 1:3 by volume.
- A pressure of about 200–250 atmospheres.
- A temperature of about 400˚C – 500˚C.
- Iron catalyst in finely divided form and impregnated with aluminium oxide as a promoter of the catalyst

Process
A mixture of nitrogen and hydrogen in the ratio of 1:3 by volume is reacted to form ammonia.

High pressure gives a high yield, but the higher the pressure, the greater the cost of equipment to produce and maintain the pressure. Most industries use a pressure of about 500 atmospheres.

Low temperatures are required to give the highest yield of ammonia. However at a low temperature, the rate of reaction is slow. Therefore a compromise has to be reached between a reasonable rate of reaction and a reasonable yield of ammonia. A temperature of 400˚ – 500˚C is used.

To increase the rate of reaction, iron is used as a catalyst. The catalyst, is mixed with aluminium oxide to increase its efficiency. A substance such as aluminium oxide which increases the efficiency of a catalyst is called a promoter.
Reactions taking place
When nitrogen and hydrogen react under the conditions mentioned, only about 10% of ammonia is formed. The reaction is exothermic.

\[
\text{Nitrogen} + \text{hydrogen} \rightarrow \text{ammonia} + \text{heat} \\
N_2(g) + 3H_2(g) \rightarrow 2\text{NH}_3(g) + \text{heat}
\]

The heat produced during the reaction is taken back to the heat exchanger together with the unreacted nitrogen, hydrogen and the 10% ammonia formed. This helps to maintain the temperature at 500°C needed for the reaction. The ammonia produced and unreacted hydrogen and nitrogen are then taken to the liquifier or condenser where ammonia is liquified or dissolved in water. The unreacted nitrogen and hydrogen are re-circulated to the compressor. Study Fig. 4.18 carefully.

![Diagram of the ammonia manufacturing process](image)

**Fig. 4.18: Manufacture of ammonia**

**Note:**
Nitrogen and hydrogen must be dried and purified to remove dust particles, which may poison the catalyst. Argon which enters the plant with the nitrogen is also removed.

**4.24: Uses of ammonia**
Some of the uses of ammonia are illustrated in Fig. 4.19(a-e).

1. Large quantities of ammonia gas are used to make fertilizers (Fig. 4.19a).
2. Ammonia gas is used in the manufacture of nitric acid. (Fig. 4.19b)
3. Liquid ammonia is used as a refrigerant in large scale refrigerating plants and factories (Fig. 4.19c).
4. Ammonia solution is used as a solvent in cleaning (in laundries).
5. Ammonia is used in the manufacture of ammonium salts such as ammonium chloride (Fig. 4.19e) which is used in dry cells and ammonium carbonate used in smelling salts. Smelling salts produce ammonia gas slowly at room temperature.
This gas acts on the heart and prevents fainting and dizziness.

6. Ammonia is used in the manufacture of dyes, wood pulp, plastics and fibres such as nylon.

4.25: Nitrogen-containing fertilisers

Nitrogen-containing fertilisers are produced from ammonia gas. Ammonia is reacted with acids like sulphuric acid to form ammonium sulphate.

\[
\text{Ammonia} + \text{sulphuric acid} \rightarrow \text{ammonium sulphate} \\
2\text{NH}_3(g) + \text{H}_2\text{SO}_4(aq) \rightarrow (\text{NH}_4)_2\text{SO}_4(aq)
\]

Nitric acid forms ammonium nitrate.

\[
\text{Ammonia} + \text{nitric acid} \rightarrow \text{ammonium nitrate} \\
\text{NH}_3(g) + \text{HNO}_3(aq) \rightarrow \text{NH}_4\text{NO}_3(aq)
\]

Phosphoric(V) acid forms ammonium phosphate.

\[
\text{Ammonia} + \text{phosphoric(V) acid} \rightarrow \text{ammonium phosphate} \\
3\text{NH}_3(g) + \text{H}_3\text{PO}_4(aq) \rightarrow (\text{NH}_4)_3\text{PO}_4(aq)
\]

Nitrogen containing fertilisers are soluble and hence are leached out. They are generally hygroscopic and should therefore be stored under dry conditions.
We have various types of nitrogenous fertilisers but the most important are discussed below.

1. **Ammonium sulphate, (NH$_4$)$_2$SO$_4$**

Ammonium sulphate contains 20-21% of nitrogen. It is an acidic fertiliser. Because of its acidic nature, it should not be applied to acidic soils unless specifically advised by an agricultural chemist for specific crops. Ammonium sulphate is good for soils which are deficient in sulphur content.

To calculate the percentage of nitrogen in fertilisers, we first add the relative atomic masses of each element in the compound.

(N=14, O = 16, H = 1, S = 32)

\[
\begin{align*}
\text{Ammonium sulphate} & : (NH_4)_2SO_4 \\
2 \times N & = 2 \times 14 = 28 \\
8 \times H & = 8 \times 1 = 8 \\
1 \times S & = 1 \times 32 = 32 \\
4 \times O & = 4 \times 16 = 64 \\
\text{RMM} & = 132
\end{align*}
\]

The R.M.M of ammonium sulphate is 132.

\[
\text{Fraction of N in (NH}_4\text{)}_2\text{SO}_4 = \frac{28}{132}
\]

\[
\% \text{ N} = \frac{28}{132} \times 100 = 21.2 \%
\]

**Exercise 4.2**

(a) Calculate the percentage of nitrogen in the following compounds.

(i) \((NH_2)_2CO\)

(ii) \((NH_4)_3PO_4\)

(b) Calculate the number of grams of nitrogen in 100 grams of \((NH_2)_2CO\)

\((N = 14, H = 1, C = 12, O = 16, P = 32)\)

2. **Ammonium sulphate nitrate**

\((NH_4)NO_3 \cdot (NH_4)_2SO_4\)

Ammonium sulphate nitrate is composed of ammonium sulphate and ammonium nitrate.

Ammonium sulphate nitrate

\((NH_4)_2SO_4 + NH_4NO_3\)
The R.M.M of ammonium sulphate nitrate is 212.

\[
\text{Fraction of } N = \frac{56}{212} \\
\text{% of } N = \frac{56}{212} \times 100 \\
= 26.4 \% 
\]

Ammonium sulphate nitrate is less acidic than ammonium sulphate. The ammonium nitrate supplies nitrate ions while ammonium sulphate supplies ammonium ions to the soils. Plants absorbs nitrate ions directly but not ammonium ions. Therefore ammonium nitrate supplies nitrogen to the plants quickly while ammonium sulphate supplies it slowly. So ammonium sulphate nitrate fertiliser is both quick and slow acting.

3. **Calcium ammonium nitrate** \((\text{NH}_4\text{NO}_3,\text{CaCO}_3)\)

Calcium ammonium nitrate is a mixture of ammonium nitrate and calcium carbonate. It is hygroscopic and gives out an odour of ammonia gas. It is a neutral fertiliser. What is the percentage of nitrogen in calcium ammonium nitrate? Use similar approach as explained above.

4. **Urea, \((\text{NH}_2)_2\text{CO}\)**

Urea has the highest nitrogen content of all nitrogenous fertilisers. It is very soluble and therefore easily leached out. It is acidic and very volatile.

- Calculate the percentage of nitrogen in urea.

5. **Diammonium Phosphate** \((\text{NH}_4)_2\text{HPO}_4\)

Diammonium phosphate supplies nitrogen and phosphorus, two of the three major nutrients needed by crops. The nitrogen is in ammonium ion \((\text{NH}_4^+)\) and therefore this fertiliser is slow acting. It is an acidic fertiliser because it produces phosphoric (V) acid, \((\text{H}_3\text{PO}_4)\).

- What is the percentage of nitrogen in diammonium phosphate?

The percentage of nitrogen in calcium ammonium nitrate, ammonium nitrate, urea and diammonium phosphate are summarised in the table below.
Table 4.9: Summary of percentage of nitrogen in some fertilisers

<table>
<thead>
<tr>
<th>Calcium ammonium nitrate</th>
<th>Ammonium nitrate</th>
<th>Urea</th>
<th>Diammonium phosphate</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₄NO₃.CaCO₃</td>
<td>NH₄NO₃</td>
<td>(NH₂)₂CO</td>
<td>(NH₄)₂HPO₄</td>
</tr>
<tr>
<td>N × 2 = 14 × 2 = 28</td>
<td>N × 2 = 14 × 2 = 28</td>
<td>C × 1 = 12 × 1 = 12</td>
<td>N × 2 = 14 × 2 = 28</td>
</tr>
<tr>
<td>H × 4 = 1 × 4 = 4</td>
<td>H × 4 = 1 × 4 = 4</td>
<td>O × 1 = 16 × 1 = 16</td>
<td>H × 9 = 1 × 9 = 9</td>
</tr>
<tr>
<td>O × 6 = 16 × 6 = 96</td>
<td>O × 3 = 16 × 3 = 48</td>
<td>N × 2 = 14 × 2 = 28</td>
<td>P × 1 = 31 × 1 = 31</td>
</tr>
<tr>
<td>Ca × 1 = 40 × 1 = 40</td>
<td>RMM = 80</td>
<td>H × 4 = 1 × 4 = 4</td>
<td>O × 4 = 16 × 4 = 64</td>
</tr>
<tr>
<td>C × 1 = 12 × 1 = 12</td>
<td>R.M.M = 180</td>
<td>R.M.M = 60</td>
<td>R.M.M = 132</td>
</tr>
<tr>
<td>% of N</td>
<td>% of N</td>
<td>% of N</td>
<td>% of N</td>
</tr>
<tr>
<td>= 28 × 100/180</td>
<td>= 28 × 100/80</td>
<td>= 28 × 100/60</td>
<td>= 28 × 100/132</td>
</tr>
<tr>
<td>= 15.5%</td>
<td>= 35%</td>
<td>= 46.7%</td>
<td>= 21.2%</td>
</tr>
</tbody>
</table>

Exercises 4.3
1. Study the flow chart and answer the questions that follow:

   (i) Identify substances D and E.
   (ii) Give the optimum conditions for the reaction in chamber 1.
   (iii) Name a catalyst and promoter needed in chamber 1.
   (iv) Write an equation for the reaction in chamber 2.

2.

   (a) (i) Name liquid L.
       (ii) Name gas X.
       (b) (i) State the observation made in step 1.
(ii) Write a chemical equation for the reaction between ammonia and copper(II) oxide.

(iii) What is the colour of litmus paper in solution L? Explain.

3. Ammonia gas can be reacted with a certain reagent to produce ammonium sulphate which is a fertiliser.
   (i) Write an equation to show how this fertiliser is formed.
   (ii) Calculate the percentage by mass of nitrogen in the fertiliser.
        \[ N = 14, \quad O = 16, \quad H = 1, \quad S = 32. \]
   (iii) Both ammonium sulphate and sodium nitrate are used as fertilisers. Which of the two is fast-acting? Explain your answer.

4. (i) There are four main reactions that take place during the catalytic oxidation of concentrated ammonia solution. Write the four chemical equation for the main reactions.
   (ii) What three conclusions can be made about ammonia when it reacts with copper(II) oxide to form water, nitrogen and copper?

4.26: Nitric acid
Nitric acid is a familiar mineral acid in the school laboratory. It is an acid which has been known for many years by scientists. Nitric acid is a powerful oxidising agent.

4.27: Laboratory preparation of nitric acid from a nitrate and concentrated sulphuric acid

Experiment 4.19
Laboratory preparation of nitric acid.

Apparatus and chemicals
- retort
- round or flat-bottomed flask
- concentrated sulphuric acid
- potassium nitrate

Procedure
1. Arrange the apparatus as shown in Fig. 4.20
2. Put some potassium nitrate in a retort. Add concentrated sulphuric acid to cover it.
3. Open the tap water to cool the retort.
4. Heat the retort gently. What do you observe in the retort and the flask? Record your observations.
5. Keep the acid formed for Experiment 4.20.
   • Why are all the apparatus used in this preparation made of glass?
   • What do you conclude from this experiment?

When concentrated sulphuric acid is added to potassium nitrate then heated, the potassium nitrate gradually dissolves and effervescence occurs.

Nitric acid is formed as a result and it distils because it is more volatile than concentrated sulphuric acid. Its vapours condense to form a yellow liquid.

\[
\text{Potassium nitrate} + \text{sulphuric acid} \rightarrow \text{potassium hydrogensulphate} + \text{nitric acid}
\]

\[
\text{KNO}_3(s) + \text{H}_2\text{SO}_4(l) \rightarrow \text{KHSO}_4(aq) + \text{HNO}_3(aq)
\]

The red-brown vapours seen in the retort flask are as a result of the decomposition of some of the nitric acid by heat, forming red-brown nitrogen(IV) oxide.

\[
\text{Nitric acid} \xrightarrow{\text{heat}} \text{water} + \text{nitrogen(IV)} + \text{oxygen oxide}
\]

\[
4\text{HNO}_3(aq) \xrightarrow{\text{heat}} 2\text{H}_2\text{O}(l) + 4\text{NO}_2(g) + \text{O}_2(g)
\]
The nitric acid collected is yellow because it contains some dissolved nitrogen(IV) oxide, otherwise, pure nitric acid is colourless. The yellow colour can be removed by bubbling air through the acid.

A retort is used for the preparation of nitric acid because it is made entirely of glass. Hot nitric acid vapours attack cork or rubber stoppers and tubings. Generally, heating nitrates with concentrated sulphuric acid produces nitric acid. However, potassium nitrate is preferred in preparing nitric acid in the laboratory because it has no water of crystallisation which can dilute the acid produced. Sodium nitrate can also be used but it is hygroscopic.

\[
\text{Sodium nitrate} + \text{sulphuric acid} \rightarrow \text{sodium hydrosulphate} + \text{nitric acid}
\]

\[
\text{NaNO}_3(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{NaHSO}_4(aq) + \text{HNO}_3(aq)
\]

NB: Nitric acid is kept in dark bottles. This is because light decomposes the acid into the water, nitrogen(IV) oxide and oxygen.

4.28: Manufacture of nitric acid by Ostwald Process

Raw materials
Nitric acid is manufactured by the catalytic oxidation of ammonia. Ammonia is obtained through the Haber process. The catalyst used is platinum - rhodium gauze. The raw materials required are ammonia, air and water.

Optimum conditions needed:
- Excess air
- Ammonia
- A temperature of 850°C-900°C
- Platinum - rhodium catalyst

The air/ammonia must be free of dust and other impurities to avoid poisoning the catalyst.

Reactions taking place
The first reaction in this process produces nitrogen(II) oxide and water. The heat produced in this step maintains the catalyst temperature and therefore no further heating is required.

\[
\text{Ammonia} + \text{oxygen} \rightarrow \text{nitrogen(II) oxide} + \text{water}
\]

\[
4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g)
\]

Nitrogen(II) oxide and steam are cooled by air from the compressor. Nitrogen(II) oxide is oxidised further by the air to produce nitrogen(IV) oxide.
Nitrogen(II) oxide + oxygen $\rightarrow$ nitrogen(IV) oxide

$2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$

Nitrogen(IV) oxide is then passed through water to form nitric acid ($\text{HNO}_3$) and nitrous acid ($\text{HNO}_2$).

Nitrogen(IV) oxide + water $\rightarrow$ nitrous acid + nitric acid.

$2\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HNO}_2(aq) + \text{HNO}_3(aq)$

Excess air is passed to oxidise nitrous acid to nitric acid.

Nitrous acid + oxygen $\rightarrow$ nitric acid

$2\text{HNO}_2(aq) + \text{O}_2(g) \rightarrow 2\text{HNO}_3(aq)$

Study Fig. 4.21 below.

**Fig. 4.21: Manufacture of nitric acid**

About 65 per cent nitric acid and 35 per cent water are produced from a single converter. The unreacted air, ammonia and nitrogen(IV) oxide are re-circulated back to the compressor. The concentration of the nitric acid produced can be increased by distillation.

**4.29: Properties of dilute nitric acid**

**Chemical properties**

Dilute nitric acid ionises in water to produce hydrogen ions and therefore acts as an acid. Concentrated acid readily decomposes on slight heating to nitrogen(IV) oxide, water and oxygen. The oxygen formed makes the acid a powerful oxidising agent.
Experiment 4.20
To determine properties of dilute nitric acid.

Apparatus and chemicals
- test tubes
- wooden splint
- sodium carbonate
- sodium hydroxide
- copper(II) oxide
- magnesium ribbon
- dilute nitric acid

Procedure
1. Carry out the following experiments/tests with dilute acid collected in Experiment 4.19.
2. Draw Table 4.10 in your notebook and record your observations and conclusions.

Table 4.10: Properties of dilute nitric acid

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Observations</th>
<th>Conclusions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Add dilute nitric acid to sodium carbonate in a test tube. Test the gas evolved.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Add phenolphthalein indicator to a solution of sodium hydroxide in a test tube. Then add dilute nitric acid drop by drop.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Add dilute nitric acid to copper(II) oxide in a test tube.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Put magnesium ribbon in a test tube. Add dilute nitric acid. Repeat the same experiment with very dilute cold nitric acid.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Put zinc in a test tube. Add dilute nitric acid.</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Dilute nitric acid reacts as an acid in its action on carbonates, hydroxides and oxides.

**Examples**

(i) \[ \text{Metal} + \text{dilute nitric acid} \rightarrow \text{metal nitrate} + \text{water} + \text{carbon(IV) oxide} \]

Calcium + nitric acid \[\rightarrow\] calcium + water + carbon(IV) oxide

\[ \text{CaCO}_3(s) + 2\text{HNO}_3(aq) \rightarrow \text{Ca(NO}_3)_2(aq) + \text{H}_2\text{O(l)} + \text{CO}_2(g) \]

(ii) Metal hydroxide + dilute nitric acid \[\rightarrow\] metal nitrate + water

Sodium hydroxide + nitric acid \[\rightarrow\] sodium nitrate + water

\[ \text{NaOH(aq)} + \text{HNO}_3(aq) \rightarrow \text{NaNO}_3(aq) + \text{H}_2\text{O(l)} \]

(iii) Metal oxide + dilute nitric acid \[\rightarrow\] metal nitrate + water

Copper(II) oxide + nitric acid \[\rightarrow\] copper(II) nitrate + water

\[ \text{CuO(s)} + 2\text{HNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{H}_2\text{O(l)} \]

Because of its oxidising properties (produces oxygen very easily), nitric acid does not usually give off hydrogen with metals. The hydrogen produced is immediately oxidised to form water.

For example, zinc metal reacts with dilute nitric acid to form zinc nitrate, nitrogen (IV) oxide and water.

\[ \text{Zinc} + \text{nitric acid} \rightarrow \text{zinc nitrate} + \text{nitrogen(IV) oxide} + \text{water} \]

\[ \text{Zn(s)} + 4\text{HNO}_3(aq) \rightarrow \text{Zn(NO}_3)_2(aq) + 2\text{NO}_2(aq) + 2\text{H}_2\text{O} \]

An exception is the reaction of magnesium with very dilute and cold nitric acid.

\[ \text{Magnesium} + \text{nitric acid} \rightarrow \text{magnesium nitrate} + \text{hydrogen} \]

\[ \text{Mg(s)} + 2\text{HNO}_3(aq) \rightarrow \text{Mg(NO}_3)_2(aq) + \text{H}_2(g) \]

**4.30: Properties of concentrated nitric acid**

Concentrated nitric acid is a very powerful oxidising agent. Its oxidising power depends on the concentration, temperature and the type of reactions.

**Experiment 4.21**

Reaction of concentrated nitric acid with iron(II) sulphate, copper and sulphur.
Apparatus and chemicals
- iron(II) sulphate solution
- dilute sulphuric acid
- sulphur powder
- concentrated nitric acid
- barium chloride or barium nitrate solution
- lead(II) nitrate solution
- copper turnings

Procedure
(a) Iron(II) Sulphate
1. Add a few drops of dilute sulphuric acid to about 2 cm$^3$ of a solution of iron(II) sulphate in a boiling tube,
2. Add a few drops of concentrated nitric acid to the solution formed in 1.
   - Record your observations in your notebook.
   - What is your conclusion?

Green iron(II) sulphate solution acidified with dilute sulphuric acid reacts with concentrated nitric acid to give a yellow-brown solution of iron(III) sulphate.

\[
\text{Iron(II) sulphate + sulphuric acid + nitric acid} \rightarrow \text{iron(III) sulphate + water + nitrogen(II) oxide}
\]

\[
6\text{FeSO}_4(aq) + 3\text{H}_2\text{SO}_4(aq) + 2\text{HNO}_3(aq) \rightarrow 3\text{Fe}_2(\text{SO}_4)_3(aq) + 4\text{H}_2\text{O}(l) + 2\text{NO}(g)
\]
(yellow-brown)

This is a good example of oxidation of iron(II) ion to iron(III) ion reaction.

\[
\text{Fe}^{2+}(aq) \rightarrow \text{Fe}^{3+}(aq) + e^- \text{ (oxidation)}
\]

Can the same type of reaction take place with iron(II) chloride?
The same type of reaction can take place with iron(II) chloride solution (addition of sulphuric acid is not necessary).

The nitrogen(II) oxide produced is oxidised to red-brown nitrogen(IV) oxide gas by oxygen from the air. i.e

\[
\text{Nitrogen(II) oxide + oxygen} \rightarrow \text{nitrogen(IV) oxide}\n\]

\[
2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)
\]

(b) Copper
1. Put a few copper turnings in a boiling tube and then add 50% concentrated nitric acid.
2. Repeat the procedure with concentrated nitric acid.
3. Record your observations in your note book.
   • What do you observe in the boiling-tube?
   • What is the colour of the gas produced?
   • What is your conclusion?
   • Predict what happens if zinc is used instead of copper.

When copper reacts with 50% concentrated nitric acid, copper(II) nitrate, water and nitrogen(II) oxide are formed.

\[
\text{Copper} + 50\% \text{ conc. nitric acid} \rightarrow \text{copper(II) nitrate} + \text{water} + \text{nitrogen(II) oxide}
\]
\[
3\text{Cu}(s) + 8\text{HNO}_3(\text{aq}) \rightarrow 3\text{Cu(NO}_3)_2(\text{aq}) + 4\text{H}_2\text{O(}l\text{)} + 2\text{NO(g)}
\]

Nitrogen(II) oxide is oxidised to nitrogen(IV) oxide by oxygen from the air.

Concentrated nitric acid reacts with copper to form copper(II) nitrate, water and nitrogen(IV) oxide.

\[
\text{Copper} + \text{nitric acid} \rightarrow \text{copper(II) nitrate} + \text{water} + \text{nitrogen(IV) oxide}
\]
\[
\text{Cu}(s) + 4\text{HNO}_3(\text{aq}) \rightarrow \text{Cu(NO}_3)_2(\text{aq}) + 2\text{H}_2\text{O(l)} + 2\text{NO}_2(\text{g})
\]

The same type of reaction takes place when concentrated nitric acid reacts with zinc.

\[
\text{Zinc} + \text{nitric acid} \rightarrow \text{zinc nitrate} + \text{water} + \text{nitrogen(IV) oxide}
\]
\[
\text{Zn}(s) + 4\text{HNO}_3(\text{aq}) \rightarrow \text{Zn(NO}_3)_2(\text{aq}) + 2\text{H}_2\text{O(l)} + 2\text{NO}_2(\text{g})
\]

(c) Sulphur
1. Add about 2 cm$^3$ of concentrated nitric acid to sulphur powder in a boiling tube and warm it.
2. Filter the mixture.
3. Dilute the filtrate with some water and divide it into two portions in two test tubes.
   (i) To the first test-tube, add a few drops of barium chloride or barium nitrate solution.
      • Record your observations in your notebook.
   (ii) To the second test-tube, add a few drops of lead(II) nitrate solution.
      • Record your observations in your notebook.

What is your conclusion?
   • A student added dilute nitric acid to the second test tube before drawing his/her conclusion. What doubt was the student trying to eliminate?

Concentrated nitric acid reacts with sulphur to form sulphuric acid, water and nitrogen(IV) oxide.
Sulphur + nitric acid → sulphuric acid + water + nitrogen(IV) oxide
S(s) + 6HNO₃(aq) → H₂SO₄(aq) + 2H₂O(l) + 6NO₂(g)

When barium chloride or barium nitrate solution is added to the sulphuric acid produced in the reaction, a white precipitate of barium sulphate is formed. The sulphuric acid also forms a white precipitate of lead(II) sulphate when reacted with lead (II) nitrate solution. This is the test for SO₄²⁻ ions. However, if sulphite (SO₃²⁻) ions were present, the precipitate would have dissolved in dilute nitric acid.

**Ionic equations for the reactions are:**

\[ \text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s) \]
\[ \text{Pb}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{PbSO}_4(s) \]

**Note:** In all these reactions, nitric acid is reduced to either nitrogen (II) oxide or nitrogen(IV) oxide and water. This shows that nitric acid is an oxidising agent.

**Exercise 4.3**
1. State four physical properties of nitrogen(IV) oxide.
2. Arrange the nitrogen oxides in order of their increasing stability.
3. Which two acids are formed when nitrogen(IV) oxide reacts with water?
4. Which two salts are formed when nitrogen(IV) oxide reacts with sodium hydroxide solution?
5. Explain the formation of the two salts in question 4 above.

**4.31 Uses of nitric acid**
1. Manufacture of nitrate fertilisers (Fig. 4.22a)
2. Manufacture of explosives e.g. TNT and dynamite (Fig. 4.22b)
3. Manufacture of dyes (Fig. 4.22c)
4. Manufacture of plastics (Fig. 4.22d)
5. It is used as an oxidising agent in textile industries (Fig. 4.22e)
6. It is used in the refining of gemstones (Fig. 4.22f)
7. It is used in the identification of gold. Copper, bronze and copper iron disulphide (CuFeS₂) react with concentrated nitric acid. The acid does not react with gold, hence it can be used to distinguish gold from those other substances which have similar appearance. (Fig. 4.22g).
Fig. 4.22: Uses of nitric acid

4.32 Action of heat on nitrates

Experiment 4.22
Action of heat on metal nitrate.

Apparatus and chemicals
- 5 test tubes
- blue litmus paper
- wooden splint
- sodium nitrate
• potassium nitrate
• copper(II) nitrate
• lead(II) nitrate
• silver nitrate (if available)

Procedure
1. Heat a spatula full of sodium nitrate, potassium nitrate, copper(II) nitrate, lead(II) nitrate and silver nitrate in different test tubes.
2. Test any gas evolved using a glowing splint and moist blue litmus paper.
3. Observe the residue both when hot and cold.
   • What is the colour of the gas or the gas mixture?
   • What is the colour of the hot and cold solid residues?
   • Record your observations in your notebook.
   • What do you conclude about nitrates?
Check whether your results agree with those in Table 4.11.

Table 4.11: Action of heat on metal nitrates

<table>
<thead>
<tr>
<th>Metal nitrate</th>
<th>Observations and products formed</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium nitrate</td>
<td>White residue, glowing splint re-lights. Sodium nitrite and oxygen are produced.</td>
</tr>
<tr>
<td>Potassium nitrate</td>
<td>Glowing splint re-lights. Potassium nitrite and oxygen are produced.</td>
</tr>
<tr>
<td>Copper(II) nitrate</td>
<td>Glowing splint re-lights. A black residue of copper(II) oxide is produced. Nitrogen(IV) oxide and oxygen are evolved.</td>
</tr>
<tr>
<td>Lead(II) nitrate</td>
<td>Glowing splint re-lights. A hot red-brown residue of lead(II) oxide (yellow on cooling), nitrogen(IV) oxide and oxygen are produced.</td>
</tr>
<tr>
<td>Silver nitrate</td>
<td>Glowing splint re-lights. Pure solid silver, nitrogen(IV) oxide and oxygen gases are produced.</td>
</tr>
</tbody>
</table>

Note:
Lead(II) oxide formed from the decomposition of lead(II) nitrate fuses with the test tube, making it difficult to clean such a test tube.
• Predict the products formed when nitrates of calcium, magnesium and zinc are heated. Write the chemical equations.
Lead(II) nitrate produces a crackling sound when heated. Hydrated copper(II) nitrate decomposes in a similar way but it does not produce a crackling sound. When heated, copper(II) nitrate melts first and dissolves in its water of crystallisation forming a green solution. Decomposition of this nitrate only starts when the water in the solution evaporates.

Nitrates of calcium, magnesium, and zinc give their corresponding metal oxides, nitrogen (IV) oxide and oxygen on heating. Nitrate of mercury(II) gives pure mercury, nitrogen(IV) oxide and oxygen on heating. Mercury fumes are poisonous, do not heat.

We have seen that all nitrates decompose on heating to form different products. The ease with which metal nitrates decompose increases down the reactivity series of the metals. Let us now find out how we can test for a nitrate ion.

4.33: Tests for nitrates
The nitrates should be in solution form before the tests are carried out. Remember, all nitrates are soluble in water.

Experiment 4.23
Test for nitrates

Apparatus and chemicals
- test tubes
- metal nitrates
- freshly prepared iron(II) sulphate solution
- concentrated sulphuric acid

Procedure
1. In a test-tube, mix equal volumes of a suspected nitrate solution with an equal volume of freshly prepared iron(II) sulphate solution.
2. With the test-tube tilted as shown in Fig. 4.23, carefully pour concentrated sulphuric acid down its side, so that two layers of solutions are formed.
   - Why should we use freshly prepared iron(II) sulphate solution?
   - Explain what happened when you shook one of the test tubes with a brown ring?

![Fig. 4.23: “Brown ring” test for soluble nitrates](image-url)
3. Record observations in your notebook.
   • What do you observe where the two layers meet?
4. Repeat the procedure with the other nitrate solutions and record your observations.
5. Shake one of test tube. Why should we use freshly prepared iron(II) sulphate solution which has a “brown ring”? Record your observation.

Concentrated sulphuric acid is denser than the layer containing the suspected nitrate and iron(II) sulphate. If the solution is a nitrate, a “brown ring” forms between the two layers of solutions.

Concentrated sulphuric acid reacts with any nitrate salt to produce nitric acid. The nitric acid is reduced to nitrogen(II) oxide by iron(II) sulphate. The nitrogen(II) oxide formed reacts with unreacted iron(II) sulphate to form a brown nitroso-iron(II) sulphate (FeSO₄.NO) which appears as a brown ring.

\[
\text{Iron(II) sulphate } + \text{ nitrogen(II) oxide } \rightarrow \text{ nitroso-iron(II) sulphate}
\]

\[
\text{FeSO}_4 \ (aq) + \text{ NO(g)} \rightarrow \text{ FeSO}_4 \cdot \text{NO(aq)}
\]

The ring disappears if the solution is shaken. This is because concentrated sulphuric acid mixes with the aqueous solution producing a lot of heat. This heat decomposes nitroso-iron(II) sulphate as shown.

\[
\text{Nitroso-iron(II) sulphate } \rightarrow \text{ iron(II) sulphate } + \text{ nitrogen(II) oxide}
\]

\[
\text{FeSO}_4 \cdot \text{NO(aq)} \rightarrow \text{ FeSO}_4 \ (aq) + \text{ NO(g)}
\]

Iron(II) sulphate solution is oxidised by oxygen from the air to iron(III) sulphate within a short period. This is the reason why it should always be freshly prepared. All nitrate salts are soluble in water. They all form nitrate ions when in solution which gives the “brown ring”.

Can dilute or concentrated nitric acid give a positive “brown ring” test?
It will give because it forms nitrate ions when in aqueous form.

4.34 Pollution effects of nitrogen compounds on the environment

(a) Oxides of nitrogen
Nitrogen(II) oxide combines with oxygen in the atmosphere to form nitrogen(IV) oxide. Nitrogen(IV) oxide, is also found in the atmosphere as a result of industrial processes or from automobile exhausts. Nitrogen(IV) oxide is toxic and cause respiratory diseases. Fig. 4.24.
Nitrogen(IV) oxide also reacts with water vapour in the air and/or rain to form dilute nitric acid. Rain containing this acid or other acids is called “acid rain”. Acid rain kills animals and plants on land and in water bodies. This is because the acid lowers the pH of soil and water making it difficult for some organisms to survive. Acid rain also attacks most metallic objects, e.g damages iron roofs by corroding them. This causes leaking in roofs.

What is formed when “acid rain” containing nitric acid reacts with rust-free iron sheets?

(b) Nitrate fertilisers
The continued use of large quantities of nitrate fertilisers pollutes the environment. This is because excess nitrates in the soil are washed down into rivers, lakes and other water bodies.

   In water, nitrates cause aquatic plants and animals (e.g. plankton and algae) to grow in large quantities. The plankton and algae support higher plants and fish. Dead plants and animal material accumulate faster than can be decomposed by bacteria and they settle to the bottom of the water bodies. Dead plant matter has a high biological oxygen demand (BOD) and some water bodies give up nearly all their dissolved oxygen in order to allow natural decay (oxidation) of the dead plant material to take place.

   It is obvious that any substance which has a high BOD or which results in the formation of something with a high BOD is a potential pollutant in water. Such substances cause decrease in the concentration of dissolved oxygen in water and this leads to death of aquatic animals such as fish.

   • Due to large masses of dead plants and animals, the penetration of light in water decreases thereby affecting photosynthesis in aquatic plants.

   • If nitrates are consumed in drinking water, they are converted by bacteria in the intestines to more toxic nitrites. After absorption into the blood stream, the nitrites combine with haemoglobin forming metahemoglobin. As a result the oxygen carrying capacity of blood is reduced. Babies suffering from excess intake of nitrates have been known to go “blue” due to lack of oxygen in their blood, a condition reffered to as ‘blue babies’.
Fig. 4.25: Pollution from chemicals e.g. foliar sprays

Note: Pollutants know no boundaries. Neighbouring countries should minimise the use of chemical fertilisers. They can be substituted with organic manure which releases nitrates very slowly giving time for crops to use nitrates released.

Exercises 4.4
1. (a) In the preparation of nitric acid, all apparatus are made of glass. Explain why?
   (b) Explain the formation of a “brown ring” during the test for nitrates.
   (c) Why does the “brown ring” disappear when the mixture in the test-tube is shaken?
2. Explain briefly why an industry which produces fumes of nitrogen(IV) oxide should be located away from people and farming areas. Use chemical equations where possible.
3. (a) In the laboratory preparation of nitric acid, how are red-brown fumes formed in the retort?
   (b) When nitric acid is left in a clear glass container, bubbles are observed. Explain this observation using a chemical equation.
   (c) Give two reasons why potassium nitrate is preferred to sodium nitrate in the preparation of nitric acid.
4. Dilute nitric acid is neutralised with potassium hydroxide. The resulting solution is carefully evaporated to dryness to give a white solid D. On heating D, a new yellow solid E and a gas F are formed.
   (a) Name the substances: D, E and F
   (b) Calculate the percentage of nitrogen in D and E
      \[ (K = 39, \, N = 14, \, O = 16) \]
   (c) Which gases are formed when ammonium chloride is heated with (i) sodium nitrite (ii) D
      Support both cases with chemical equations.
Summary
1. Nitrogen can be prepared in the laboratory from air by removing oxygen, carbon(IV) oxide and water vapour. It will have some traces of noble gases, especially argon.
2. Nitrogen is inert at room temperature because of the strong triple covalent bonds between the two nitrogen atoms which make the molecule.
3. Nitrogen(I) oxide re-lights a brightly glowing splint but unlike oxygen it does not react with nitrogen(II) oxide.
4. Nitrogen(II) oxide forms nitrogen(IV) oxide when it reacts with pure oxygen or oxygen from the air.
5. The ease of reduction (loss of oxygen) decreases from N₂O, NO, NO₂ in that order.
6. Ammonia gas is very soluble in water. It is dried by anhydrous calcium oxide. Ammonia is the only common alkaline gas.
7. Aqueous ammonia also called ammonia solution or ammonium hydroxide reacts with some metal ions to produce precipitates of metal hydroxides.
8. Ammonia does not burn, but can burn in oxygen or air enriched with oxygen.
9. Ammonia reacts with oxygen in the presence of platinum catalyst to produce nitrogen(II) oxide.
10. In the manufacture of ammonia, about 500 atmospheres pressure, a temperature of about 500°C and an iron catalyst are needed.
11. In the manufacture of nitric acid a temperature of 850-900°C and platinum-rhodium catalyst are needed.
12. Nitric acid acts as an oxidising agent because it easily produces oxygen, especially when heated.
13. When nitric acid oxidises a substance it is always reduced to red-brown nitrogen(IV) oxide and water.
14. Potassium and sodium nitrates produce a metal nitrite and oxygen when heated.
15. Mercury(II) nitrate and silver nitrate produce metal, nitrogen(IV) oxide and oxygen when heated.
16. Nitrates of calcium, magnesium, zinc, iron(II), iron(III), lead(II) and copper(II) produce metal oxide, nitrogen(IV) oxide and oxygen when heated.
17. Brown ring formed during the test for nitrates is as a result of the formation of nitroso-iron(II) sulphate (FeSO₄·NO).

Revision Exercise 4
1. (i) Explain how nitrogen, oxygen and argon are obtained during the isolation of nitrogen from air for industrial use.
   (ii) Name two gases which remain gaseous when liquid air is at –200°C.
2. (i) Describe briefly what happens when ammonium nitrite is heated in a boiling tube. Use a chemical equation.
   (ii) Explain why a brightly glowing splint relights in nitrogen(I) oxide. Use a chemical equation.
3. (a) Explain why anhydrous calcium chloride or concentrated sulphuric acid are not used to dry ammonia gas.

(b) What do you observe when aqueous ammonia is added drop by drop and in excess to 10cm³ of the following solutions in separate test tubes:
   (i) zinc sulphate.
   (ii) copper(II) chloride.
   (iii) iron(III) sulphate.

4. (a) Name the two raw materials needed in the manufacture of ammonia.

(b) What are the three conditions needed in the manufacture of ammonia.

5. (a) Explain the chemical changes which take place in the test for nitrates using concentrated sulphuric acid and iron(II) sulphate solution.

(b) Why should iron(II) sulphate solution be freshly prepared in this test?

6. (a) Explain the term “acid rain” in terms of oxides of nitrogen.

(b) Explain how nitrates in lakes or rivers affect:
   (i) oxygen concentration in water.
   (ii) penetration of light in waters of the lakes or rivers.
Unit 5: Sulphur and its compounds

Introduction

Sulphur is in Group VI, Period 3, of the Periodic Table as shown in Fig. 5.1. It has an atomic number of 16. The electronic arrangement of an atom of sulphur is therefore 2.8.6. Sulphur is a yellow solid.

Fig 5.1: Position of sulphur in the Period Table

5.1: Occurrence and extraction of sulphur

Sulphur is one of the elements known since ancient times. It occurs naturally in volcanic regions or regions where volcanoes were formerly active. It is found in its elementary form in several countries such as Italy, Mexico, Japan and the USA. Sulphur can occur as a “free” element or combined in compounds such as:

(i) hydrogen sulphide found in petroleum, coal gas and natural gas.
(ii) metal sulphides e.g. zinc blende (ZnS), iron pyrites (FeS₂), copper pyrites (CuFeS₂), galena (PbS) etc.
(iii) metal sulphates e.g. gypsum (CaSO₄ · 2H₂O).

Extraction of sulphur by Frasch process

The sulphur beds in Texas and Louisiana in USA are as much as 30 metres thick and occur at a depth of 160 metres. Between the surface sulphur and deposit, there is a layer of quicksand which makes it difficult to use common methods in mining.

Hermen Frasch (1852 – 1914) an American chemist devised a method of mining sulphur, hence the process is named after him.
The basic principle of Frasch process is that sulphur is melted underground and pumped up to the surface of the earth. A borehole is made to the underground deposits of sulphur and three concentric pipes are sunk into the bore as illustrated in Fig 5.2.
Functions of the three pipes during the extraction of sulphur

The outermost pipe which has the widest diameter of the three pipes is sunk to the sulphur bed. This pipe delivers super heated water. The second pipe is held a little above the surface of sulphur deposits. This pipe is the outlet for molten sulphur. The inner pipe brings hot compressed air into the sulphur deposits to force molten sulphur up the middle pipe to the storage tanks.

The extraction process

Superheated water at 170°C (under 10 atmospheres pressure) is pumped down the outermost pipe. The purpose of the superheated water is to melt sulphur at 113°C. Hot compressed air at about 15 atmospheres pressure is pumped down the innermost pipe to force froth of molten sulphur up. This molten sulphur rises through the annular space between the inner most pipe and the next. The molten sulphur is collected and stored in tanks. Sulphur cools and solidifies to solid blocks. The sulphur that is obtained from this process is over 99% pure.

5.2: Allotropes of sulphur (i.e. the different forms of sulphur)

Experiment 5.1
To prepare the different forms of sulphur.
A. Rhombic sulphur

**Apparatus and chemical**
- beaker
- filter funnel
- filter paper
- hand lens
- sulphur
- carbon disulphide or methylbenzene

**Note:** Carbon disulphide is a solvent that is very poisonous and inflammable. The experiment should be performed in a fume cupboard.

**Procedure**
1. Half fill a 250 cm$^3$ beaker with water. Heat it to a temperature of about 85$^\circ$C. Turn off the bunsen flame.
2. Place about 25 cm$^3$ of methylbenzene or carbon disulphide in a boiling tube.
3. Warm the methylbenzene or carbon disulphide liquid to about 40$^\circ$C using the hot water in the beaker and add sulphur powder into the boiling tube while shaking well.
4. Filter the solution into another boiling tube or an evaporating dish.
5. Allow the filtrate to cool and evaporate slowly at room temperature in the fume cupboard.
6. Observe the crystals of sulphur with a hand lens.
   - Draw a diagram of one of the crystals observed above.
   - Describe the shape of the crystal you have drawn.

B. Monoclinic sulphur

**Apparatus and chemicals**
- evaporating dish
- source of heat
- a piercing instrument/metal spatula
- hand lens
- powdered sulphur

**Procedure**
1. Transfer about 4 spatula-endful of powdered sulphur into an evaporating dish.
2. Heat slowly until all the powdered sulphur has melted.
3. Allow the liquid to cool until a crust just forms.
4. Pierce two holes in the crust and pour out the molten sulphur through one hole.
   - What is the purpose of the other hole?
5. Remove the crust to expose the crystals.
6. Observe the shape of the crystals formed with the help of a hand lens in the evaporating dish.
• Draw a diagram of one of the crystals observed above.
• Describe the shape of the crystal you have drawn.

**Rhombic sulphur or alpha sulphur** (α – sulphur)
Rhombic sulphur which is also known as alpha (α) sulphur is a yellow crystalline solid with an octahedral shape. It crystallizes from a solution of carbon disulphide. Rhombic sulphur is stable below 96˚C. Above this temperature, needle-shaped crystals of monoclinic sulphur are formed slowly.

![Fig. 5.3: Octahedrally shaped rhombic sulphur](image)

![Fig. 5.4: prism shaped monoclinic sulphur](image)

**Monoclinic sulphur or beta-sulphur** (β-sulphur)
Monoclinic sulphur also known as beta-sulphur (β), is a yellow crystalline solid. It is obtained by allowing molten sulphur to solidify. Long prism shaped needles as shown in Fig 5.4 form on the walls of the container that can be separated from the molten sulphur by pouring off the latter (Experiment 5.1 B). If monoclinic sulphur is kept at room temperature for a few days, it gradually changes to rhombic sulphur.

**Table 5.1: Characteristics of rhombic and monoclinic sulphur**

<table>
<thead>
<tr>
<th>Characteristics</th>
<th>Rhombic sulphur</th>
<th>Monoclinic sulphur</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Shape of the crystals</td>
<td>octahedral</td>
<td>Long prism needle like shaped</td>
</tr>
<tr>
<td>2. Colour of the crystal</td>
<td>Yellow</td>
<td>Yellow</td>
</tr>
<tr>
<td>3. Stability</td>
<td>Below 96˚C</td>
<td>Above 96˚C or 96˚C to 119˚C</td>
</tr>
<tr>
<td>4. Density m.p</td>
<td>2.07 g/cm³</td>
<td>1.96g/cm³</td>
</tr>
<tr>
<td></td>
<td>113˚C</td>
<td>119˚C</td>
</tr>
</tbody>
</table>
Transition temperature of the two allotropes is 96°C.

\[
\text{Rhombic sulphur above 96°C} \quad \text{monoclinic sulphur below 96°C}
\]

There are other two non-crystalline forms of sulphur known as **amorphous sulphur** and **plastic sulphur**

**C. Amorphous sulphur**

**Apparatus and chemicals**

- beakers
- filter funnel
- filter paper
- sodium thiosulphate solution
- dilute hydrochloric acid

**Procedure**

1. Add dilute hydrochloric acid in a beaker containing sodium thiosulphate solution.
2. Wait for about two minutes.
3. Filter to get amorphous sulphur as residue on the filter paper. Examine the solid and describe its appearance.
4. Put some of the solid in carbon disulphide and stir. Record your observation.
5. Put the rest of the solid in water and stir. Record your observation.

\[
\text{Na}_2\text{S}_2\text{O}_3(aq) + 2\text{HCl}(aq) \rightarrow 2\text{NaCl}(aq) + \text{SO}_2(g) + \text{H}_2\text{O}(l) + \text{S}(s)
\]

Amorphous sulphur is non-crystalline and colloidal in appearance. It is finely divided and therefore difficult to filter. It is insoluble in carbon disulphide. In water, it forms a colloidal solution called ‘**milk of sulphur**’.

**D. Plastic sulphur**

**Apparatus and chemicals**

- sulphur powder
- beaker
- boiling tube
**Procedure**

1. Heat sulphur in a boiling tube until it boils.
2. Pour it slowly into cold water in a beaker as shown in Fig. 5.5. Record your observation.

![Diagram of hot molten sulphur being poured into cold water]

**Fig 5.5: Preparation of plastic sulphur**

You will observe a brown rubber-like substance. This is plastic sulphur, a supercooled liquid. This liquid gradually changes to a hard yellow opaque substance or rhombic sulphur. Plastic sulphur is insoluble in carbon disulphide.

**Action of heat on sulphur**

**Experiment 5.2**
To investigate the action of heat on sulphur.

**Apparatus and chemicals**
- boiling tube
- source of heat
- pair of tongs
- sulphur
- water

**Procedure**

1. Transfer about a spatula full of powdered sulphur into a boiling tube.
2. Heat the powdered sulphur gradually as you move it in and out of the flame.
3. Observe any changes and record your observations as you heat. Continue the heating process until a dark brown vapour is formed.
4. Pour the boiling sulphur into a beaker of cold water. What do you observe?
   
   **CAUTION:** The brown vapour formed is poisonous. This experiment must be done in a fume chamber or an open space.
   - Describe all the changes observed while the sulphur was being heated.
   - Describe the appearance of sulphur after it had been cooled in water.
From experiment 5.1 A and B, sulphur can exist in different forms in the same state. This is described as **allotropy**. The crystalline allotropes are rhombic and monoclinic sulphur. Table 5.1 shows a summary of characteristics of rhombic and monoclinic sulphur.

Sulphur undergoes a series of changes when heated as observed in Experiment 5.2. These observations are summarised in Table 5.2

**Table 5.2: Observations of action of heat on sulphur**

<table>
<thead>
<tr>
<th>Temp</th>
<th>Observation</th>
</tr>
</thead>
<tbody>
<tr>
<td>113˚C</td>
<td>Sulphur melts to give a mobile amber liquid. As the temperature rises further; the liquid darkens.</td>
</tr>
<tr>
<td>180˚C</td>
<td>The liquid becomes very thick or viscous. (The liquid becomes like a syrup). The colour changes gradually from, red to black.</td>
</tr>
<tr>
<td>250˚C</td>
<td>The liquid is so thick and viscous that it cannot flow.</td>
</tr>
<tr>
<td>above 250˚C</td>
<td>The colour becomes brighter and the liquid once more becomes mobile.</td>
</tr>
</tbody>
</table>

**Table 5.3: Explanation of what happens when sulphur is heated**

1. Sulphur exists as $S_8$ molecules at room temperature. The crown-shaped molecules are packed neatly together to form crystals.

2. Melts at 113˚C- intermolecular forces are broken.

3. The rings of 8 atoms open up as the molten sulphur is heated more strongly. The long chains entangle and make liquid sulphur viscous.

4. Long chains start to break up to form smaller units. Liquid sulphur becomes less viscous (mobile).

5. Short units escape from liquid as gas – sulphur vapour.
5.3: Physical and chemical properties of sulphur

Physical properties of sulphur

- sulphur is a yellow, non-metallic solid.
- has low melting point.
- it is insoluble in water.
- sulphur is soluble in organic solvents like methylbenzene/carbon disulphide among others.

Chemical properties of sulphur

Experiment 5.3
To investigate the chemical properties of sulphur.

Reaction of sulphur with oxygen
(i) To investigate reaction of sulphur with oxygen.

Apparatus and chemicals
- deflagrating spoon
- source of heat
- sulphur powder

Procedure
1. Place a small amount of sulphur in a deflagrating spoon and heat.
2. What do you observe?
3. Record your observation.

Sulphur burns in air or oxygen with a blue flame to form sulphur(IV) oxide.

\[ \text{Sulphur} + \text{oxygen} \rightarrow \text{sulphur(IV) oxide} \]
\[ S(s) + O_2(g) \rightarrow SO_2(g) \]

When heated it combines directly with other non-metals like hydrogen.

\[ \text{Sulphur} + \text{hydrogen} \xrightarrow{\text{heat}} \text{hydrogen sulphide} \]
\[ S(s) + H_2(g) \rightarrow H_2S(g) \]

Reaction of sulphur with metals
(ii) To investigate the reaction of sulphur with metals

Apparatus and chemicals
- spatula
- source of heat
• ignition tube
• test tube holder
• iron powder
• sulphur powder

Procedure
1. Place spatula-full of sulphur powder and powdered iron in an ignition tube.
2. Heat the mixture.
3. Examine the residue when it has cooled down.
   • What did you observe in each case? Record your observation in your notebook.
   • Suggest other metals that can combine with sulphur.
   • What do we call the above method of preparing salts?
   • What do you conclude?
   • Write balanced equations for the reactions.

Sulphur reacts with most metals to form metal sulphides
A mixture of iron powder and sulphur reacts to form iron(II) sulphide. It is an exothermic reaction. That means heat is produced.

(a) \[ \text{Iron} + \text{sulphur} \rightarrow \text{iron(II) sulphide} + \text{heat} \]
   \[ \text{Fe(s)} + \text{S(s)} \rightarrow \text{FeS(s)} + \text{heat} \]

Other metals such as zinc and copper combine with sulphur to form metal sulphides. For example, copper and sulphur react to form copper sulphide and the reaction is exothermic

(b) \[ \text{Copper} + \text{sulphur} \rightarrow \text{copper(II) sulphide} + \text{heat} \]
   \[ \text{Cu(s)} + \text{S(s)} \rightarrow \text{CuS(s)} + \text{heat} \]

(iii) To investigate the reaction of sulphur with dilute and concentrated acids.

Apparatus and chemicals
• sulphur powder.
• test tubes.
• test tube holders.
• dilute sulphuric, nitric and hydrochloric acids.
• concentrated sulphuric, nitric and hydrochloric acids.
• source of heat

Procedure
1. Place about 1 cm³ of dilute sulphuric acid in a test tube and add a small amount of sulphur powder. What do you observe?
2. Repeat with dilute nitric and hydrochloric acids.
3. Place a small amount of sulphur in a test tube. Add one to two drops of concentrated sulphuric acid and warm. What do you observe? (Caution: reaction of hot concentrated acids with sulphur should be done in a fume chamber or open space.)
4. Repeat with concentrated nitric acid and hydrochloric acid.
   - Do dilute acids react with sulphur?
   - Which concentrated acids react with sulphur?

Sulphur does not react with dilute acids, however, it is oxidised by hot concentrated oxidizing acids like sulphuric acid and nitric acid. With concentrated sulphuric acid, sulphur is oxidised to sulphur(IV) oxide.

\[
\text{Sulphur + sulphuric acid} \rightarrow \text{sulphur(IV) oxide + steam}
\]
\[
S(s) + 2\text{H}_2\text{SO}_4(l) \rightarrow 3\text{SO}_2(g) + 2\text{H}_2\text{O}(g)
\]

Concentrated nitric acid oxidises sulphur to sulphuric acid, while it is reduced to red-brown nitrogen(IV) oxide.

\[
\text{Sulphur + nitric acid} \rightarrow \text{sulphuric acid + nitrogen(IV) oxide + steam}
\]
\[
S(s) + 6\text{HNO}_3(l) \rightarrow \text{H}_2\text{SO}_4(l) + 6\text{NO}_2(g) + 2\text{H}_2\text{O}(g)
\]

**5.4 Uses of sulphur**

Sulphur is used in many ways. Some of the uses have been illustrated in Fig. 5.6(a-i).
1. In the manufacture of sulphuric acid (\(\text{H}_2\text{SO}_4\)) which is used in the manufacture of many other compounds. For example, detergents, plastics, etc., Fig 5.6 (a).
2. For making gun powder, matches etc., Fig 5.6 (b).
3. For manufacture of fire works. Fig 5.6 (c).
4. In the vulcanization (hardening) of rubber. Fig 5.6 (d).
5. For manufacturing germicides, fungicides Fig 5.6 (e).
6. In manufacture of medicines Fig 5.6 (f).
7. In photographic development Fig 5.6(g).
8. For making bleaching agents Fig 5.6 (h).
9. In making artificial hair colours or dyes Fig 5.6 (i).
Fig 5.6: Uses of sulphur

Exercise 5.1
1. Describe the allotropes of sulphur.
2. What is the difference between rhombic and monoclinic sulphur?
3. List the physical properties of sulphur.
4. What are the uses of sulphur.

5.5 Sulphur(IV) oxide (SO₂)

5.6 Preparation of sulphur(IV) oxide

Experiment 5.4
To prepare sulphur(IV) oxide and investigate its properties.
Apparatus and chemicals
- thistle funnel or dropping funnel
- round bottomed flask
- tripod stand
- wire gauze
- source of heat
- delivery tube
- wash bottle
- gas jars
- copper turnings and concentrated sulphuric acid or dilute hydrochloric acid and sodium sulphite.

Procedure
There are two methods of preparing sulphur(IV) oxide in the laboratory.

A. Preparation of sulphur(IV) oxide from concentrated sulphuric acid and copper turnings.
1. Place a few pieces of copper turnings in the round bottomed flask.
2. Release the concentrated sulphuric acid just to cover the copper turnings.
3. Heat the mixture.
4. If the gas is required dry, pass it through concentrated sulphuric acid as shown in Figure 5.7.
5. Collect several gas jars of sulphur(IV) oxide.

![Diagram of the preparation of sulphur(IV) oxide using concentrated sulphuric acid and copper](image)

Fig 5.7: Preparation of sulphur(IV) oxide using concentrated sulphuric acid and copper
Sulphur(IV) oxide is prepared in the laboratory by heating copper turnings with concentrated sulphuric acid. Why do we have to use concentrated sulphuric acid and not dilute sulphuric acid? Dilute sulphuric acid does not react with copper because copper is lower than hydrogen in the reactivity series. The reaction with concentrated sulphuric acid will not occur without heating. When it is heated, sulphur(IV) oxide is evolved.

\[
\text{Copper + sulphuric acid} \xrightarrow{\text{heat}} \text{copper(II) sulphate + water + sulphur(IV) oxide}
\]

\[
\text{Cu(s)} + 2\text{H}_2\text{SO}_4(\text{l}) \xrightarrow{\text{heat}} \text{CuSO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) + \text{SO}_2(\text{g})
\]

**B. Preparation of sulphur(IV) oxide from dilute hydrochloric acid and sodium sulphite**

1. Place some of sodium sulphite in the flat bottomed flask.
2. Put the dilute acid through the dropping funnel to cover sodium sulphite.
   - Record your observations in your notebook.
3. A gas will be evolved, but if the process is slow, heat the flask gently. Strong heat is not required.
4. Follow the rest of the procedure as described in Experiment 5.4A. The set-up of apparatus is as shown in Fig 5.8.

![Diagram](image)

*Fig 5.8: Preparation of sulphur(IV) oxide using dilute hydrochloric acid and sodium sulphite*
• What is the colour of the gas?
• Does it smell?
• What is the density of the gas compared to that of the air?
• What method is suitable for collection of sulphur(IV) oxide? Why is this method preferable?
• Why can’t we collect the gas over water?

In the laboratory, sulphur(IV) oxide can be prepared by action of dilute acid on a sulphite. For example

\[
\text{Sodium sulphite} + \text{hydrochloric acid} \rightarrow \text{sulphur(IV) oxide} + \text{water}
\]

\[
\text{Na}_2\text{SO}_3(s) + 2\text{HCl(aq)} \rightarrow \text{SO}_2(g) + 2\text{NaCl(aq)} + \text{H}_2\text{O(l)}
\]

5.7: Physical and chemical properties of sulphur(IV) oxide

5.8: Physical properties of sulphur(IV) oxide
• It is a colourless gas.
• It has a pungent choking smell. It is a poisonous gas and can cause respiratory diseases.
• It is denser than air and that is why it is collected by downward delivery.
• It is soluble in water.

5.9: Chemical properties of sulphur(IV) oxide

Experiment 5.5
To investigate the chemical properties of sulphur(IV) oxide

I Acid character of sulphur(IV) oxide
To investigate acid character of sulphur(IV) oxide

Apparatus and chemicals
• source of sulphur(IV) oxide
• inverted funnel
• litmus papers (blue and red)
• beaker
• water

Procedure
1. Pass the gas directly into the water using an inverted funnel as shown in Fig 5.9.
2. Test the solution formed with blue litmus paper or a universal indicator.
   • Why do we have to use an inverted funnel and not a delivery tube?
   • What do you observe when blue litmus paper or universal indicator is dropped
     into the solution? What does it tell you about the nature of the gas?
   • What do you conclude about the solubility of sulphur(IV) oxide?

Sulphuric(IV) oxide is very soluble in water. Sulphur(IV) oxide actually reacts with
water to form a solution of sulphurous acid. Sulphurous acid is a weak acid.

\[
\text{Sulphur(IV) oxide} + \text{water} \rightarrow \text{sulphurous acid}
\]

\[
\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq})
\]

That is the reason for change of colour of the universal indicator and the blue litmus
paper. These show the acid nature of the gas. Note that the filter funnel is used to
prevent “sucking back“ because sulphur(IV) is soluble in water.

II Bleaching action of sulphur(IV) oxide.

To investigate the bleaching action of sulphur(IV) oxide

Apparatus and chemicals
   • gas jar full of sulphur(IV) oxide
   • moist coloured flower petals
   • wet red or blue litmus paper
   • card board cover
   • dry red or blue litmus paper

Procedure
Place moist coloured flower petals, wet red or blue litmus paper and dry blue or red
litmus papers in separate gas jars full of sulphur(IV) oxide and observe them over a
period of about 15 – 20 minutes as illustrated in Fig 5.10.
• What happens to the flowers and litmus papers?
• Is there any difference between the wet and the dry litmus papers after 20 minutes?
• What can be the reason(s) for this difference?
• What conclusion(s) can you make from your observations?
• Record your observations in your note book.

![Diagram of the experiment with sulphur(IV) oxide gas, delivery tube, gas jar, coloured flowers or wet and dry litmus papers.]

**Fig 5.10: Bleaching action of sulphur(IV) oxide**

The flowers placed in the gas jars of sulphur(IV) oxide are bleached. The wet blue and red litmus papers are also bleached but there was no effect on the dry blue and red litmus papers. From the above observations, it is quite clear that for bleaching to take place, there must be presence of water because dry sulphur(IV) oxide did not bleach the dry litmus papers.

• What then causes the bleaching?

Sulphur(IV) oxide first reacts with water to form sulphurous acid.

\[
\text{Sulphur(IV) oxide + water} \rightarrow \text{sulphurous acid}
\]

\[
\text{SO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_3(aq)
\]

The sulphurous acid formed then reduces the dye by removing an oxygen atom. Hence the dye is bleached.

\[
dye + \text{sulphurous acid} \rightarrow (\text{dye - oxygen atom}) + \text{sulphuric acid}
\]

\[
dye + \text{H}_2\text{SO}_3(aq) \rightarrow (\text{dye - O}) + \text{H}_2\text{SO}_4(aq)
\]

(coloured) (colourless)

The sulphite (SO$_3^{2-}$) ion readily accepts an oxygen atom from the substance being bleached or reduced (flowers, or wet litmus papers). It is then easily oxidized to sulphate ion (SO$_4^{2-}$).

\[
\text{Sulphite ion} + \text{O} \rightarrow \text{sulphate ion}
\]

\[
\text{SO}_3^{2-}(aq) + \text{O} \rightarrow \text{SO}_4^{2-}(aq)
\]
In this example, sulphur(IV) oxide acts as a bleaching agent and a reducing agent at the same time. We can say that sulphur(IV) oxide bleaches by reducing the dye. This means that sulphur(IV) oxide bleaches by reduction. After sometime, oxygen from the air may combine with the colourless substance to regain the original colour. This is why you find newspaper and straw gradually becoming yellow-brown with time.

III. Reducing action of sulphur(IV) oxide
To investigate the action of sulphur(IV) oxide with potassium dichromate(VI) solution

Apparatus and chemicals
- test tube
- acidified potassium dichromate(VI) solution
- source of sulphur(IV) oxide

Procedure
1. Place about 3 cm$^3$ of acidified potassium dichromate(VI) solution in a test tube. Bubble sulphur(IV) oxide gas through the acidified potassium dichromate(VI)
   - what do you observe? What reaction has taken place?
   - what conclusion can you draw from this reaction?
   - record your observation in your notebook.
   - predict reaction of potassium manganate(VII) with sulphur(IV) oxide gas.

(i) Reaction of sulphur(IV) oxide with potassium dichromate(VI), (K$_2$Cr$_2$O$_7$)
As sulphur(IV) oxide is bubbled through acidified orange potassium dichromate(VI) solution, it turns green. The dichromate(VI), (Cr$_2$O$_7^{2-}$) ions are reduced to chromium ions (Cr$^{3+}$) which are green. The ionic equation for this reaction is as follows:

\[
\text{Reduction:} \quad \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 3\text{SO}_3^{2-}(\text{aq}) + 8\text{H}^+(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 3\text{SO}_4^{2-}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})
\]

The complete equation for the reaction is as follows:

\[
\text{Potassium + sulphur(IV) + sulphuric acid + dichromate(VI) oxide + water} \rightarrow \text{potassium + chromium + water + sulphate + sulphate}
\]

\[
\text{K}_2\text{Cr}_2\text{O}_7(\text{aq}) + 3\text{SO}_2(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + \text{Cr}_2(\text{SO}_4)_3(\text{aq}) + \text{H}_2\text{O}(\text{l})
\]

Orange   \rightarrow   Green
Reaction of sulphur(IV) oxide with oxygen

Sulphur(IV) oxide reacts with oxygen in presence of a heated catalyst, vanadium(V) oxide or platinised asbestos to form sulphur(VI) oxide.

\[
\text{Sulphur(IV) oxide} + \text{oxygen} \rightarrow \text{sulphur(VI) oxide} \\
2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})
\]

In this reaction, sulphur(IV) oxide has been oxidized to sulphur(VI) oxide by addition of oxygen.

(IV) Oxidizing action of sulphur(IV) oxide

To investigate the oxidizing action of sulphur(IV) oxide

Apparatus and chemicals

- deflagrating spoon
- a gas jar full of sulphur(IV) oxide
- magnesium ribbon
- source of heat

Procedure

1. Wrap a piece of magnesium ribbon on a deflagrating spoon and light it.
2. Lower the burning magnesium ribbon into a gas jar full of sulphur(IV) oxide as shown in Fig. 5.11

\[\text{deflagrating spoon} \quad \text{card board cover} \quad \text{gas jar} \quad \text{sulphur(IV) oxide} \quad \text{burning magnesium}\]

3. What do you observe on the sides of the gas jar after magnesium stops burning?
   - Record your observations in the notebook.
   - Explain your observations.
   - What conclusion can you make from the above observations?
   - Suppose the gas jar was full of carbon(IV) oxide. What would you observe? Explain.
Sulphur(IV) oxide does not burn or support burning but heat from a piece of burning magnesium splits sulphur(IV) oxide to sulphur and oxygen. Magnesium continues to burn in oxygen forming magnesium oxide. We observe yellow particles of sulphur and white ash which is magnesium oxide on the sides of the gas jar.

\[
\text{Sulphur(IV) oxide + magnesium} \rightarrow \text{magnesium oxide + sulphur}
\]

\[
\text{SO}_2(g) + 2\text{Mg}(s) \rightarrow 2\text{MgO}(s) + \text{S}(s)
\]

In this experiment, the oxygen from the sulphur(IV) oxide react with the magnesium to form magnesium oxide.

\[\begin{array}{ccc}
\text{Reduction} & & \\
\text{Sulphur(IV) oxide} + \text{magnesium} & \rightarrow & \text{magnesium oxide + sulphur} \\
\text{SO}_2(g) + 2\text{Mg}(s) & \rightarrow & 2\text{MgO}(s) + \text{S}(s)
\end{array}\]

\[\begin{array}{ccc}
\text{Oxidation} & & \\
\text{Hydrogen sulphide} & \rightarrow & \text{sulphur(IV) oxide} + \text{water}
\end{array}\]

\[\begin{array}{ccc}
\text{Reduction} & & \\
\text{Hydrogen sulphide} & \rightarrow & \text{sulphur(IV) oxide + water}
\end{array}\]

**Reduction**

\[
\text{Sulphur(IV) oxide + magnesium} \rightarrow \text{magnesium oxide + sulphur}
\]

\[
\text{SO}_2(g) + 2\text{Mg}(s) \rightarrow 2\text{MgO}(s) + \text{S}(s)
\]

**Oxidation**

\[\begin{array}{ccc}
\text{Hydrogen sulphide} & \rightarrow & \text{sulphur(IV) oxide + water}
\end{array}\]

(V) **Action of sulphur(IV) oxide on hydrogen sulphide**

To investigate the action of sulphur(IV) oxide on hydrogen sulphide.

**Apparatus and chemicals**

- gas jar full of sulphur(IV) oxide
- gas jar full of hydrogen sulphide
- water

1. Add a little water just to wet the two gas jars.
2. Invert the gas jar of hydrogen sulphide over the gas jar of sulphur(IV) oxide to mix the gases as shown in Fig. 5.12.

![Fig. 5.12: Reducing sulphur(IV) oxide with hydrogen sulphide](image-url)
3. Repeat the procedure with no water added.
   • What do you observe?
   • Critically observe the gas jars after the reaction. What do you see?
   • What conclusion can you make from your observation.
   • Record your observations and conclusions in your notebook.

When the gases mix a reaction takes place and a yellow deposit of sulphur is formed on the sides of the gas jars. The dry gases do not react at all. Both sulphur(IV) oxide and hydrogen sulphide gases are reducing agents but hydrogen sulphide is a more powerful reducing agent than sulphur(IV) oxide gas. The sulphur deposit comes from both sulphur(IV) oxide and hydrogen sulphide.

\[
\text{Reduction} \quad \begin{align*}
\text{Hydrogen sulphide} + \text{sulphur(IV) oxide} & \quad \rightarrow \quad \text{water} + \text{sulphur} \\
2H_2S(g) + SO_2(g) & \quad \rightarrow \quad 2H_2O(l) + 3S(s)
\end{align*}
\]

\[
\text{Oxidation}
\]

**Test for sulphur(IV) Oxide**

Sulphur(IV) oxide is a colourless gas which bleaches moist litmus papers.
- Sulphur(IV) oxide decolourises an acidified solution of potassium manganate(VII)
- Sulphur(IV) oxide changes the colour of acidified solution of potassium dichromate(VI) from orange to green.
- Sulphur(IV) oxide changes the colour of a solution of iron(III) chloride from yellow to green. It reduces iron(III) chloride (yellow) to iron(II) chloride (green)

**5.10 Test for sulphites and sulphates**

**Experiment 5.6**
To test for sulphates and sulphites.

**Apparatus and chemicals**
- test tubes
- sodium chloride solution
- dilute sulphuric acid
- copper(II) sulphate solution
- iron(II) sulphate solution
- dilute nitric acid
• barium chloride, barium nitrate solution
• copper(II) sulphite solution
• iron(II) sulphite solution

**Procedure**

1. Place about 5 cm$^3$ of sodium chloride solution, sodium hydroxide solution, dilute sulphuric acid, copper(II) sulphate solution, iron(II) sulphate, copper(II) sulphite and iron(II) sulphite into seven separate test tubes. Label them 1–7 as in Table 5.4.
2. Add a small amount of barium chloride solution into the test tube containing sodium chloride solution.
   • What do you observe?
   What can you conclude from the above observations?
3. To the other 6 test tubes, acidify with excess dilute nitric acid and add barium chloride solution.
   • What do you observe?
   • What can you conclude from the above reactions?

Copy table 5.4 in your notebooks and record your observations and conclusions.

**Table 5.4: Observations and conclusion on test for sulphate and sulphite**

<table>
<thead>
<tr>
<th>Solution tested with barium chloride</th>
<th>Observations</th>
<th>Conclusions</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>without acid</td>
<td>with excess acid</td>
</tr>
<tr>
<td>1. Sodium chloride</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. Sodium hydroxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. Dilute sulphuric acid</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4. Copper(II) sulphate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5. Iron(II) sulphate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6. Copper(II) sulphite</td>
<td></td>
<td></td>
</tr>
<tr>
<td>7. Iron(II) sulphite</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Table 5.5 Tests for sulphites and sulphates**

<table>
<thead>
<tr>
<th>Test Action of barium nitrate or barium chloride solution</th>
<th>Sulphites</th>
<th>Sulphates</th>
</tr>
</thead>
<tbody>
<tr>
<td>A white precipitate is formed when barium chloride or barium nitrate is added. No precipitate when added to an acidified sulphite solution.</td>
<td>A white precipitate is formed when barium chloride or barium nitrate is added to an acidified sulphate solution.</td>
<td></td>
</tr>
<tr>
<td>Ba$^{2+}$(aq) + SO$_3^{2-}$(aq) $\rightarrow$BaSO$_3$(s)</td>
<td>Ba$^{2+}$(aq) + SO$_4^{2-}$(aq) $\rightarrow$BaSO$_4$(s)</td>
<td></td>
</tr>
<tr>
<td><strong>NOTE:</strong> The precipitate dissolves in excess acid.</td>
<td><strong>NOTE:</strong> precipitate is insoluble in excess acid.</td>
<td></td>
</tr>
</tbody>
</table>
Uses of sulphur(IV) oxide
1. It is used for bleaching, as a fumigant and in food preservatives. Fig 5.13 (a), (b) and (f) respectively.
2. Large quantities of sulphur(IV) oxide are used in the contact process for the manufacture of sulphuric acid. Fig 5.13 (c)
3. Sulphur(IV) oxide is used in bleaching straws and as a disinfectant. Fig 5.13 (d) and (e) respectively.
4. Liquid sulphur(IV) oxide has been used in purifying petroleum products. Fig 5.13 (g)

Exercise 5.2
When sulphur is burned in air, sulphur(IV) oxide gas is formed and white fumes of sulphur(VI) oxide.

\[
\begin{align*}
\text{Sulphur} & \quad + \quad \text{oxygen} & \longrightarrow & \quad \text{sulphur(IV) oxide} \\
S(s) & \quad + \quad O_2(g) & \longrightarrow & \quad SO_2(g) \\
2S(s) & \quad + \quad 3O_2(g) & \longrightarrow & \quad 2SO_3(g)
\end{align*}
\]

Sulphur(IV) oxide dissolves in water to form an acidic solution. Most chemical industries including some in Kenya, release sulphur(IV) oxide as a by-product, during the manufacture of some products.
• Name some of the industries you know in our country that produce sulphur(IV) oxide.
• What are some of the pollution effects of sulphur(IV) oxide in the environment?
• Suggest and outline a method of removing sulphur(IV) oxide from waste gases escaping from a factory chimney.
• What is the economic advantage of recovering sulphur(IV) oxide?

5.11 Sulphur(VI) oxide

5.12 Preparation of sulphur(VI) oxide

Note: Preparation of sulphur(VI) oxide might not be possible in the laboratory. This will be studied theoretically.

Let us review reaction of sulphur(VI) oxide with water.

\[
\text{Sulphur (VI) oxide} + \text{water} \rightarrow \text{sulphuric acid}
\]

\[
\text{SO}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(aq)
\]

How then do we obtain sulphur(VI) oxide? It is difficult to produce sulphur(VI) oxide directly. Sulphur(VI) oxide must be obtained by reacting sulphur(IV) oxide with oxygen. Let us study the following experiment.

Experiment 5.7
Preparation of sulphur(VI) oxide

Note: This will be discussed theoretically since sulphur(IV) oxide does not react with oxygen under ordinary conditions, but it is important to know how it can be done.

Apparatus and chemicals
• delivery tubes
• combustion tube
• boiling tube
• beaker
• U-tube
• source of oxygen
• source of sulphur(IV) oxide
• platinised asbestos catalyst
• freezing mixture
• fused calcium chloride
Procedure
1. Dry oxygen and sulphur(IV) oxide can be passed over heated platinised asbestos catalyst in a combustion tube.
2. The resulting product can be passed through a wash bottle placed in a freezing mixture to convert sulphur(VI) oxide gas produced into a solid Fig 5.14. While crystals of sulphur(VI) oxide are formed, fused calcium chloride placed in the U-tube absorbs the moisture that might enter in the apparatus.

\[
2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g)
\]

**Fig 5.14: Preparation of sulphur(VI) oxide**

**Note:** Sulphur(VI) oxide in any form is corrosive to eyes skin, mucous membranes and respiratory tract. Inhalation of sulphur(VI) oxide may cause damage to the respiratory tract and even lung tissue.

**Other methods of preparing sulphur(VI) oxide**
It can also be prepared by heating concentrated sulphuric acid with a large excess of phosphorous(V) oxide.

\[
\text{H}_2\text{SO}_4(l) + \text{P}_2\text{O}_5(s) \rightarrow \text{SO}_3(g) + 2\text{HPO}_3(aq)
\]

Sulphur(VI) oxide is more conveniently prepared on a small scale by distilling fuming sulphuric acid in a retort and collecting the sulphur(VI) oxide in a perfectly dry receiver cooled in a freezing mixture.

\[
\text{H}_2\text{S}_2\text{O}_7(l) \xrightarrow{\text{heat}} \text{H}_2\text{SO}_4(l) + \text{SO}_3(g)
\]
Physical properties of sulphur(VI) oxide

- It is a white crystalline solid which is hygroscopic.
- It can also occur as a colourless liquid.
- It has a sharp irritating smell.
- It reacts vigorously with water producing sulphuric acid. The reaction is highly exothermic.

Chemical properties of sulphur(VI) oxide

**Reaction of sulphur(VI) oxide with water**

Sulphur(VI) oxide reacts vigorously with water to form sulphuric acid. This reaction can be very dangerous if not handled with care as the reaction is extremely vigorous and occurs with hissing sound giving out corrosive vapour i.e. sulphuric acid as shown in Fig. 5.15 below.

\[
\text{Sulphur(VI)} + \text{water} \rightarrow \text{sulphuric acid} + \text{heat}
\]

\[
\text{SO}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(aq) + \text{heat}
\]

*Fig. 5.15: Dissolving sulphur(IV) oxide in water*

Sulphur(VI) oxide produced is directly passed into water via an inverted funnel. Sulphuric acid cannot be prepared directly by inserting the delivery tube into the water because sulphur(VI) oxide is soluble in water and the water would ‘suck back’ through the delivery tube into the heated combustion tube causing the tube to crack. When sulphur(VI) oxide is dissolved in water the reaction is highly exothermic and produces acid spray.

So sulphur(VI) oxide reacts with water vapour in the atmosphere to form white fumes consisting of tiny droplets of sulphuric acid.

2. **Reaction of sulphur(VI) oxide with basic oxides**

Sulphur(VI) oxide easily reacts with basic oxides to form sulphates

eg. Calcium oxide + sulphur(VI) oxide → calcium sulphate

\[
\text{CaO(s)} + \text{SO}_3(g) \rightarrow \text{CaSO}_4(s)
\]
Uses of sulphur(IV) oxide
Sulphur(VI) oxide is used in the industrial manufacture of sulphuric acid.

5.13 Manufacture of sulphuric acid by Contact process
Many tonnes of sulphuric acid are made every year by the Contact process. The stages in Contact process are as follows:

Stage 1: Production of sulphur(IV) oxide
Sulphur(IV) oxide is produced by burning sulphur or roasting metal sulphides in air.
  • Sulphur is burned in a furnace in presence of purified air to form sulphur(IV) oxide.
    \[
    \text{Sulphur} + \text{oxygen} \rightarrow \text{sulphur(IV) Oxide} \\
    S(s) + O_2(g) \rightarrow SO_2(g)
    \]
  • Sulphur(IV) oxide produced from heating metal sulphides e.g.
    \[
    \text{Zinc sulphotde} + \text{oxygen} \rightarrow \text{zinc oxide} + \text{sulphur(IV) oxide} \\
    2ZnS(s) + 3O_2(g) \rightarrow 2 \text{ZnO(s)} + 2 \text{SO}_2(g)
    \]

Therefore, the raw materials required are sulphur, air and water

Stage 2: Conversion of sulphur(IV) oxide to sulphur(VI) oxide
Sulphur(IV) oxide is dried in a drying tower and cooled to an optimum temperature of 450 – 500°C and a pressure of about 2 - 3 atmospheres.
More air is passed through the reactor packed with vanadium(V) oxide (V_2O_5) catalyst. The reaction is exothermic and no external heating is required. Sulphur(IV) oxide is converted to sulphur(VI) oxide.

\[
\text{Sulphur(IV) oxide} + \text{oxygen} \rightarrow \text{sulphur(VI) oxide} + \text{heat} \\
2\text{SO}_2(g) + O_2(g) \rightarrow 2\text{SO}_3(g)
\]

This is a reversible reaction. A reversible reaction is a reaction that can go in either direction.
Any sulphur(IV) oxide, \(\text{(SO}_2\)) not converted to sulphur(VI) oxide, \(\text{SO}_3\), might be released to the atmosphere. This will pollute the atmosphere. Hence all measures must be taken to make sure that no sulphur(IV) oxide \(\text{(SO}_2\)) or very little if any gets into the atmosphere. The unreacted sulphur(IV) oxide, \(\text{SO}_2\), is reacted with calcium hydroxide in the chimneys forming a salt and water Fig. 5.16.

\[
\text{Calcium hydroxide} + \text{sulphur(IV) oxide} \rightarrow \text{calcium sulphite} + \text{water} \\
\text{Ca(OH)}_2(s) + \text{SO}_2(g) \rightarrow \text{CaSO}_3(aq) + \text{H}_2\text{O(l)}
\]
This process is called **scrubbing** the gas.

![Diagram of the contact process]

**Fig. 5.16: Contact process**

**Note:** Vanadium(V) oxide is used as a catalyst because it is cheaper and less easily poisoned by impurities than platinum

**Stage 3: Conversion of sulphur(VI) oxide to sulphuric acid**

Finally the sulphur(VI) oxide is absorbed in concentrated sulphuric acid. An oily liquid called **oleum** is formed.

\[
\text{SO}_3(g) + \text{H}_2\text{SO}_4(l) \rightarrow \text{H}_2\text{S}_2\text{O}_7(l)
\]

Sulphuric acid is obtained by diluting oleum with water.

\[
\text{Oleum} + \text{water} \rightarrow \text{sulphuric acid}
\]

Contact process may be represented as shown in Fig. 5.16 and Fig. 5.17.
Optimum conditions for contact process
- catalyst – vanadium(V) oxide.
- temperatures of 450°C – 500°C.
- pressure of about 2-3 atmospheres.

Application of contact process in the manufacture of sulphuric acid:
East Africa Chemical Plant in Webuye
One of the raw materials used in the manufacture of sulphuric acid is sulphur. The sulphur is imported from the Gulf region. It docks in Mombasa where it is transported to the heavy chemical plant in Webuye. Sulphur arrives in the plant packed in 50 kg bags. They are unpacked into an open yard because if kept in compact form under pressure, sulphur is highly explosive.

Stage 1: Production of sulphur(IV) oxide
Sulphur from the sulphur yard is put in a sulphur pit. At the pit, sulphur is liquified by steam at 120°C. The sulphur pit is designed with many chambers to remove impurities such as silica, mud, pebbles etc. These impurities settle down at the bottom of the chambers. The liquified sulphur is then pumped through a pumping lane to a sulphur burner. Through another pipe, a hot air is added to the sulphur burner at the same time. The liquid sulphur reacts with oxygen in the hot air to form sulphur(IV) oxide.

\[
\text{Sulphur} + \text{oxygen} \rightarrow \text{sulphur(IV) oxide}
\]
\[
S(l) + O_2(g) \rightarrow SO_2(g)
\]
This reaction releases a lot of heat. The temperature ranges from 750°C — 800°C. The heat released is not conducive for the reaction that follows. To reduce the temperature, sulphur(IV) oxide produced is passed through a waste heat boiler. At the waste heat boiler, hot air is cooled with water and air. In the process, because of the high temperatures, steam is released which is used to liquify the sulphur. The sulphur(IV) oxide leaves the waste heat boiler at lower temperature of about 500°C.

Stage 2: Conversion of sulphur(IV) oxide to sulphur(VI) oxide
The sulphur(IV) oxide at 500°C is then passed through gas filters to remove impurities like moisture and any other impurities because if allowed to come into contact with the catalyst, the catalyst gets poisoned.

The purified sulphur(IV) oxide is passed through a converter with vanadium(V) oxide (V₂O₅) catalyst. The catalyst is arranged in four chambers namely 1st bed, 2nd bed, 3rd bed and 4th bed as illustrated in Fig. 5.18.

![Fig. 5.18: Sulphur(IV) oxide converter](image)

The catalyst helps to take the oxygen very fast from air and fix it to sulphur(IV) oxide to form sulphur(VI) oxide. The reaction in this chamber is reversible.

\[
\text{Sulphur(IV) oxide} + \text{oxygen} \rightleftharpoons \text{sulphur(VI) oxide} \\
2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})
\]

To make sure that the reaction goes forward, the temperatures must be kept between 450°C — 500°C.

Stage 3: Conversion of sulphur(VI) oxide to sulphuric acid
At the absorption tower, much lower temperatures are required for maximum absorption to take place. The hot sulphur(VI) oxide from the converter is passed through a heat exchanger which reduces the gas temperature to between 180°C — 200°C.

Sulphur(VI) oxide at 180°C — 200°C is passed to the absorption tower where there is circulation of 98.5% sulphuric acid. This concentrated sulphuric acid absorbs the sulphur(VI) oxide. The optimum temperatures for absorption by 98.5% sulphuric acid is also between 180°C — 200°C. At higher temperatures very little of sulphur(VI)
oxide is absorbed and much go to waste through the tower which release the gas high up into the atmosphere. At correct temperatures, almost all the gas is absorbed. The 98.5% sulphuric acid absorbs sulphur(VI) oxide to form an oily liquid called oleum.

\[
\text{Sulphuric acid} + \text{sulphur(VI) oxide} \rightarrow \text{oleum}
\]

\[
H_2SO_4(l) + SO_3(g) \rightarrow H_2S_2O_7(l)
\]

The oleum is then passed to the circulating tank. At the circulating tank, oleum is added to water. Dilution takes place and oleum is converted to 98.5% concentrated sulphuric acid. This process also produces heat. To reduce the temperature, the concentrated sulphuric acid is passed through the cooling towers. From the cooling towers the acid is stored in storage tanks. The process as applied in the heavy chemical industry at Webuye is illustrated in Fig. 5.19.

**Fig. 5.19: A flow chart of sulphuric acid plant Webuye**

5.14: Pollution control in Contact process

**Environmental aspects in Contact process**

1. In stage 2, 99.5% of the sulphur(IV) oxide gets converted into sulphur(VI) oxide. Sulphur(VI) oxide causes acid rain. In Kenya, particularly Thika where (Kel Chemicals Limited) and Webuye East Africa Heavy Chemicals among other sites is situated, residents have complaints regarding:
   - roofs disintegrating due to corrosion.
   - plants not doing well in the region.
   - dying of fish in rivers.
   - people especially children complaining of respiratory complications.
   - paint coating in cars to wear out very fast.
Control of Pollution in Contact process
To avoid this, controls are needed on the waste gas during Contact process. These include:

- recycling sulphur(IV) oxide to the catalytic chambers.
- converting the toxic wastes into harmless products.
- people working in these industries should be well protected and not exposed to these toxic waste.
- Solid calcium hydroxide or calcium carbonate can be used to neutralise sulphur(IV) oxide before it leaves the chimney. In the process calcium sulphate is formed which can be used in the building industry to make plaster.

2. The reaction in each stage gives out heat. As much heat energy as possible should be conserved and reused.

5.15: Properties of sulphuric acid

(a) Concentrated sulphuric acid

(i) Physical properties of concentrated sulphuric acid
- concentrated sulphuric acid is a colourless liquid.
- it is a viscous dense oily liquid whose density is 1.83g/cm³.
- it is a non-volatile acid.
- it has a high boiling point of 338°C

(ii) Chemical properties of concentrated sulphuric acid
Note: Sulphuric acid is extremely corrosive. Whenever you handle it, do it with a lot of care. In case of any accidents, like spilling on the skin, wash with plenty of water. Then rinse with a dilute alkali.

Experiment 5.8
To prepare a solution of dilute sulphuric acid

Apparatus and chemicals
- boiling tube
- concentrated sulphuric acid
- water
- teat pipette

Procedure
1. Half fill a test-tube with cold water.
2. Add 2 -3 drops of concentrated sulphuric acid slowly and carefully down the side of the slanted tube.
3. Hold the bottom of the tube in your palm.
• What do you observe?
• How do you feel when you hold the test tube in your palm?
• What do you conclude?

When concentrated acid is mixed with water, it produces a lot of heat. That is why when you touch the sides of the test tubes, it feels hot. The heat produced is so much that if the water is poured directly to the acid, it can produce acid sprays which are harmful.

Note: When diluting concentrated sulphuric acid never add water to the acid. The acid should be added to a large volume of water carefully and slowly down the sides of the container.

Concentrated sulphuric acid as a dehydrating agent
What is a dehydrating agent?
A dehydrating agent is a substance that removes water molecules or elements that make water from other substances.

Experiment 5.9
To investigate properties of concentrated sulphuric acid as a dehydrating agent.

Apparatus and chemicals
• hard glass tubes
• test tube rack
• teat pipette
• spatula
• concentrated sulphuric acid
• sucrose (common sugar)
• hydrated copper(II) sulphate
• ethanol

Procedure
(i) Sucrose
1. Place sucrose to a quarter of a 100 cm$^3$ beaker.
2. Pour concentrated sulphuric acid using a measuring cylinder to cover the sugar and leave the beaker to stand for sometime. What do you observe?
   Copy Table 5.6 in your notebook and record your observations and conclusions.

(ii) Hydrated copper(II) sulphate
1. Place a few crystals of copper(II) sulphate in a test tube.
2. Add enough concentrated sulphuric acid to cover the crystals completely.
3. Leave it for a while.
4. Record your observations and conclusions in your notebook as illustrated in Table 5.6.
(ii) Ethanol
1. Place about 10 cm³ of ethanol in a hard glass tube.
2. Add 3 to 4 drops of concentrated sulphuric acid to the ethanol in the glass tube and heat the mixture.
3. Record your observation and conclusions in your notebook.

Table 5.6: Reactions of concentrated sulphuric acid with sucrose, ethanol and hydrated copper(II) sulphate

<table>
<thead>
<tr>
<th>Reactions</th>
<th>Observation</th>
<th>Conclusion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reaction of concentrated sulphuric acid with sucrose.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Reaction of concentrated sulphuric acid with hydrated copper(II) sulphate.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Reaction of concentrated sulphuric acid with ethanol</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Reaction of concentrated sulphuric acid and hydrated copper(II) sulphate**
Concentrated sulphuric acid dehydrates copper(II) sulphate - 5 - water i.e. CuSO₄·5H₂O by removing the water of crystallization. The blue hydrated copper(II) sulphate changes to white anhydrous copper(II) sulphate powder.

\[
\text{Hydrated copper(II) sulphate} \rightarrow \text{con. H}_2\text{SO}_4 \rightarrow \text{anhydrous copper(II) sulphate} + \text{water}
\]

\[
\text{CuSO}_4 \cdot 5\text{H}_2\text{O(s)} \rightarrow \text{conc. H}_2\text{SO}_4 \rightarrow \text{CuSO}_4(s) + 5\text{H}_2\text{O(1)}
\]

Because of its great affinity for water, the acid can absorb water from the atmosphere. Because of that, concentrated sulphuric acid is said to be **hygroscopic** hence used for **drying gases**.

**Note:** Concentrated sulphuric acid is not used to dry gases that react with it. This include an alkaline gas like ammonia.

**Reaction of concentrated sulphuric acid with sucrose**
Concentrated sulphuric acid dehydrates sucrose. The mixture turns brown then black and the beaker becomes very hot. A black spongy substance is formed. This black solid is carbon. This black spongy carbon, swells and fills the container. Steam is also produced.

\[
\text{Sucrose} \rightarrow \text{conc. H}_2\text{SO}_4 \rightarrow \text{carbon} + \text{water}
\]

\[
\text{C}_{12}\text{H}_{22}\text{O}_{11(s)} \rightarrow \text{conc. H}_2\text{SO}_4 \rightarrow 12\text{C(s)} + 11\text{H}_2\text{O(g)}
\]
Concentrated sulphuric acid has dehydrated sucrose by removing hydrogen and oxygen to form water leaving behind a black spongy mass of carbon.

**Concentrated sulphuric acid and ethanol**
Ethanol is dehydrated by concentrated sulphuric acid at about 170°C to produce a colourless gas, ethene. That is why production of bubbles of colourless gas is observed.

\[
\text{Ethanol} \xrightarrow{\text{conc. H}_2\text{SO}_4} \text{ethene + water}
\]

\[
\text{CH}_3\text{CH}_2\text{OH(aq)} \xrightarrow{\text{conc. H}_2\text{SO}_4} \text{CH}_2=\text{CH}_2 + \text{H}_2\text{O(l)}
\]

• From the dehydrating properties of concentrated sulphuric acid, explain how concentrated sulphuric acid attack organic substances such as clothes, wood, paper etc.

**Characteristics of dehydrating agent**
• it has a very strong affinity for water.
• it removes water of crystallization from compounds.
• it decomposes compounds by removing hydrogen and oxygen to form water.

**Concentrated sulphuric acid as an oxidizing agent**

**Experiment 5.10**
To investigate concentrated sulphuric acid as an oxidizing agent.

**Apparatus and chemicals**
• test tubes
• test tube rack
• spatula
• red and blue litmus papers
• glass rod
• concentrated sulphuric acid
• copper turnings
• zinc
• sulphur
• carbon
• filter paper soaked in acidified potassium dichromate(VI) or potassium manganate(VII) solution.

**Procedure**
1. Place a spatulaful of copper turnings into a test tube.
2. Add about 5cm³ of concentrated sulphuric acid and heat the mixture.
3. Test the gas produced by holding the filter paper soaked in acidified potassium dichromate(VI) solution near the mouth of the test tube. What do you observe?
4. Repeat the same procedure with zinc, sulphur, and carbon. Heating is not required for zinc.
5. Record your observation and conclusions in your notebook as illustrated in Table 5.7.

**Check for the following**
- colour change of potassium dichromate(VI) paper.
- colour of substances formed. Predict the substances formed from these reactions.

*Table 5.7: Observations and conclusions of reaction of concentrated sulphuric acid on metals and non-metals*

<table>
<thead>
<tr>
<th>Reaction of conc H₂SO₄ with</th>
<th>Observations</th>
<th>Conclusions</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Copper turnings</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. Zinc</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. Sulphur</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4. Carbon</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Write word and balanced chemical equations for the reactions.

Concentrated sulphuric acid oxidizes all metals except those that are very low in the reactivity series such as gold and platinum.

Copper turnings (red-brown solid) is oxidized by concentrated sulphuric acid to form copper(II) sulphate which is a blue solution.

\[
\text{Copper (brown) + sulphuric acid} \rightarrow \text{copper(II) sulphate + water + sulphur(IV) oxide}
\]

\[
\text{Cu(s) + 2H₂SO₄(l) } \rightarrow \text{CuSO₄(aq) + SO₂(g) + 2H₂O(l)}
\]

The potassium dichromate(VI) paper or solution turns from orange to green because of the presence of sulphur(IV) oxide which is produced when sulphuric acid is reduced. Copper metal is oxidized by losing electrons to form copper sulphate. Copper in copper sulphate solution is in ionic form.

\[
\text{Cu(s) } \rightarrow \text{Cu}^{2+}(aq) + 2e^{-}
\]

Oxidation
Zinc reacts in the same way as copper with concentrated sulphuric acid.

\[
\text{Zinc} + \text{sulphuric acid} \rightarrow \text{zinc sulphate + water}
\]

\[
\text{Zn}(s) + \text{H}_2\text{SO}_4(l) \rightarrow \text{ZnSO}_4(aq) + \text{SO}_2(g) + \text{H}_2\text{O}(l)
\]

Oxidation

\[
\text{Zn}(s) \rightarrow \text{Zn}^{2+} + 2e^-
\]

It also oxidizes non-metals like sulphur and carbon.

\[
\text{Sulphur} + \text{sulphuric acid} \rightarrow \text{sulphur(IV) oxide + water}
\]

\[
\text{S}(s) + 2\text{H}_2\text{SO}_4(l) \rightarrow 3\text{SO}_2(g) + 2\text{H}_2\text{O}(l)
\]

Oxidation

\[
\text{Carbon} + \text{sulphuric acid} \rightarrow \text{carbon(IV) oxide + sulphur(IV) oxide + water}
\]

\[
\text{C}(s) + 2\text{H}_2\text{SO}_4(l) \rightarrow \text{CO}_2(g) + 2\text{SO}_2(g) + 2\text{H}_2\text{O}(l)
\]

Oxidation

Note: When the filter paper soaked in solution of potassium manganate(VII) is used, it becomes decolourised.

Displacement reactions of concentrated sulphuric acid

Concentrated sulphuric acid has a higher boiling point than concentrated solutions of nitric acid and hydrochloric acid. It is therefore a non-volatile acid. Hydrochloric and nitric acids are more volatile than sulphuric acid. They are displaced from the reaction mixture as vapour more easily. Concentrated sulphuric acid therefore displaces more volatile acids such as nitric and hydrochloric acids from their salts. Usually an acid salt is formed.

- Why is a normal salt not formed?

    \[
    \text{Potassium} + \text{sulphuric acid} \rightarrow \text{potassium nitrate + nitric acid}
    \]

    \[
    \text{KNO}_3(s) + \text{H}_2\text{SO}_4(q) \rightarrow \text{KHSO}_4(aq) + \text{HNO}_3(aq)
    \]

    \[
    \text{Sodium} + \text{sulphuric acid} \rightarrow \text{sodium chloride + hydrogen chloride}
    \]

    \[
    \text{NaCl}(s) + \text{H}_2\text{SO}_4(l) \rightarrow \text{NaHSO}_4 + \text{HCl}(g)
    \]

The flask cannot stand the temperature at which a normal salt would be formed. The high temperature would melt the glass.
(b) Properties of dilute sulphuric acid

**Exercise 5.3**
Follow the procedures indicated in Table 5.8 and write correct observations and conclusions. Write balanced chemical equations for each reaction.

**Table 5.8: Observation and conclusion of reactions of dilute sulphuric acid**

<table>
<thead>
<tr>
<th>Procedure</th>
<th>Observation</th>
<th>Conclusion</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Take three test tubes. In each of the test tubes, place small amount of magnesium ribbon, zinc granules and copper turnings respectively. Add about 2 -3 cm³ of dilute sulphuric acid. Test the gas evolved with a burning splint.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. Take another three test tubes. In each of the test tubes, place small amounts of potassium carbonate, sodium hydrogen-carbonate and calcium carbonate respectively. Test for any gas produced with a burning splint and aqueous calcium hydroxide.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. Place a small amount of copper(II) oxide powder in a test tube. Add about 10 cm³ of dilute sulphuric acid and warm. Repeat with copper(II) hydroxide.</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Physical properties of dilute sulphuric acid**
- it has a sour taste
- it turns blue litmus red
- it is a colourless solution

**Chemical properties of dilute sulphuric acid**

1. **Reaction of dilute sulphuric acid with metals.**
   Dilute sulphuric acid generally reacts with some metals to form salt and water only. The general equation for the reaction is:

   Sulphuric acid + metal $\rightarrow$ a salt + hydrogen
Let us consider Exercise 5.3 Table 5.7 procedure 1. Did all the metals used react with the dilute sulphuric acid? Dilute sulphuric acid will react with most metals. As you go down the reactivity series the reaction is less vigorous. Metals like copper, silver and gold do not react with dilute sulphuric acid. The other reactions produce hydrogen gas.

\[
\text{Sulphuric acid} + \text{magnesium} \rightarrow \text{magnesium sulphate} + \text{hydrogen} \\
\text{H}_2\text{SO}_4(\text{aq}) + \text{Mg}(s) \rightarrow \text{MgSO}_4(\text{aq}) + \text{H}_2(\text{g}) \\
\]

\[
\text{Sulphuric acid} + \text{zinc} \rightarrow \text{zinc sulphate} + \text{hydrogen} \\
\text{H}_2\text{SO}_4(\text{aq}) + \text{Zn}(s) \rightarrow \text{ZnSO}_4(\text{g}) + \text{H}_2(\text{g}) \\
\]

(ii) **Reaction of dilute sulphuric acid with carbonates and hydrogen carbonates**

Dilute sulphuric acid reacts with metal carbonates and hydrogen carbonates to form a salt, carbon(IV) oxide and water. The general equations are:

\[
\text{Metal} + \text{sulphuric} \rightarrow \text{salt} + \text{carbon(IV)} + \text{water} \\
\text{carbonate} \text{acid} \text{oxide} \\
\text{M}_2\text{CO}_3(s) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{M}_2\text{SO}_4(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \\
\]

\[
\text{Metal} + \text{sulphuric} \rightarrow \text{salt} + \text{carbon(IV)} + \text{water} \\
\text{hydrogencarbonate} \text{acid} \text{oxide} \\
2\text{MHC}_2\text{O}_3(s) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{M}_2\text{SO}_4(\text{aq}) + 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \\
\]

The gas produced extinguishes a burning splint and forms a white precipitate with aqueous calcium hydroxide. This is the test for carbon(IV) oxide.

Potassium carbonate and sodium hydrogen carbonates react with dilute sulphuric acid to produce soluble salts, carbon(IV) oxide gas and water.

\[
\text{Potassium} + \text{sulphuric} \rightarrow \text{potassium} + \text{carbon(IV)} + \text{water} \\
\text{carbonate} \text{acid} \text{oxide} \\
\text{K}_2\text{CO}_3(s) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O} \\
\]

\[
\text{Sodium} + \text{sulphuric} \rightarrow \text{sodium} + \text{carbon(IV)} + \text{water} \\
\text{hydrogencarbonate} \text{acid} \text{oxide} \\
2\text{NaHCO}_3(s) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \\
\]

When dilute sulphuric acid reacts with calcium carbonate, the reaction starts and soon stops. This is evident by the evolution of carbon(IV) oxide for sometime. The calcium sulphate produced is a slightly soluble salt. It forms as a coating on the surface of the unreacted calcium carbonate. This prevents any further reaction to take place. Dilute sulphuric acid therefore reacts with both calcium carbonate and hydrogen carbonates to form a slightly soluble salt, water and carbon(IV) oxide.
Calcium + sulphuric acid → calcium + carbon(IV) + water

CaCO₃(s) + H₂SO₄(aq) → CaSO₄(s) + CO₂(g) + H₂O(l)

- Identify soluble and insoluble sulphates.
- Predict from which compounds they are formed (consider only carbonates and hydrogen carbonates reacting with dilute sulphuric acid)

(iii) Reaction of dilute sulphuric acid with oxides and hydroxides
Basic oxides and hydroxides react with dilute sulphuric acid to form neutral solutions of salt and water. They are known as neutralization reactions.

The general equation for the reaction is

\[
\text{Acid} + \text{metal oxide} \rightarrow \text{a salt} + \text{water} \\
\text{Acid} + \text{metal hydroxide} \rightarrow \text{a salt} + \text{water}
\]

When warm dilute sulphuric acid is reacted with copper(II) oxide powder, the solid dissolves and a blue solution of copper(II) sulphate is formed. Dilute sulphuric acid also neutralizes sodium hydroxide to form salt and water only.

\[
\text{Copper(II) oxide} + \text{sulphuric acid} \rightarrow \text{copper(II) sulphate} + \text{water} \\
\text{CuO(s)} + \text{H₂SO₄(aq)} \rightarrow \text{CuSO₄(aq)} + \text{H₂O(l)}
\]

\[
\text{Sodium hydroxide} + \text{sulphuric acid} \rightarrow \text{sodium sulphate} + \text{water} \\
2\text{NaOH(aq)} + \text{H₂SO₄(aq)} \rightarrow \text{Na₂SO₄(aq)} + \text{2H₂O(l)}
\]

Uses of sulphuric acid
Sulphuric acid is used in many ways. Some of them are illustrated in Fig. 5.20.
1. To manufacture fertilisers like ammonium sulphate. Fig 5.20 (c)
2. In the production of synthetic fibres. Fig 5.20 (b)
3. In petroleum refining
4. In car batteries. Fig 5.20 (e)
5. In cleaning of metals. Fig 5.20 (d)
6. To make soaps and detergents. Fig 5.20 (a)
7. To make dyes, drugs, paints, explosives and many other compounds.
Exercise 5.4
1. Sulphuric acid is manufactured by Contact process.
   (a) What are the raw materials for the process?
   (b) Draw a flow diagram showing the steps in making sulphuric acid. Include equations.
   (c) What is the catalyst used in the process?
   (d) In the last step, why don’t we dissolve the sulphur(VI) oxide in water?
   (e) List four uses of sulphuric acid.

2. Oleum (H₂S₂O₇) is an intermediate in the industrial manufacture of sulphuric acid. How is oleum converted into sulphuric acid?

5.16: Hydrogen sulphide
How would hydrogen sulphide gas be prepared in the laboratory?

5.17: Preparation and physical properties of hydrogen sulphide

Experiment 5.11
Laboratory preparation of hydrogen sulphide gas.

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Note: This experiment will be treated only theoretically.

Apparatus and chemicals
- thistle funnel
- flat bottomed flask
- delivery tube
- trough
- beehive shelf
- gas jar
- dilute hydrochloric acid
- iron(II) sulphide

Procedure
1. The apparatus is set up as illustrated in Fig 5.21.
2. Iron(II) sulphide is placed in flat bottomed flask.
3. Then dilute hydrochloric acid is added to the iron(II) sulphide.
4. The gas is collected by downward delivery.

Fig 5.21: Preparation of hydrogen sulphide
- Suggest a test for hydrogen sulphide. What is the purpose of the water in the wash bottle?
- The reducing action of hydrogen sulphide gas can be tested by bubbling the gas through a solution of bromine water, Iron(III) chloride solution, acidified potassium manganate(VII) solution, concentrated nitric acid and iron(III) sulphate separately.
- What property is tested by the method of collection?

Hydrogen sulphide is generally prepared in the laboratory by the action of a dilute acid on a metal sulphide.

\[
\text{Metal} + \text{dilute \ acid} \rightarrow \text{metal salt} + \text{hydrogen sulphide}
\]
In our case, dilute hydrochloric acid and iron(II) sulphide have been used

\[
\begin{align*}
\text{Iron(II)} & \ + \ \text{hydrochloric} & \rightarrow \ \text{iron(II)} & \ + \ \text{hydrogen} \\
\text{sulphide} & \quad \text{acid} & \quad \text{chloride} & \quad \text{sulphide} \\
\text{FeS(s)} & \ + \ 2\text{HCl(aq)} & \rightarrow \ \text{FeCl}_2(s) & \ + \ \text{H}_2\text{S(g)} \\
\end{align*}
\]

Hydrogen sulphide gas is bubbled through distilled water to remove hydrogen chloride gas and then collected by downward delivery because it is denser than air. The gas can also be collected over warm water. It is fairly soluble in water.

**Test for hydrogen sulphide**

Hydrogen sulphide is tested by a strip of filter paper soaked in lead(II) ethanoate solution. The paper turns black. The colour change is caused by precipitation of black lead(II) sulphide.

\[
\begin{align*}
\text{Lead(II) ethanoate} & \ + \ \text{hydrogen} & \rightarrow \ \text{lead} & \ + \ \text{ethanoic} \\
\text{sulphide} & \quad \text{sulphide} & \quad \text{acid} \\
\text{Pb(C}_2\text{H}_3\text{O}_2\text{)_2(aq)} & \ + \ \text{H}_2\text{S(g)} & \rightarrow \ \text{PbS(s)} & \ + \ 2\text{C}_2\text{H}_4\text{O}_2\text{(aq)} \\
\end{align*}
\]

**Physical properties of hydrogen sulphide**

Hydrogen sulphide gas
- is colourless.
- has a repulsive smell of rotten eggs - the proteins in eggs forms the gas when they rot.
- is denser than air.
- is fairly soluble in water.

**Chemical properties of hydrogen sulphide**

**Reducing action of hydrogen sulphide**

Hydrogen sulphide is readily oxidised and therefore acts as a very powerful reducing agent. Table 5.9 shows what is observed when hydrogen sulphide gas is bubbled through various solutions.
Table 5.9: The reactions of hydrogen sulphide as a reducing agent

<table>
<thead>
<tr>
<th>Solution</th>
<th>Observation</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Bromine water</td>
<td>• Solution turns from orange to colourless.</td>
</tr>
<tr>
<td></td>
<td>• a yellow solid is formed.</td>
</tr>
<tr>
<td>2. Iron(III) chloride solution</td>
<td>• solution turns from yellow to pale green.</td>
</tr>
<tr>
<td></td>
<td>• a yellow solid is formed.</td>
</tr>
<tr>
<td>3. Acidified potassium manganate(VII) solution</td>
<td>• solution turns from purple to colourless.</td>
</tr>
<tr>
<td></td>
<td>• a yellow solid is formed.</td>
</tr>
<tr>
<td>4. Acidified potassium dichromate(VI) solution</td>
<td>• solution turns from orange to green.</td>
</tr>
<tr>
<td></td>
<td>• a yellow solid is formed.</td>
</tr>
<tr>
<td>5. Hydrogen peroxide solution</td>
<td>• a yellow solid is formed.</td>
</tr>
<tr>
<td>6. Concentrated nitric acid</td>
<td>• a red-brown gas is produced.</td>
</tr>
<tr>
<td></td>
<td>• a yellow solid is formed.</td>
</tr>
<tr>
<td>7. Iron(III) sulphate solution</td>
<td>• Solution turns from yellow to pale green.</td>
</tr>
<tr>
<td></td>
<td>• a yellow solid is formed.</td>
</tr>
</tbody>
</table>

• What is common in all the observations in Table 5.9? Explain your answer.
• Write balanced equations for the reactions.

The equations for the above equations are as follows:

1. Bromine water $+$ hydrogen sulphide $\rightarrow$ hydrogen bromide $+$ sulphur

$$\text{Br}_2(aq) \quad \text{H}_2\text{S(g)} \quad \rightarrow \quad 2\text{HBr(aq)} \quad + \quad \text{S(s)}$$

Bromine has been reduced to bromide ions

Reduction

$$\text{Br}_2(aq) + 2e^- \rightarrow 2\text{Br}^{-}(aq)$$
2. Hydrogen sulphide reduces iron(III) ions to iron(II) ions.

\[
\text{Iron(III)} + \text{hydrogen} \rightarrow \text{iron(II)} + \text{hydrochloric} + \text{sulphur}
\]

\[
\text{chloride sulphide chloride acid}
\]

\[
2\text{FeCl}_3(\text{aq}) + \text{H}_2\text{S(g)} \rightarrow 2\text{FeCl}_2(\text{aq}) + 2\text{HCl(}aq) + \text{S(s)}
\]

\[
(\text{yellow}) \quad (\text{pale green}) \quad (\text{yellow solid})
\]

i.e. \(\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})\)

3. Hydrogen sulphide reduces acidified manganate(VII) ions to manganese(II) ions. That is why the colour changes from purple to colourless. In the process, hydrogen sulphide is oxidized to sulphur (yellow solid).

\[
\text{Manganate(VII)} + \text{hydrogen} + \text{hydrogen} \rightarrow \text{manganese(II)} + \text{sulphur} + \text{water}
\]

\[
\text{ion sulphide ion ion}
\]

\[
2\text{MnO}_4^{-}(\text{aq}) + 5\text{H}_2\text{S(g)} + 6\text{H}^{+}(\text{aq}) \rightarrow 2\text{Mn}^{2+}(\text{aq}) + 5\text{S(s)} + 8\text{H}_2\text{O(l)}
\]

\[
(\text{purple}) \quad (\text{colourless}) \quad (\text{yellow solid})
\]

4. Hydrogen sulphide reduces acidified orange dichromate(VI) ions to green chromium(III) ions

\[
\text{Dichromate(VI)} + \text{hydrogen} + \text{hydrogen} \rightarrow \text{chromium(III)} + \text{sulphur} + \text{water}
\]

\[
\text{ion sulphide ion ion}
\]

\[
(\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 3\text{H}_2\text{S(g)} + 8\text{H}^{+}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 3\text{S(s)} + 7\text{H}_2\text{O(l)}
\]

\[
(\text{orange}) \quad (\text{green}) \quad (\text{yellow solid})
\]

Note: \(\text{H}^+\) ions in equations 3 and 4 are given by dilute sulphuric acid used in acidifying the solutions.

5. Hydrogen sulphide reduces hydrogen peroxide (which is colourless) to water which is colourless. This results in formation of sulphur.

\[
\text{Hydrogen} + \text{hydrogen} \rightarrow \text{water} + \text{sulphur}
\]

\[
\text{peroxide sulphide}
\]

\[
\text{H}_2\text{O}_2(\text{aq}) + \text{H}_2\text{S(g)} \rightarrow 2\text{H}_2\text{O(l)} + \text{S(s)}
\]

\[
(\text{colourless}) \quad (\text{yellow solid})
\]

6. Hydrogen sulphide reduces concentrated nitric acid to nitrogen(IV) oxide gas. This is evident by the red-brown fumes. In the process, hydrogen sulphide is oxidized to sulphur which is a yellow solid.
Nitric acid + hydrogen sulphide $\rightarrow$ nitrogen(IV) oxide + sulphur + water

\[2\text{HNO}_3(aq) + \text{H}_2\text{S}(g) \rightarrow 2\text{NO}_2(g) + \text{S}(s) + 2\text{H}_2\text{O}(l)\]

(red-brown) (yellow solid)

7. Being a stronger reducing agent, hydrogen sulphide reduces sulphur(IV) oxide in solution to form yellow solid of sulphur.

Sulphur(IV) oxide in solution forms sulphite ion (\(\text{SO}_3^{2-}\)) and hydrogen ion (\(\text{H}^+\)).

\[
\text{Hydrogen sulphide} + \text{sulphite} + \text{hydrogen} \rightarrow \text{water} + \text{sulphur sulphide ion} + \text{ion}
\]

\[2\text{H}_2\text{S}(g) + \text{SO}_3^{2-}(aq) + 2\text{H}^+(aq) \rightarrow 3\text{H}_2\text{O}(l) + 3\text{S}(s)\]

(yellow solid)

5.18: Pollution of the environment by compounds of sulphur (Hydrogen sulphide and oxides of sulphur)

Remember that pollution means adding things to our environment which are undesirable. One type of pollution is air pollution. Air is said to be polluted when it contains substances in quantities that have adverse effects on life. Air pollution occurs due to the presence of unnatural gases, particles or chemicals in the air. The presence of hydrogen sulphide, sulphur(IV) oxide and sulphur(VI) oxide in the air cause air pollution. Think about the following questions.

- Under what conditions is air said to be polluted?
- Why is air pollution a serious problem?

Sources of air pollution by hydrogen sulphide and oxides of sulphur

Most fossil fuels contain sulphur as an impurity. In most cities and towns, the major source of air pollution is the burning of fossil fuels such as oils, diesel, petrol and coal. When the fuel is burned, sulphur is oxidised to sulphur(IV) oxide and sulphur(VI) oxide.

Sulphur(IV) oxide (\(\text{SO}_2\)) and sulphur(VI) oxide (\(\text{SO}_3\)) sulphide are also produced by power stations and industries especially those processing sulphide ores. These gases are released into the environment. When sulphur(IV) oxide and sulphur(VI) oxide are released to the atmosphere, they dissolve in rain water to form acidic solutions. Consequently, the rain comes down as acid rain. This has been illustrated in Fig. 5.22 below.

\[\text{e.g. Sulphur(IV) oxide} + \text{water} \rightarrow \text{sulphurous acid}\]

\[\text{SO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_3(aq)\]

Sulphurous acid + oxygen $\rightarrow$ sulphuric acid

\[2\text{H}_2\text{SO}_3(aq) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{SO}_4(aq)\]
Effects of acid rain on the environment

• **On human health**
The harm to people from acid rain is not direct. However, the pollutants that cause acid rain e.g. sulphur(IV) oxide and sulphur(VI) oxide are injurious to human health. These gases can be transported long distance by wind. When inhaled into people’s lungs, the effect is increased illness and premature death from heart and lung disorders such as asthma and bronchitis.

• **On useful materials**
Acid rain contribute to the corrosion of metals and deterioration of paint and stone. These effects seriously reduce the value of buildings, bridges, statues, cultural objects and cars. This can be observed in Webuye town due to sulphur(IV) oxide released by Trans Africa Webuye paper mills.

• **On coating in vehicles**
Acid rain damages coating in vehicles so that their surfaces appear as irregularly shaped and permanently etched. It can easily be observed in vehicles. It occurs after evaporation of moisture droplet. Usually the damage is permanent; once it has occurred, the only solution is to repaint the vehicle.

• **Aluminium compounds in the soil**
The action of acid rain on insoluble aluminium compounds in the soil makes them soluble. Insoluble aluminium compounds are fairly harmless to plants and animals, but with increased acidity, the compounds become soluble and seep
into the soil. Plants then absorb the aluminium ions through the roots, become poisoned and die. Plant roots are also weakened and trees often fall. Acid rain can pass through the soil and end up in rivers and lakes. Fish living in such kind of water die or become poorly developed Fig. 5.23.

![Diagram showing the effects of increased solubility of aluminium compounds due to acid rain.](image)

*Fig. 5.23: Effects of increased solubility of aluminium compounds due to acid rain.*

**Hydrogen sulphide**
- Small quantities of hydrogen sulphide are common in the environment. It is found dissolved in mineral springs, water bodies, waste water and in natural gases. Much of the gas is produced in marshes, standing water, rotting vegetables and sewage treatment plants. It is also produced in various processes in industries such as in the manufacture of chemical fibres, rotting vegetables and oil refineries.

**Effects of hydrogen sulphide**
- **On humans and animals**
  Hydrogen sulphide is an irritant. It is toxic to the nerves and cells of human and animals. It may cause irritation of the eye and respiratory organs, nausea, malfunction of nerves, numbness and finally death due to respiratory paralysis. Illness caused by hydrogen sulphide must be reported immediately to the doctor once detected.

- **Plants**
  Hydrogen sulphide causes only slight damage to plants. The most sensitive plants to the gas are tomatoes, cucumbers and soya beans.

**Steps to reduce pollution**
What can be done to reduce pollution caused by oxides of sulphur and hydrogen sulphide?
It is important that we reduce the level of air pollution if we are to have a healthier environment and sustain life. To achieve this goal, the following can be done.
1. Burn less fossil fuels that contain sulphur by using energy more efficiently. Pollutants like sulphur should be removed from fuels before they are burned. We should use alternative forms of energy.
   - Do you know any ‘cleaner’ ways to produce electricity?
   - Would these ways be environmentally friendly? Discuss.

2. Sulphur(IV) oxide and sulphur(VI) oxide can be converted into harmless compounds before they leave the power stations or industries. For example, a mixture of calcium carbonate and water neutralises the sulphur(IV) oxide.
   - Can you think of other ways to alleviate this problem?

3. There should be strict laws in the country to limit the amount of oxides of sulphur and hydrogen sulphide released to the environment from industries.

Exercise 5.5
1. Describe the preparation and collection of gas jar of hydrogen sulphide starting with powdered sulphur.
2. Describe experiments to show action of hydrogen sulphide gas.
   (a) as a combustible substance.
   (b) as a reducing agent

Summary
1. Sulphur is a yellow solid.
2. Many of the deposits of sulphur are found underground. Sulphur is mined by Frasch process.
3. Sulphur has four allotropes namely rhombic sulphur, monoclinic sulphur, plastic sulphur and amorphous sulphur.
4. Rhombic sulphur is more stable than monoclinic sulphur. The transition temperature between the allotropes is 96°C.
5. Sulphur burns in air or oxygen to form sulphur(IV) oxide.
6. Sulphur(IV) oxide is an acidic gas which turns blue litmus red. It has a pungent irritating smell and is denser than air.
7. Sulphur(IV) oxide is an oxidising agent and also reducing agent.
8. Sulphur(IV) oxide turns acidified potassium dichromate(VI) paper from orange to green. This is a test for sulphur(IV) oxide gas.
9. Sulphuric acid is manufactured by the contact process.
10. Concentrated sulphuric acid is a non-volatile acid. It is a powerful dehydrating and oxidizing agent.
11. Hydrogen sulphide is poisonous and has an unpleasant smell (of a rotten egg). It is denser than air. It is slightly soluble in water forming a weakly acidic solution.
12. Hydrogen sulphide is a reducing agent.
Revision Exercise 5

1. In an experiment, 2.5g of sulphur was obtained by reacting hydrogen sulphide and chlorine as shown by the equation below:

$$\text{H}_2\text{S(g)} + \text{Cl}_2(g) \rightarrow \text{S(s)} + 2\text{HCl(g)}$$

(a) Which of the reactants act as a reducing agent in the above reaction. Explain.

(b) Given that the yield of sulphur in the above reaction is 80%, calculate the number of moles of $\text{H}_2\text{S(g)}$ used in the reaction.

2. List the effects of hydrogen sulphide on the environment.

3. Powdered sulphur was placed in a deflagrating spoon and heated in a Bunsen flame until the sulphur starts to burn.

(a) What is seen when the sulphur burns.

(b) Write an equation for the reaction above.

(c) When the gas in (b) is dissolved in water, name the solution that is formed and write the equation for the reaction.

(d) If a few drops of litmus solution was added to (c) above, what would be the colour of the solution.

(e) Write an equation for the reaction that will occur between the solution formed in (c) above and sodium hydroxide solution.

4. Suggest ways through which air pollution can be reduced.

5. (a) To what extent is acid rain a problem in Kenya?

(b) Is the problem of acid rain likely to improve or get worse in future if nothing is done about it in Kenya? Explain your answer.

6. (a) Acid rain is an international problem. Why?

(b) Should governments of different countries be doing something about the problem? Why do you think so?
6.1: Chlorine
Chlorine is the second element in Group VII of the Periodic Table as shown in Fig. 6.1 below. It is located in Period 3. Using this information, work out the electronic arrangement of an atom of chlorine.

![Fig 6.1: Position of chlorine in the Periodic Table]

Chlorine has an atomic number of 17 and its electronic arrangement is 2.8.7. It is a reactive gaseous element and therefore it does not occur in nature as “free” element. It is usually found combined with sodium forming sodium chloride (rock salt). Rock salt is the main source of chlorine. Sea water contains about 2.6 per cent sodium chloride.

6.2: Laboratory preparation of chlorine
Chlorine is prepared by oxidation of hydrochloric acid using a suitable oxidising agent. The common oxidising agents used are potassium manganate(VII) or manganese(IV) oxide. When potassium manganate(VII) is used as an oxidising agent, no heating is required. However, with manganese(IV) oxide, heat is required for oxidation to take place. The oxidising agents provide the oxygen atoms needed to oxidise hydrochloric acid.

\[
\text{Hydrochloric acid} + \text{oxygen atoms} \rightarrow \text{chlorine} + \text{water}
\]

\[
2\text{HCl(aq)} + \text{O} \rightarrow \text{Cl}_2(g) + \text{H}_2\text{O(l)}
\]

Experiment 6.1
Laboratory preparation of chlorine

Caution: Chlorine gas is very poisonous. It should therefore be prepared in a fume cupboard or an open space.
Apparatus and chemicals
• round-bottomed flask
• thistle funnel
• delivery tubes
• two wash bottles
• several gas jars
• cardboard covers
• concentrated hydrochloric acid
• manganese(IV) oxide
• concentrated sulphuric acid
• blue and red litmus papers

Procedure
1. Put some manganese(IV) oxide in a round-bottomed flask.
2. Arrange the apparatus as shown in Fig 6.2.

Fig 6.2: Laboratory preparation of chlorine using manganese(IV) oxide

3. Add concentrated hydrochloric acid into the flask.
4. Heat the mixture
   • What do you observe in the flask and the gas jar?
4. Test the gas collected in the gas jar with moist blue and red litmus papers.
   • What happens to both papers?
5. Smell the gas by wafting the gas towards your nose. Do not smell the gas directly.
   • Describe the smell of the gas?
   • Collect several gas jars full of chlorine for the following experiments.
Which method of collection have you used? Give a reason why you used this method.

6. Invert one of the gas jars full of this gas into water in a trough.
   • What do you observe?

7. Record the observations in your notebook
   • What conclusions can you make from the above observations?

After heating the mixture of concentrated hydrochloric acid and manganese(IV) oxide, effervescence occurs and a yellow-green gas is given out. Hydrogen chloride gas which is an impurity dissolves in the water in the first wash bottle. Chlorine is then dried by passing it through concentrated sulphuric acid. It is collected by downward delivery because it is denser than air.

The moist blue litmus paper turns red and then white in chlorine.

\[
\text{Hydrochloric } + \text{ manganese(IV) oxide } \xrightarrow{\text{heat}} \text{ manganese(II) + water + chlorine chloride}
\]

\[
4\text{HCl(aq)} + \text{MnO}_2(s) \xrightarrow{\text{heat}} \text{MnCl}_2(aq) + 2\text{H}_2\text{O(l)} + \text{Cl}_2(g)
\]

If potassium manganate(VII) is used, heating is not required as shown in Fig. 6.3. This is because it is a stronger oxidising agent (releases oxygen very easily) than manganese(IV) oxide.

\[
\text{Hydrochloric } + \text{ potassium manganate(VII) } \longrightarrow \text{ potassium + manganese(II) + water + chlorine chloride}
\]

\[
16\text{HCl(aq)} + 2\text{KMnO}_4(s) \longrightarrow 2\text{KCl(aq)} + 2\text{MnCl}_2(aq) + 8\text{H}_2\text{O(l)} + 5\text{Cl}_2(g)
\]

Fig. 6.3: Laboratory preparation of chlorine using potassium manganate(VII)
Another method of preparing chlorine is by the reaction of sodium chloride with concentrated sulphuric acid and manganese(IV) oxide. In this reaction, hydrogen chloride gas is produced and then oxidised by manganese(IV) oxide to chlorine. Note that heat is required in this reaction.

\[
\text{Sulphuric acid} + \text{sodium chloride} \rightarrow \text{hydrochloric acid} + \text{sodium hydrosulphate acid}
\]
\[
\begin{align*}
H_2SO_4(aq) & + NaCl(aq) \rightarrow HCl(g) + NaHSO_4(aq) \\
\text{Hydrochloric acid} + \text{manganese(IV)} & \rightarrow \text{manganese(II)} + \text{water} + \text{chlorine}
\end{align*}
\]
\[
4HCl(aq) + MnO_2(aq) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)
\]

Hydrogen chloride gas is oxidised to chlorine gas.

- Why does moist blue litmus paper turn red first before being bleached by chlorine?

**Note:** If not required dry, chlorine can be collected over brine (concentrated sodium chloride solution).

### 6.3 Properties of chlorine.

(a) **Physical properties of chlorine**

- is a yellow-green gas.
- has an unpleasant, pungent and irritating smell.
- is fairly soluble in water.
- is denser than air.
- is easily liquified under pressure. This makes it easy to store and transport chlorine in liquid state in steel cylinders.

(b) **Chemical properties of chlorine**

**NB:** Chlorine is very poisonous and therefore most of its properties should be demonstrated in a fume cupboard or an open place.

**Experiment 6.2**

Reaction of chlorine with hydrogen.

**Apparatus and chemicals**

- wooden splint
- source of dry hydrogen
- gas jar full of dry chlorine
- glass rod
- concentrated ammonia solution
Procedure
1. Prepare dry hydrogen by reacting zinc granules with dilute sulphuric acid.
2. Allow hydrogen gas to pass through a delivery tube with a jet at the end.
   Let the gas pass through this tube for 2 - 3 minutes before lighting it.
   • Why should this precaution be taken before lighting hydrogen?
3. Light the end of the tube so that hydrogen burns.
   • Do you remember another gas that burns with this type of flame?
4. Lower the tube containing burning hydrogen into a gas jar of dry chlorine as shown in Fig 6.4.

Fig 6.4: Reaction of chlorine and hydrogen

• What do you observe?
5. Test the gas formed in the gas jar with a glass rod previously dipped in concentrated ammonia solution.
   • What do you observe?
   • What conclusion do you make from the above test?
6. Record your observations in your notebook.

Hydrogen is allowed to pass through the delivery tube for sometime so that all the air is displaced from the apparatus. Air-hydrogen mixture is explosive when ignited. Hydrogen burns with a blue flame. Carbon(II) oxide also burns with a blue flame. Dense white fumes are observed when the glass rod containing concentrated ammonia solution is dipped in the gas jar previously containing chlorine. This is because the hydrogen chloride formed from the reaction of hydrogen and chlorine reacts with ammonia to form ammonium chloride.

\[
\text{Ammonia} + \text{hydrogen chloride} \rightarrow \text{ammonium chloride} \\
\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s) \\
\text{white fumes}
\]

We can conclude that chlorine combines with hydrogen to form hydrogen chloride gas.

\[
\text{Hydrogen} + \text{chlorine} \rightarrow \text{hydrogen chloride} \\
\text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2\text{HCl}(g)
\]
Experiment 6.3  
Reaction of chlorine with magnesium and iron.

**Apparatus and chemicals**
- source of dry chlorine
- gas jars
- conical flask
- magnesium ribbon
- iron wire
- deflagrating spoon
- combustion tube

**Procedure**

**A. Magnesium**
1. Clean a magnesium ribbon with a knife or emery paper.
   - Why should we clean the magnesium ribbon?
2. Wrap the magnesium ribbon firmly in a deflagrating spoon and burn it.
   - What do you observe?
3. Insert the burning magnesium ribbon in a gas jar of dry chlorine as shown in Fig. 6.5.
   - What do you observe in the gas jar?
4. Record your observations in your notebook.
   - What conclusions can you make from your observations?

Magnesium continues to burn in dry chlorine with a sputtering sound forming white fumes of magnesium chloride. This shows that magnesium reacts with chlorine.

\[
\text{Magnesium} + \text{chlorine} \rightarrow \text{magnesium chloride} \\
\text{Mg(s)} + \text{Cl}_2(g) \rightarrow \text{MgCl}_2(s)
\]

**Fig 6.5: Reaction of chlorine with magnesium**
- What do you observe in the gas jar?
B. Iron
1. Put a piece of iron wire in the combustion tube.
2. Arrange the apparatus as shown in Fig 6.6.

Fig. 6.6: Reaction of chlorine with iron

![Diagram of experiment setup]

The whole apparatus should be kept dry. Give a reason for this.

3. Heat the iron wire until it glows red-hot.
4. Pass dry chlorine through the combustion tube over the red-hot iron.
5. Warm the end of the combustion tube near the receiver.
   Record your observation.
   • what do you observe in the combustion tube?
   • what do you observe in the receiver?
   • why is it possible to collect the product from the combustion tube into the receiver?
   • what is your conclusion?

When dry chlorine is passed over heated iron wire, black crystals of iron(III) chloride form on the sides of the combustion tube.

\[
\text{Iron} \quad + \quad \text{chlorine} \quad \rightarrow \quad \text{iron(III) chloride}
\]

\[
2\text{Fe(s)} \quad + \quad 3\text{Cl}_2(g) \quad \rightarrow \quad 2\text{FeCl}_3(s)
\]

When the end of combustion tube near the receiver is warmed, the black crystals of iron(III) chloride sublime to brown fumes and are deposited in the receiver, where they can be collected. When exposed to air, iron(III) chloride slightly fumes. This is because the iron(III) chloride combines with water vapour in the atmosphere to form white fumes of hydrogen chloride gas. This gas dissolves in water vapour to form hydrochloric acid.

\[
\text{Iron(III) chloride} \quad + \quad \text{water} \quad \rightarrow \quad \text{iron(III) hydroxide} \quad + \quad \text{hydrochloric acid}
\]

\[
\text{FeCl}_3(s) \quad + \quad 3\text{H}_2\text{O}(l) \quad \rightarrow \quad \text{Fe(OH)}_3(s) \quad + \quad 3\text{HCl}(g)
\]

Therefore, in the preparation of iron(III) chloride, the whole apparatus should be dry.
Note: Chlorine combines with magnesium or iron to form salts. A lot of heat energy is required for these reactions to take place. This energy is provided by the burning magnesium or red-hot iron.

**Experiment 6.4**
Reaction of chlorine with phosphorus.

**Apparatus and chemicals**
- gas jar full of dry chlorine
- deflagrating spoon
- dry red phosphorous

**Caution:**
Use red phosphorous in place of white phosphorous whenever possible. Red phosphorus is less reactive hence less risky.

**Procedure**
1. Warm phosphorus in a deflagrating spoon.
2. Put the warm phosphorus in a gas jar full of dry chlorine as shown in Fig 6.7
   - What do you observe?

![Diagram of reaction](image)

*Fig 6.7: Reaction of chlorine with phosphorus*

3. Record your observations in your notebook.
   - What conclusions can you make from your observations?

When lowered into a jar full of dry chlorine, warm phosphorus catches fire. A mixture of two phosphorus chlorides are formed.

\[
P_4(s) + 6Cl_2(g) \rightarrow 4PCl_3(s)
\]
Phosphorus + chlorine → phosphorus(V) chloride
\[ P_4(s) + 10Cl_2(g) → 4PCl_5(s) \]

If cold white phosphorus is used, it catches fire spontaneously forming the same type of chlorides. Phosphorus is very reactive towards chlorine. White phosphorus is even more reactive towards chlorine than red phosphorus.

**Experiment 6.5**
Reaction of chlorine with hydrogen sulphide.

**Apparatus and chemicals**
- gas jar full of dry chlorine.
- gas jar full of dry hydrogen sulphide.

**Procedure**
1. Invert a covered gas jar of hydrogen sulphide over a covered gas jar of chlorine as shown in Fig 6.8.

   ![Fig 6.8: Reaction of chlorine and hydrogen sulphide.](image)

   \[
   \text{gas jar} \quad \rightarrow \quad \text{dry hydrogen sulphide} \quad \rightarrow \quad \text{cardboard cover} \quad \rightarrow \quad \text{dry chlorine} \quad \rightarrow \quad \text{gas jar}
   \]

2. Remove the cover.
   - What do you observe?
   - What is the colour of the solid?
3. Wet a glass rod with concentrated ammonia solution and insert it into the remaining gas in the lower gas jar to test the gas produced in the reaction.
   - What do you observe?
   - What can you conclude from your observations?

When dry hydrogen sulphide gas reacts with dry chlorine gas, a yellow solid of sulphur is seen on the sides of the lower gas jar. The hydrogen chloride gas produced in the reaction gives white fumes with ammonia gas (from concentrated ammonia solution on the glass rod) i.e.
Hydrogen + ammonia $\rightarrow$ ammonium chloride

\[ \text{HCl(g)} + \text{NH}_3(g) \rightarrow \text{NH}_4\text{Cl(s)} \] (white fumes)

Chlorine oxidises hydrogen sulphide gas to sulphur at room temperature. Chlorine is reduced by hydrogen sulphide to hydrogen chloride. Therefore, chlorine is the **oxidising agent** and hydrogen sulphide is the **reducing agent**.

\[ \text{H}_2\text{S(g)} + \text{Cl}_2(g) \rightarrow 2\text{HCl(g)} + \text{S(s)} \] (yellow solid)

• Compare the density of hydrogen sulphide with that of air, using information acquired in this experiment.

**Experiment 6.6**
Reaction of chlorine with sodium sulphite.

**Apparatus and chemicals**
- test tubes
- solution of sodium sulphite
- dilute nitric or hydrochloric acid
- barium nitrate or barium chloride solution
- chlorine water or a source of chlorine

**Procedure**
1. Acidify sodium sulphite solution in a test-tube by adding a few drops of either dilute nitric acid or hydrochloric acid to it.
2. Put a few drops of barium nitrate or barium chloride solution to the sodium sulphite solution. Observe what happens and record it in your notebook.
   • What can you conclude from this observation?
3. Shake an acidified solution of sodium sulphite with chlorine water or pass chlorine gas into it. Put a few drops of barium nitrate or barium chloride solution to it. What is your observation? Record this in your notebook.

No precipitate is formed when a few drops of barium nitrate or barium chloride solutions are added to the acidified solution of sodium sulphite.
When sodium sulphite solution is first shaken with chlorine water or chlorine gas is bubbled into it, the solution formed gives a white precipitate with barium nitrate or barium chloride solution. This test confirms that the sulphite ion (SO\(_3^{2-}\)) has been oxidised to sulphate ion (SO\(_4^{2-}\)) by chlorine.

\[
\text{oxidised}
\]

\[
\text{Na}_2\text{SO}_3(aq) + \text{Cl}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{HCl}(aq)
\]

\[
\text{reduced}
\]

Sodium + barium \rightarrow \text{barium + sodium}

\[
\text{Sulphate} + \text{nitrate} \rightarrow \text{sulphate + nitrate}
\]

\[
\text{Na}_2\text{SO}_4(aq) + \text{Ba(NO}_3)_2(aq) \rightarrow \text{BaSO}_4(s) + 2\text{NaNO}_3(aq)
\]

**Ionic equation,**

\[
\text{SO}_4^{2-}(aq) + \text{Ba}^{2+}(aq) \rightarrow \text{BaSO}_4(s)
\]

Why should sodium sulphite solution be acidified?
A white precipitate of barium sulphite is formed when barium nitrate or barium chloride solution is added to unacidified sodium sulphite.

\[
\text{Na}_2\text{SO}_3(aq) + \text{Ba(NO}_3)_2(aq) \rightarrow \text{BaSO}_3(aq) + 2\text{NaNO}_3(aq)
\]

The white precipitate dissolves in dilute hydrochloric or dilute nitric acid and the solution becomes clear. Therefore, acidifying the sodium sulphite solution with dilute nitric or hydrochloric acid helps differentiate when barium sulphite or barium sulphate formed in the reaction.

- Why is it not advisable to acidify sodium sulphite solution with sulphuric acid?

If sulphuric acid is used to acidify the sodium sulphite solution, a white precipitate of barium sulphate is formed. This is because the sulphuric acid provides sulphate ions (SO\(_4^{2-}\)) which react with barium ions (Ba\(^{2+}\)) to form barium sulphate.

\[
\text{SO}_4^{2-}(aq) + \text{Ba}^{2+}(aq) \rightarrow \text{BaSO}_4(s)
\]
Experiment 6.7
Reaction of chlorine with ammonia.

Apparatus and chemicals
- gas jar full of dry chlorine
- source of dry ammonia

Procedure
1. Pass dry ammonia gas into a gas jar containing dry chlorine gas as shown in Fig 6.9.

Fig 6.9: Reaction of chlorine with ammonia
- What do you observe?
2. Record your observations in your notebook.
- What conclusions can you make from the above observations?

When ammonia gas is mixed with chlorine gas, it catches fire and burns for some time at the end of the delivery tube. The flame goes off after a few minutes. White fumes of ammonium chloride fill the gas jar. The chemical reactions which take place are:

\[
2\text{NH}_3(g) + 3\text{Cl}_2(g) \rightarrow 6\text{HCl}(g) + \text{N}_2(g)
\]

The hydrogen chloride gas formed reacts with the unreacted ammonia to form dense white fumes of ammonium chloride.
Ammonia + hydrogen chloride $\rightarrow$ ammonium chloride
\[ \text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s) \]

More hydrogen chloride is formed only to react with the unreacted ammonia gas. The overall reaction is:

Ammonia + chlorine $\rightarrow$ ammonium + nitrogen.
\[ 8\text{NH}_3(g) + 3\text{Cl}_2(g) \rightarrow 6\text{NH}_4\text{Cl}(s) + \text{N}_2(g) \]

Here, chlorine is reduced to hydrogen chloride and ammonia is oxidised to nitrogen. Chlorine is the oxidising agent while ammonia is the reducing agent.

- Why should ammonia and chlorine gas be dry?
Hydrogen chloride gas formed is very soluble.

**Experiment 6.8**
Reaction of chlorine with water.

**Apparatus and chemicals**
- boiling tube
- water trough
- source of chlorine
- delivery tubes
- red and blue litmus papers
- wooden splint
- silver nitrate solution
- conical flask or beaker

**Procedure**
1. Prepare chlorine water by passing chlorine into a flask or beaker containing about 100 cm$^3$ of water in a fume cupboard or open space. It takes about 20 minutes to saturate this amount of water.
   - Record your observation in your notebook.
2. Dip red and then blue litmus paper into the chlorine water you have just prepared.
   - What is the effect of this water to litmus papers?
3. Fill a boiling tube with the chlorine water.
4. Invert the boiling tube with its contents in a beaker as shown in Fig 6.10.
5. Leave the experiment for about 2 days in sunlight.
6. Dip red and blue litmus papers into the resulting solution.
   • What is the effect of the solution on the litmus papers?
   • What is the colour and smell of the resulting solution?
7. Put about 2 cm$^3$ of the resulting solution in a test tube and then add a few drops of silver nitrate solution.
   • What is formed in the test tube?
8. Test the gas formed in the boiling tube with a glowing splint.
   • What happens to the glowing splint?
   • What can you conclude about reaction taking place when chlorine water is exposed in sunlight?

When chlorine is dissolved in water, some of it reacts with water to form hydrochloric acid (HCl) and hypochlorous acid (HOCl).

Chlorine water is a yellow-green solution. This water bleaches blue and red litmus papers.

After two days the gas produced re-lights a glowing splint showing that it is oxygen. The resulting solution in the beaker is colourless. It turns blue litmus paper red showing that it is acidic.

$$\text{Cl}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HCl}(aq) + \text{HOCl}(aq)$$

Hypochlorous acid is a very unstable acid and light energy from the sun decomposes it to form hydrochloric acid and an oxygen atom.

$$\text{HOCl}(aq) \rightarrow \text{HCl}(aq) + [O]$$

Since oxygen is a diatomic gas, the atoms join to form molecules.

$$[O] + [O] \rightarrow \text{O}_2(g)$$
Overall equation

Hypochlorous acid $\rightarrow$ hydrochloric acid + oxygen

\[ 2\text{HOCl(aq)} \rightarrow 2\text{HCl(aq)} + \text{O}_2(g) \]

The final solution is colourless because of the formation of hydrochloric acid and oxygen. The chloride ions (Cl$^-$) from hydrochloric acid give a white precipitate of silver chloride with silver nitrate solution.

Silver nitrate + hydrochloric acid $\rightarrow$ silver chloride + nitric acid

\[ \text{AgNO}_3(aq) + \text{HCl(aq)} \rightarrow \text{AgCl(s)} + \text{HNO}_3(aq) \]

white precipitate

Ionically; the reaction can be represented as follows:

\[ \text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl(s)} \]

white precipitate

- What do you think would happen if hypochlorous acid is heated?
  It will produce oxygen gas and hydrochloric acid will also be formed.

Experiment 6.9
Reaction of chlorine with alkali solutions.

Apparatus and chemicals
- dilute sodium hydroxide solution.
- concentrated sodium hydroxide solution
- source of chlorine.
- boiling tubes or test tubes
- combustion tube

Procedure
1. Bubble chlorine through a dilute solution of sodium hydroxide in a test tube
   - What is your observation?
2. Repeat the same with hot concentrated sodium hydroxide solution (about 70$^\circ$C).
   - What is your observation? Record your observation in your notebook.

Chlorine reacts with cold dilute sodium hydroxide solution to form a yellow solution of sodium hypochlorite and sodium chloride

\[ \text{Sodium} + \text{chlorine} \rightarrow \text{sodium hypochlorite} + \text{sodium chloride} \]

\[ \text{Sodium hydroxide} + \text{chlorine} \rightarrow \text{sodium hypochlorite} + \text{sodium chloride} + \text{water} \]
\[2\text{NaOH(aq)} + \text{Cl}_2(\text{g}) \rightarrow \text{NaOCl(aq)} + \text{NaCl(aq)} + \text{H}_2\text{O(l)}\]

Chlorine also reacts with hot concentrated sodium hydroxide solution to form sodium chlorate and sodium chloride solutions.

\[
\text{sodium hydroxide} + \text{chlorine} \rightarrow \text{sodium chlorate} + \text{sodium chloride} + \text{water}
\]

\[6\text{NaOH(aq)} + 3\text{Cl}_2(\text{g}) \rightarrow \text{NaClO}_3(\text{aq}) + 5\text{NaCl(aq)} + 3\text{H}_2\text{O(l)}\]

Dilute or concentrated alkali can be used to absorb unreacted or excess chlorine. For example in the preparation of magnesium chloride as shown in Fig 6.11.

![Fig. 6.11: Preparation of magnesium chloride](image)

Can similar reactions occur with potassium hydroxide solution?

When dilute potassium hydroxide is used instead of sodium hydroxide, potassium hypochlorite, potassium chloride and water are formed.

\[
\text{Potassium hydroxide} + \text{chlorine} \rightarrow \text{potassium hypochlorite} + \text{potassium chloride} + \text{water}
\]

\[2\text{KOH(aq)} + \text{Cl}_2(\text{aq}) \rightarrow \text{KClO(aq)} + \text{KCl(aq)} + \text{H}_2\text{O(l)}\]

When hot concentrated potassium hydroxide is used, potassium chlorate, potassium chloride and water are formed.

\[
\text{Potassium hydroxide} + \text{chlorine} \rightarrow \text{potassium chlorate} + \text{potassium chloride} + \text{water}
\]

\[6\text{KOH(aq)} + 3\text{Cl}_2(\text{aq}) \rightarrow \text{KClO}_3(\text{aq}) + 5\text{KCl(aq)} + 3\text{H}_2\text{O(l)}\]
• Sodium chlorate and sodium chloride from the reaction above have different solubilities. How can you separate them when the mixture is in solution form?

A solution of sodium chlorate and sodium chloride can be separated by fractional crystallisation.

6.4 Displacement reactions of chlorine
The trend in the oxidising power of halogens is most easily seen in their displacement reactions. Chlorine displaces bromine and iodine from solutions of their salts such as potassium iodide and potassium bromide.

Experiment 6.10
To study displacement reactions of chlorine on bromides and iodides.

Apparatus and chemicals
• test tubes
• test tube rack
• potassium chloride solution
• potassium bromide solution
• potassium iodide solution
• chlorine water or source of chlorine gas.

Procedure
1. Fill a test tube with about 4 cm$^3$ of potassium bromide solution and label it.
2. Fill another test tube with about 4 cm$^3$ of potassium iodide solution and label it.
3. Place the labelled test tubes on a test tube rack.
4. Add chlorine water or bubble chlorine gas through potassium bromide and potassium iodide solutions. If you are using chlorine water, shake the mixture well.
5. Record your observations in your notebook.
   • Describe the colour changes in the potassium bromide and potassium iodide solutions.
   • What is the colour of the solid deposited, if any?
   • What can you conclude from the above observations?

When chlorine water or chlorine gas is added to colourless potassium bromide solution, the solution turns red due to the formation of bromine water. Chlorine oxidises bromide ions to bromine. Chlorine gas is reduced to chloride ions.

$$\text{Potassium bromide} + \text{chlorine} \rightarrow \text{potassium chloride} + \text{bromine}$$

$$\text{2KBr(aq)} + \text{Cl}_2(\text{g}) \rightarrow \text{2KCl(aq)} + \text{Br}_2(\text{l})$$
Ionically, the reaction between potassium bromide and chlorine can be represented as follows:

$$2\text{Br}^-(aq) + \text{Cl}_2(g) \rightarrow 2\text{Cl}^-(aq) + \text{Br}_2(l)$$

Chlorine is more reactive than bromine hence it displaces bromine from potassium bromide solution.

Similarly chlorine displaces iodide ions from potassium iodide solution. The formation of iodine is seen when the clear solution of potassium iodide turns dark brown (a characteristic of iodine colour) and finally a black solid (iodine) is deposited.

$$2\text{KI}(aq) + \text{Cl}_2(g) \rightarrow 2\text{KCl}(aq) + \text{I}_2(s)$$

Ionically:

$$2\text{I}^-(aq) + \text{Cl}_2(g) \rightarrow 2\text{Cl}^-(aq) + \text{I}_2(s)$$

Chlorine is, therefore more reactive than iodine. Therefore, chlorine displaces more easily than iodine.

We can also prove that bromine displaces iodide ions from potassium iodide solution. Iodine is deposited as a result.

$$2\text{KI}(aq) + \text{Br}_2(l) \rightarrow 2\text{KBr}(aq) + \text{I}_2(s)$$

Ionically:

$$2\text{I}^-(aq) + \text{Br}_2(l) \rightarrow 2\text{Br}^-(aq) + \text{I}_2(s)$$

From these observations, we can conclude that the order of reactivity is chlorine > bromine > iodine. That is, chlorine has greater displacement power than bromine whereas bromine has greater displacement power than iodine. This has been illustrated in Fig. 6.12.
• Explain the order chlorine > bromine > iodine in terms of ease of electron gain by halogens.

The ease of gain of electrons depends on the atomic radius of an atom of an element. The smaller the atomic radius the more easily an atom can gain an electron or electrons. The atomic radius increases down the group and therefore in Group (VII), reactivity of elements decreases down the group.

6.5: Bleaching action of chlorine

Experiment 6.11
Bleaching action of chlorine

Apparatus and chemicals
• coloured flower petals
• litmus papers
• source of dry and moist chlorine.
• two gas jars

Procedure
1. Fill a gas jar with moist chlorine.
2. Fill another gas jar with dry chlorine.
3. In each gas jar, drop coloured flower petals and litmus papers.
4. Leave the set-up undisturbed for some time.
   • Record your observations in a notebook.
   • What can you conclude about the bleaching action of chlorine?
Coloured materials such as coloured flowers and litmus paper contain dyes and can be bleached by chlorine. Hypochlorous acid is formed by the reaction between chlorine and water. The acid gives up its oxygen atoms to these dye molecules. In the process the flower petals and litmus papers are bleached.

\[
\text{Water} + \text{chlorine} \rightarrow \text{hydrochloric acid} + \text{hypochlorous acid}
\]

\[
\text{H}_2\text{O}(l) + \text{Cl}_2(g) \rightarrow \text{HCl}(aq) + \text{HOCl}(aq)
\]

\[
dye + \text{HOCl}(aq) \rightarrow \text{HCl}(aq) + (\text{dye} + \text{O}) \text{colourless}
\]

Water must be present before any bleaching by chlorine can take place. If absent no hypochlorous acid can be formed and therefore the dye will not be bleached.

- What would happen if a wet newsprint is inserted into a gas jar of chlorine?
  The newsprint is not affected because writings on it are not dyes.
- Explain why bleaching by chlorine cannot occur in the dark despite the presence of water.
  Energy is required to release oxygen atoms from hypochlorous acid (HOCl). Light provides this energy.

**Household bleaching agents**

Household bleaching agents such as *jik*, *topex* and *milton* contain hypochlorous acid or sodium hypochlorite.

Sodium hypochlorite (NaOCl) contains active oxygen for bleaching like hypochlorous acid (HOCl). The bleaching agents that contain sodium hypochlorite bleach in the following manner:

\[
\text{Sodium hypochlorite} + \text{dye} \rightarrow \text{sodium chloride} + (\text{dye} + \text{O})
\]

\[
\text{NaOCl}(aq) + \text{dye} \rightarrow \text{NaCl}(aq) + (\text{dye} + \text{O}) \text{colourless}
\]

The ionic equation for the reaction between bleaching agents and dyes is show below.

\[
\text{OCl}^-\text{(aq)} + \text{dye} \rightarrow \text{Cl}^-\text{(aq)} + (\text{dye} + \text{O}) \text{colourless}
\]

The bleaching agent oxidises the dye.

Calcium hypochlorite (bleaching powder), CaOCl₂, is also a powerful bleaching agent.

- Explain how calcium hypochlorite acts as a bleaching agent.
Note
- Calcium hypochlorite can also be used for bleaching.
- Calcium hypochlorite is very sensitive to acids. It can even react with carbonic acid present in the atmosphere to form calcium carbonate, water and chlorine.
- When a cloth is bleached, it must be thoroughly washed afterwards to remove the hydrochloric acid produced during bleaching process, which if left will attack the cloth.

6.6 Test for chlorides
The chlorides can be in solid form or in aqueous solution.

Experiment 6.12
Test for chlorides in solid form.

Apparatus and chemicals
- dry test tube
- concentrated sulphuric acid
- concentrated ammonia solution
- dilute nitric acid
- silver nitrate solution
- glass rod
- any solid chloride.

Procedure
1. Put the solid chloride sample in a dry test tube.
2. Add concentrated sulphuric acid.
3. Heat the mixture.
4. Test the gas produced with a glass rod which has previously been dipped in concentrated ammonia solution or acidified silver nitrate solution.
   - What do you observe?

White fumes of ammonium chloride are formed when ammonia solution on the glass rod is brought near the mouth of the test tube. This is a test for hydrogen chloride gas.

Hydrogen chloride gas also forms a white precipitate with acidified silver nitrate solution on the glass rod. All solid chlorides produce hydrogen chloride gas when reacted with concentrated sulphuric acid. This is the test for solid chlorides.

Note: Silver nitrate should be acidified with dilute nitric acid. A precipitate of silver chloride can be confused with that of silver carbonate. When the acid is present, silver carbonate reacts with it and therefore no precipitate will be observed.
- Explain why we should not acidify silver nitrate solution with hydrochloric acid.
When acidified with hydrochloric acid, silver nitrate forms a white precipitate of silver chloride.

\[
\begin{align*}
\text{Silver} + \text{hydrochloric} & \rightarrow \text{silver} + \text{nitric} \\
\text{nitr}ate & \text{acid} \quad & \text{chloride} & \text{+ acid} \\
AgNO_3(aq) + HCl(aq) & \rightarrow AgCl(s) + HNO_3(aq) \\
\end{align*}
\]

Ionically:
\[
\begin{align*}
Ag^+(aq) + Cl^-(aq) & \rightarrow AgCl(s) \\
\text{(from hydrochloric acid)} & \text{(white precipitate)}
\end{align*}
\]

**Experiment 6.13**
Test for chlorides in aqueous solution.

**Apparatus and chemicals**
- test tubes
- chloride in aqueous solution
- dilute nitric acid
- silver nitrate solution
- lead(II) nitrate solution

**Procedure**
1. Put a chloride in aqueous solution in a test-tube.
2. Add a few drops of acidified aqueous silver nitrate to this solution.
   - What do you observe?
3. In another chloride solution in a test tube, add a few drops of lead(II) nitrate solution.
   - What do you observe?
4. If a precipitate is observed in 3 above, heat the contents in the test-tube.
   - What do you observe?

A white precipitate is formed when acidified aqueous silver nitrate is added to any soluble chloride.

Ionically:
\[
Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s) \\
\text{(white precipitate)}
\]

When lead(II) nitrate solution is added to a chloride in aqueous solution, a white precipitate which is soluble on heating is formed.
All chlorides in solution form a white precipitate with acidified silver nitrate solution. Lead(II) nitrate solution also forms a white precipitate of lead(II) chloride when reacted with chloride solution. Lead(II) chloride is soluble in hot water.

6.7: Uses of chlorine

1. Sterilization of drinking water or swimming pools. Due to its poisonous nature, chlorine kills germs. (Fig 6.13(a)).
2. In the manufacture of antiseptics. (Fig 6.13b)
3. In the manufacture of insecticides, such as DDT. (Fig 6.13c)
4. In the manufacture of herbicides, like sodium and potassium chlorates. (Fig 6.13d)
5. Manufacture of bleaching agents such as Parazone, Milton, Domestos, Jik etc. Chlorine can also be used directly as a bleach. For example, the wood pulp used at Webuye Paper Mills to make paper is initially bleached by chlorine and other bleaching agent such as sulphur(IV) oxide and calcium hydrogensulphite. (Fig 6.13 e)
6. Manufacture of fluid used in fire extinguishers such as pyrene (Fig. 6.13 f).
7. Manufacture of plastics like polychloroethene, (polyvinyl chloride, PVC) (Fig 6.13 g)
8. In the manufacture of anaesthetic such as trichloromethane (CHCl₃) (Fig 6.13 h)
9. In the manufacture of dry cleaning solvents such as carbon (IV) chloride. (Fig 6.13 i)
10. Chlorine compounds as nutrients help in digestion of food in our bodies and maintenance of water balance in body cells. They are found in, table salt and foods such as meat, milk and eggs.
Fig. 6.13: Uses of chlorine

Exercises 6.1
1. The flow chart below shows how chlorine gas is prepared and reacted with burning magnesium.

Conc acid A + KMnO₄ → moist Cl₂ → Solution F → dry Cl₂ → burning magnesium → Magnesium chloride

Excess Cl₂ → Hot conc. NaOH

(i) Identify
   (a) Acid A
   (b) Solution F
(ii) What is the function of
(a) Solution F
(b) Hot concentrated NaOH?
(iii) Write chemical equation for the reaction taking place with hot concentrated sodium hydroxide.

2. A saturated solution of chlorine water was exposed to sunlight as shown in the figure below.

(a) Name gas X
(b) Write two equations for the reactions that take place.
(c) What was the original colour of the solution?
(d) Write a chemical and ionic equations for the reaction which takes place when aqueous silver nitrate is added to the final solution.
(e) Why should silver nitrate solution be acidified with dilute nitric acid when testing for chlorides in aqueous solutions?

3. (a) If a piece of dry litmus paper is placed in dry chlorine, there is no observable change on the paper. If another litmus paper which is moist is placed in the same dry gas, it is bleached very rapidly. Explain these observations and use balanced chemical equations where possible.
(b) Why does blue moist litmus paper first turn red and then bleached by chlorine?
(c) The writings on a newsprint are not affected by moist chlorine gas. Why is this so?
(d) What is observed when chlorine water in a sealed test-tube is kept in a dark room for a week?

4. (a) Explain what happens when chlorine gas is bubbled in
   (i) Potassium bromide solution.
   (ii) Potassium iodide solution.
   Use chemical and ionic equations where possible.
(b) Arrange bromine, iodine, fluorine and chlorine in order of their displacement power, starting with the strongest element.
5. Study the flow chart below and then answer the questions that follow.

(a) Name gas X
(b) What is the use of (a) anhydrous calcium chloride in step 1?
(c) What would happen to iron(III) chloride if the air does not pass through anhydrous calcium chloride? Support your answer with a chemical equation.
(d) Write a chemical equation for the reaction in step 2.

6.8: Hydrogen chloride

Introduction
Hydrogen chloride is one of the three common hydrogen halides. In gaseous form, it is called hydrogen chloride gas. The solution of this gas in water is called hydrochloric acid.

6.9: Preparation of hydrogen chloride
Hydrogen chloride gas is usually prepared in the laboratory by heating a mixture of a solid chloride salt and concentrated sulphuric acid.

Chloride salts which form insoluble sulphate salts such as chlorides of barium, lead(II) and calcium should not be used. The precipitate will form around the surface of the compound and prevent any further reaction.

Experiment 6.14
Laboratory preparation of hydrogen chloride.

Apparatus and chemicals
- round-bottomed flask
- thistle funnel or dropper
- wash-bottle
- delivery tube
- gas jars
- glass rod
- sodium chloride
• concentrated sulphuric acid
• concentrated ammonia solution

**Procedure**
1. Put some solid sodium chloride in a round-bottomed flask. Lumps rather than powdered sodium chloride should be used because the reaction rate is moderate.
2. Arrange the apparatus as shown in Fig 6.14.

![Diagram of laboratory preparation of hydrogen chloride](image)

**Fig. 6.14: Laboratory preparation of hydrogen chloride**

3. Add moderately concentrated sulphuric acid drop by drop. What do you observe?
4. Heat the mixture gently.
   • What is the use of concentrated sulphuric acid in the wash bottle?
   • Why does concentrated sulphuric acid not distill off?
   • What is the colour and smell of the gas produced in the gas jar?
   • What is the test for this gas?
   • Record observations and conclusions in your notebook.

When the mixture of sodium chloride and concentrated sulphuric acid is heated, colourless hydrogen chloride gas which has a pungent irritating smell is produced. The gas forms white fumes of ammonium chloride with ammonia gas. This is a test for hydrogen chloride gas.

\[
\text{Sodium chloride} + \text{sulphuric acid} \rightarrow \text{sodium hydrogensulphate} + \text{hydrogen chloride}
\]

\[
\text{NaCl(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{NaHSO}_4(\text{aq}) + \text{HCl(g)}
\]

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Hydrogen chloride gas is dried by passing it through concentrated sulphuric acid. The temperature used in this experiment is not enough to produce sodium sulphate (normal salt) but only sodium hydrogensulphate (acid salt). Sulphuric acid does not distill off because the temperature used in the experiment is below the boiling point of this acid.

Collect four gas jars of hydrogen chloride gas for Experiment 6.15.

**6.10: Properties of hydrogen chloride**

**Experiment 6.15**
To determine the properties of hydrogen chloride

**Apparatus and chemicals**
- four jars of dry hydrogen chloride
- trough
- concentrated ammonia solution
- blue and red litmus papers
- dilute nitric acid
- silver nitrate solution
- glass rod

**Procedure**
1. Carry out the following experiment/tests with hydrogen chloride gas collected in Experiment 6.14. Follow the procedures as outlined in Table 6.14. Draw the table in your notebook and record your observations and conclusions.

**Table 6.1: Determination of properties of hydrogen chloride gas**

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Observations</th>
<th>Conclusions</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Note the colour and smell of the gas</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. Invert the first gas jar full of dry hydrogen chloride in a trough of water.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. Bring a glass rod which has been dipped in concentrated aqueous ammonia or dilute acidified (using HNO$_3$) silver nitrate to the mouth of the second gas jar full of dry hydrogen chloride as shown in Fig. 6.15(a) and (b)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4. Put moist blue and red litmus paper in the third gas jar full of hydrogen chloride.</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

323
5. Ignite the gas in the fourth jar with a burning splint.

6. Pour the gas from a gas jar on a burning splint or on a moist blue litmus paper.

---

**Fig. 6.15 (a) Test for hydrogen chloride with ammonia**

**Fig. 6.15 (b) Test for hydrogen chloride with silver nitrate solution**

**Physical properties of hydrogen chloride gas**

- It is a colourless gas with a pungent irritating smell.
- It is very soluble in water. The water rapidly fills the gas jar, if the jar is full of dry hydrogen chloride.
- It is denser than air. It can therefore be poured down to extinguish a burning candle or turn moist blue litmus red. It is therefore collected by downward delivery.

**Chemical properties of hydrogen chloride gas**

**Combustion**

Hydrogen chloride does not burn neither does it support burning. This is why it extinguishes a burning splint.

**Effect on litmus**

Hydrogen chloride turns moist blue litmus paper red. Therefore the gas is acidic.

**Reaction with ammonia**

Concentrated ammonia solution on a glass rod reacts with hydrogen chloride gas to form white fumes of ammonium chloride.
Ammonia + hydrogen chloride ——> ammonium chloride

\[
\begin{align*}
\text{NH}_3(g) & + \quad \text{HCl}(g) & \longrightarrow & \quad \text{NH}_4\text{Cl}(s) \\
\text{(white fumes)} & & & 
\end{align*}
\]

This is the test for hydrogen chloride gas.

- **Reaction with acidified silver nitrate solution**
  Hydrogen chloride gas forms white precipitate of silver chloride with acidified silver nitrate solution. The gas first forms hydrochloric acid with water in the silver nitrate solution then silver chloride and nitric acid.

\[
\begin{align*}
\text{Hydrogen chloride} & \quad \text{water} & \quad \text{hydrochloric acid} \\
\text{HCl}(g) & \quad \text{water} & \quad \text{HCl(aq)} \\
\text{Silver nitrate} & + \quad \text{hydrochloric acid} & \longrightarrow & \quad \text{silver chloride} + \quad \text{nitric acid} \\
\text{AgNO}_3(aq) & + \quad \text{HCl(aq)} & \longrightarrow & \quad \text{AgCl(s)} \quad + \quad \text{HNO}_3(aq)
\end{align*}
\]

- Do you remember why silver nitrate solution is acidified with dilute nitric acid?

**6.11: Effect of solvent on the properties of hydrogen chloride**

**Experiment 6.16**
To determine the properties of hydrogen chloride gas when dissolved in water or in methylbenzene

**Apparatus and chemicals**
- two 500 cm\(^3\) beakers
- a filter funnel
- wooden splint
- delivery tube
- source of hydrogen chloride
- fourteen test tubes
- dry methylbenzene
- a magnesium ribbon
- a carbonate salt
- a hydrogencarbonate salt
- an alkali (soluble base)
- an insoluble base
- potassium manganate(VII)
- phenolphthalein indicator
- teat pipettes
- aqueous calcium hydroxide
Procedure
1. Prepare aqueous hydrogen chloride by passing hydrogen chloride gas into water in a beaker using a filter funnel. See Fig 6.16.

![Diagram of preparing hydrogen chloride solution in water]

Fig 6.16: Preparation of hydrogen chloride solution in water

2. Label this solution X.
3. Repeat procedure 1 and 2 using dry methylbenzene instead of water. The hydrogen chloride gas should be completely dry.
4. Label the solution produced in 3 above as Y.
5. Use seven test tubes and to each add about 10 cm$^3$ of solution X.
6. Using other seven test tubes, add to each about 10 cm$^3$ of solution Y.
7. Carry out the tests as shown in Table 6.2 below. Copy Table 6.2 below in your note book and record your observations and conclusions.

<table>
<thead>
<tr>
<th>Tests</th>
<th>Observation in Solution X</th>
<th>Observation in Solution Y</th>
<th>Conclusion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Put blue litmus paper to X and Y</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Add a small piece of magnesium ribbon to X and Y.</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Add a small amount of a carbonate or hydrogen carbonate to X and Y.</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Add a small amount of potassium manganate(VII) to X and Y.</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Put 2-3 drops of phenolphthalein indicator in two separate test tubes containing aqueous sodium hydroxide. Add X into the first test tube and Y into the second test tube drop by drop.</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Place copper(II) oxide to X and Y.</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Table 6.2: Comparison of reactions of hydrogen chloride dissolved in water and in methylbenzene
• Identify the gases produced.
• What are your conclusions about reactions of hydrogen chloride dissolved in water or methylbenzene?
  What are the reasons for the above conclusions?

When dissolved in water, hydrogen chloride gas forms hydrochloric acid which ionises to form hydrogen ions (H⁺) and chloride ions (Cl⁻)

\[
\text{HCl}(g) \xrightarrow{\text{water}} \text{H}^+(aq) + \text{Cl}^-(aq)
\]

These hydrogen ions give aqueous hydrochloric acid acidic properties
When hydrogen chloride gas is dissolved in organic solvents like methylbenzene, it does not ionise. The gas retains the properties of a covalent compound as shown in Fig. 6.17 below.

**Fig. 6.17: Solvent effect on properties of hydrogen chloride**

**Hydrochloric acid:**
• has a sour taste and turns blue litmus paper red.
• reacts with some metals to give a salt and hydrogen gas. For example:

\[
\text{Mg}(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g)
\]

• reacts with carbonates and hydrogen carbonates to give a salt, water and carbon(IV) oxide. For example

\[
\text{Na}_2\text{CO}_3(s) + 2\text{HCl}(aq) \rightarrow 2\text{NaCl}(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)
\]

Carbon(IV) oxide forms a white precipitate with aqueous calcium hydroxide. i.e.

\[
\text{Ca(OH)}_2(aq) + \text{CO}_2(g) \rightarrow \text{CaCO}_3(s) + \text{H}_2\text{O}(l)
\]

• reacts with strong oxidising agents like potassium manganate(VII) without heating to give chlorine gas. Chlorine gas bleaches moist litmus papers.
• neutralises alkalis (soluble bases) to form a salt and water only. For example:

\[
\text{Sodium hydroxide + hydrochloric acid} \rightarrow \text{sodium chloride + water}
\]

\[
\text{NaOH(aq) + HCl(aq) } \rightarrow \text{NaCl(aq) + H}_2\text{O(l)}
\]

\[
\text{Potassium hydroxide + hydrochloric acid} \rightarrow \text{potassium chloride + water}
\]

\[
\text{KOH(aq) + HCl(aq) } \rightarrow \text{KCl(aq) + H}_2\text{O(l)}
\]

• reacts with insoluble basic oxides to form a salt and water only. For example black copper(II) oxide reacts with hydrochloric acid to form a green solution of copper(II) chloride.

\[
\text{Copper(II) oxide + hydrochloric acid} \rightarrow \text{Copper(II) chloride + water}
\]

\[
\text{CuO(s) + 2HCl(aq) } \rightarrow \text{CuCl}_2(aq) + \text{H}_2\text{O(l)}
\]

Like in the preparation of aqueous ammonia, a filter funnel is used when dissolving hydrogen chloride gas in water. This is to prevent “sucking back” of water from the beaker to the preparation flask. This preparation flask is hot and contains hot concentrated sulphuric acid. Water can react explosively with the hot acid.

• Do you remember how we can test for hydrogen chloride gas?

6.12: Uses of hydrogen chloride gas
The major use of this gas is in the manufacture of hydrochloric acid.

6.13: Industrial manufacture of hydrochloric acid

Raw materials
Hydrogen chloride gas is obtained by direct combination of hydrogen and chlorine. Hydrogen and chlorine are obtained by the electrolysis of concentrated aqueous sodium chloride (brine).

Reactions taking place
When hydrogen is burnt in chlorine gas, hydrogen chloride is formed.

\[
\text{Hydrogen + chlorine} \rightarrow \text{hydrogen chloride}
\]

\[
\text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2\text{HCl(g)}
\]

The reaction can be explosive and therefore it is controlled by allowing a small quantity of hydrogen through a jet to burn in excess chlorine.
The hydrogen chloride formed is then dissolved in water over glass beads to form hydrochloric acid.

\[
\text{HCl(g) } \xrightarrow{\text{water}} \text{HCl(aq)}
\]

The glass beads increase the surface area over which absorption takes place as shown in Fig. 6.18 below.

**Fig 6.18: Manufacture of hydrochloric acid**

**Industrial manufacture of hydrochloric acid at Pan African Paper, Webuye**

The process of manufacture of hydrochloric acid at Pan Paper Webuye in Kenya is the same as the one described above. The process of manufacture of hydrochloric acid at Pan Africa Paper Webuye is as follows.

Salt from Magadi soda is put into a saturator basin where brine solution is made. The concentrated salt solution is pumped to raw brine tank where sedimentation takes place. Sedimentation results in formation of a filtrate and mud which settles down. The mud is released and the filtrate is left behind in the tank. The temperature of the brine is raised to 60–70°C. The solution overflows to a neutralizer. At the neutralizer, the brine is heated and neutralized with hydrochloric acid to a pH of 6.8 – 8.5.

The brine solution from the neutralizer is then pumped to the cells arranged in series and connected through copper cathodes. There are also metal anodes that are fixed to the bottom of each cell. Electrolysis of brine takes place in the cells.

The ions that are discharged are shown in the following equations.

At the Anode: \[2\text{Cl}^–(\text{aq}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^–\]

At the Cathode: \[2\text{H}^+(\text{aq}) + 2\text{e} \rightarrow \text{H}_2(\text{g})\]
Note: The (Na⁺) ions and hydroxide (OH⁻) remain in solution as sodium hydroxide. During electrolysis chlorine gas will be released at the top of the cell, sodium hydroxide percolates down and hydrogen gas is also liberated at top of the cell.

The sodium hydroxide is pumped into storage tanks. The hot chlorine liberated is cooled using water to about 40–50°C then passed through a drying chamber where it is dried by concentrated sulphuric acid. The dried chlorine gas is diverted to a compression chamber where the gas is compressed.

The hydrogen from the cell is cooled through hydrogen coolers. The cooled hydrogen is sucked by a hydrogen blower to the synthesis unit. Here, hydrogen is lighted and injected into the combustion chamber. Chlorine is then introduced into the same combustion chamber. Hydrogen burns in the presence of chlorine to form hydrogen chloride fumes. The fumes are passed through the absorber where water is added depending on the concentration. The process is as illustrated in Fig. 6.19 below.

Fig. 6.19: Manufacture of hydrochloric acid at Pan Paper, Webuye

The acid formed is stored or transported in steel tanks which are lined inside with rubber. If the acid comes into contact with a part of the steel or rust, iron (II) chloride which is later oxidised by air to iron(III) chloride is formed. This may make the acid to have a yellow colour.

At Pan Paper Mill Webuye, all the products of electrolysis of aqueous sodium chloride find use in paper manufacturing process.

The products are sodium hydroxide, chlorine and hydrochloric acid. Chlorine is used for bleaching the wood pulp and hydrochloric acid is used to control the pH of aqueous sodium hydroxide. However most of the acid is sold to other industries which manufacture chemical products.
6.14: Uses of hydrochloric acid
1. To standardize pH of beers and wines (Fig. 6.20a)
2. In the pickling of metals (Fig. 6.20b)
3. In the manufacture of dyes, drugs etc (Fig. 6.20c)
4. In the manufacture of photographic materials (Fig. 6.20d)
5. In school and research laboratories. (Fig. 6.20e)

![Figure 6.20: Uses of hydrochloric acid](image)

6.15 Pollution of the environment by chlorine and its compounds

1. Chlorine
   We said earlier that chlorine is useful in sterilizing drinking water and water in swimming pools, manufacturing antiseptics and insecticides, among other uses. Chlorine has some bad effects to the environment. Chlorine was used as the first chemical weapon in the first world war that killed millions of people. It was used in Tanzania during the First World War (1914-1918).

2. Chlorofluorohydrocarbons (CFCs)
   CFCs are nontoxic, non-flammable chemicals which are almost inert. A CFC molecule contains atoms of carbon, chlorine and flourine. CFCs are used in the manufacture of aerosol sprays, foams (eg. mattresses), coolants and solvents.
CFCs are safe to use in most appliances. However, they are harmful when released into the atmosphere.

CFCs are harmful because they destroy the ozone layer ($O_3$) that protects us from the harmful ultra violet (uv) rays from the sun. Ozone is found in the upper part of the atmosphere (10 to 50km) above the earth’s surface.

When CFCs get to the atmosphere, the chlorine atom (Cl) is released from a CFC molecule. The chlorine atom split ozone to oxygen atoms and oxygen molecule. The oxygen atom combines with oxygen ozone molecule ($O_3$) to form an oxygen molecule, $O_2$. This results in depletion of the ozone, $O_3$, layer.

\[
O_3(g) \xrightleftharpoons[Cl]{\text{chlorine atom}} O_2(g) + O(g)
\]

\[
O_3(g) + (O) \rightarrow 2O_2(g)
\]

With destruction of the ozone layer more harmful ultra-violet (U.V) rays now reach the earth’s surface. On the ground, even a low concentration of these rays are harmful Fig. 6.21(a) and 6.21(b).

Ultra-violet rays are responsible for some skin cancers and eye cataracts. They damage the lungs and reduce crop yields.

- **Sources of CFCs in the atmosphere**
  CFCs are released to the atmosphere during the manufacture and disposal of substances such as:
  (i) Coolants that contain (CFCs) used in refrigerators and air conditioners.
  (ii) Certain aerosols where CFCs are used as propellants.
(iii) Foam containing CFC’s used in the manufacture of mattresses among other things.

- **Measures to prevent depletion of ozone layer**
  It has been said that prevention is better than cure. The first step in preventing the depletion of ozone layer is to substitute the chemicals that accelerates the decomposition of ozone with ozone friendly substances. The battle to save the ozone layer is not just an environmental issue. It now involves the complex politics of development. Plans of action papers e.g. (Montreal Protocol, Vienna 1995) will come to nothing if developed countries do not provide the promised funds to help developing nations to switch to safer substances.

3. **Dichloro-diphenyl-trichloroethane (DDT)**
DDT is widely used in most developing countries as an insecticide in the control of malaria. It is effective against mosquitoes and other insects. However, DDT is harmful to animals and human beings, especially when sprayed to control mosquitoes. This happens when DDT is washed down into water bodies. It is absorbed by tiny sea plants and animals called *plankton* which are eaten by fish. When fish is eaten, the DDT is transferred into the bodies of animals and human beings. DDT has made the egg shells of some birds to become thin and so fragile that they break during incubation.

- **Alternatives to DDT**
  There are safer and locally produced natural products (pyrethroids) which are very effective against mosquitoes. For example, pyrethrum flowers are used by the Pyrethrum Board of Kenya to produce insecticides and larvicides (chemicals that kill larvae). These include:
  (i) *Pylarva* which kills adult mosquitoes larvae and pupae in their breeding habitats.
  (ii) *Pymos* for aerial and indoor residual spraying to kill adult mosquitoes.
  (iii) *Pynet* for treating mosquito nets, kills and repels adult mosquitoes.
These products have been tested and found to be extremely effective against adult, larvae and pupal stages of mosquitoes as shown in Fig. 6.22(a)

![Fig. 6.22(a) Spraying mosquitoes, with CFC free insecticides](image-url)
Pyrethrum compounds are easily decomposed. As a country which produces excessive pyrethrum flowers, we can use this crop to save our environment; while at the same time raising the living standards of pyrethrum farmers who are poor as shown in Fig. 6.22(b) below.

**Fig. 6.22(b) Picking pyrethrum flowers**

**Exercises 6.2**
1. Study the following flow chart and answer the questions that follow:

   ![Flow chart](image)

   (a) Name:
      (i) Gas X
      (ii) Solution Y
      (iii) Gas Q

   (b) Name the substance used in drying gas X

   (c) Name the method used in collecting dry:
      (i) Gas X
      (ii) Gas Q

   (d) In the preparation of gas X, sodium sulphate is not formed. Give a reason for this.

2. (a) What is observed when moist hydrogen chloride gas is bubbled in aqueous silver nitrate in a test tube?
Write the chemical equation for the reaction in 2(i).
Write the ionic equation for the reaction in 2(i).
Silver chloride and silver carbonate are insoluble in water. How can you differentiate them?

3. (a) Explain why a solution of hydrogen chloride in methlybenzene has no effect on carbonates.
(b) Which gas is produced when aqueous hydrogen chloride is reacted with
   (i) a hydrogencarbonate salt.
   (ii) potassium manganate(VII).
(c) Give the test of each gas.
(d) In the manufacture of hydrochloric acid name the two gases which are used.
(e) Write the chemical equations of the reactions taking place in 3 (d) above.
(f) In the manufacture of hydrochloric acid, the gases are dissolved in water over glass beads. What is the use of the beads?
(g) In the laboratory preparation of hydrogen chloride using concentrated sulphuric acid, it is not advisable to use solid barium chloride or lead(II) chloride or calcium chloride. Explain why.

Projects
1. Determine the chlorine content of various bleaching powders and liquids in the market.
2. Write a report on the active bleaching ingredients in various bleaching powders and liquids in the market.
3. Visit a water purification and treatment plant. Discuss with the chemist or the laboratory technologist/technician about the activities which take place here.
4. Visit Pan Paper Mill, Webuye. Discuss with the chemist how hydrochloric acid is manufactured, its various uses and the environmental impact in the surrounding areas.

Summary
1. Chlorine is prepared in the laboratory by the oxidation of concentrated hydrochloric acid using either potassium manganate(VII) (no heating) or manganese (IV) oxide (heating required).
2. Chlorine is a very poisonous gas. Its preparation and determination of its properties should be carried out in a fume cupboard or an open place.
3. The oxidising agents provide oxygen atom to hydrochloric acid to form chlorine and water.
4. When chlorine water (a solution of chlorine in water) is exposed to sunlight, hydrochloric acid and oxygen gas are formed.
5. Hypochlorous acid is unstable and decomposes in sunlight to form oxygen and the more stable hydrochloric acid.
6. Chlorine bleaches only in the presence of water. The oxygen atom from the decomposition of hypochlorous acid oxidizes the dye to a colourless substance.
7. If there is no dye, two oxygen atoms from the decomposition of hypochlorous acid combine to form oxygen molecule.
8. The oxygen atoms from the hypochlorous acid are harmful to all kinds of life. Therefore chlorine is used to kill bacteria.
9. In most of its chemical reactions chlorine acts as an oxidising agent.
10. Chlorine displaces bromine and iodine from their aqueous salt solutions.
11. From experimental evidence, the order of oxidising or displacement power for halogens is \( F > Cl > Br > I \)
12. The laboratory preparation of \( \text{HCl}(g) \) is by the reaction a chloride salt and concentrated sulphuric acid. Sodium chloride is usually used.
13. Hydrogen chloride does not form ions in non-polar solvents such as methylbenzene. In water it forms hydrogen ions (\( \text{H}^+ \)) which give its aqueous solution acidic properties.
14. In the industrial manufacture of hydrochloric acid, hydrogen chloride gas formed from direct combination of hydrogen and chlorine is dissolved in water.
15. In the manufacture of this acid, great care should be taken to prevent chlorine, hydrogen chloride and hydrogen from escaping to the atmosphere. Care is taken because chlorine is very poisonous.

**Revision Exercises 6**

1. \( \text{MnO}_2(s) + \text{HCl(aq)} \) → Gas X
   
   (a) Name gas X.
   (b) What is the use of water?
   (c) What is the use of concentrated sulphuric acid?

2. (a) Which gas is produced when sodium chloride and manganese(IV) oxide mixture is heated with concentrated sulphuric acid.
   (b) In the production of the gas in 2(i) name two chemical reactions which takes place. Use chemical equations to support your answer.

3. What is observed when:
   (a) burning magnesium
   (b) red hot iron are separately reacted with dry chlorine.

4. When excess chlorine is bubbled through dilute sodium hydroxide, the resulting solution acts as a bleaching agent.
(a) Write an equation for the reaction between chlorine and sodium hydroxide solution.
(b) Explain how the resulting solution acts as a bleaching agent.

5. Use the flow chart below to answer the questions that follow

5. Use the flow chart below to answer the questions that follow

![Flow Chart]

(a) Name Gas A and C
(b) Name precipitate D
(c) Name the solid which can be heated with B to produce gas C.

6. Explain why hydrogen chloride gas dissolved in water liberates carbon(IV) oxide with carbonates yet when it is dissolved in methylbenzene, it does not.

7. (a) What is the importance of the ozone layer to mankind?
(b) Explain how chlorofluorohydrocarbons (CFCs) affect the ozone layer.
Paper 1

1. The diagrams below represent three methods for collecting gas.

![Diagrams](a) (b) (c)

Explain which method is best for collecting ammonia gas.

2. Draw an electron dot (.) and cross (x) diagram to show bonding in a water molecule.

3. Study the diagram below and answer the questions that follow.

![Diagram](cotton wool soaked in conc hydrochloric acid, white ring formed after sometime, cotton wool soaked in conc ammonia solution)

(a) Briefly explain why the white ring forms and why it forms at a point closer to the source of hydrogen chloride.

(b) Write an equation for the reaction which produces the white ring.

4. When 0.288 g of an oxide of metal M was reduced using a suitable reducing agent, 0.256g of the pure metal was formed. Determine the empirical formula of the metal oxide. (Relative atomic mass of M = 64)

5. Calculate the relative atomic mass of an element whose isotopic masses and relative abundance are given below.

<table>
<thead>
<tr>
<th>Relative Abundance (%)</th>
<th>Isotopic Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>31</td>
<td>65</td>
</tr>
<tr>
<td>69</td>
<td>63</td>
</tr>
</tbody>
</table>
6. A chromatogram of acid enzymes X and Y and three simple sugars are shown below.

Which two simple sugars must be present in X and Y.

7. Use the information below and answer the questions that follow.

<table>
<thead>
<tr>
<th>Solution</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>S</td>
<td>1.5</td>
</tr>
<tr>
<td>T</td>
<td>7</td>
</tr>
<tr>
<td>V</td>
<td>9.7</td>
</tr>
</tbody>
</table>

Which of the above solutions will not affect the colour of the blue litmus paper. Explain.

8. 20 cm$^3$ of carbon (II) oxide is mixed with 25 cm$^3$ of oxygen, measured at the same conditions of temperature and pressure and then the mixture is ignited. Determine the excess gas and calculate the volume of the mixture when cooled to original conditions.

$$2\text{CO(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{CO}_2\text{(g)}$$

9. When a piece of burning magnesium ribbon is lowered into a gas jar of carbon (IV) oxide gas, it continues to burn for sometime.
   (a) State and explain the observations.
   (b) Write an equation for the reaction.

10. Study the flow chart below and answer the questions that follow.
(a) Identify the two ions present in solid X.
(b) Write down the ionic equation for the reaction that takes place during the formation of the white precipitate.

11. Hydrogen sulphide gas is slightly soluble in water. The reaction is given by the equation below:

\[ \text{H}_2\text{S}(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{HS}^-(\text{aq}) \]

(a) What does the equation above show about the behaviour of hydrogen sulphide gas when dissolved in water.
(b) Write down an equation for the reaction between a little solution of hydrogen sulphide an excess potassium hydroxide solution.

12. When chlorine gas is bubbled through dilute sodium hydroxide solution, the resulting solution acts as a bleaching agent.
(a) Write an equation for the reaction between chlorine gas and sodium hydroxide solution.
(b) Explain how the resulting solution acts as a bleaching agent.

13. (a) Explain briefly why anhydrous iron(III) chloride cannot be prepared in the laboratory by reacting chlorine gas with aqueous iron(II) chloride solution.
(b) Write an equation for the reaction that you would carry out to prepare anhydrous iron(III) chloride.

14. An unknown mass of anhydrous sodium carbonate was dissolved in water and the solution made up to 250 cm\textsuperscript{3}. 25.0 cm\textsuperscript{3} of this solution neutralized 20.0 cm\textsuperscript{3} of 0.25 M nitric acid. Calculate the unknown mass of sodium carbonate.

15. (a) Diamond and graphite are both allotropes of carbon, explain why graphite is used as a lubricant whereas diamond is used as an abrasive.
(b) State two uses of carbon(II) oxide

16. The table below gives the properties of compounds A, B, C and D.

<table>
<thead>
<tr>
<th>Compound</th>
<th>M. P (\degree C)</th>
<th>B.P\textsuperscript{\circ}C</th>
<th>Conductivity in water</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>–23</td>
<td>77</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>B</td>
<td>–19</td>
<td>74</td>
<td>Does not conduct</td>
</tr>
<tr>
<td>C</td>
<td>–85</td>
<td>–61</td>
<td>Conducts</td>
</tr>
<tr>
<td>D</td>
<td>714</td>
<td>1407</td>
<td>Conducts</td>
</tr>
</tbody>
</table>

(a) Which one of the compounds in the table is ionic? Explain.
(b) Give the compound(s) that is/are liquid(s) at room temperature.

17. Study the flow chart below based on pure solid compound Q and water. Answer the questions that follow.
Identify the cation and anion present in compound Q.

(b) What would be observed if compound Q was strongly heated in an ignition tube?

18. (a) From the following equation write down an ionic equation.

\[ \text{CuSO}_4(\text{aq}) + \text{Fe}(s) \rightarrow \text{Cu}(s) + \text{FeSO}_4(\text{aq}) \]

(b) Identify the substance that has been oxidized and the one that has been reduced.

19. A gaseous hydrocarbon CH has the following composition. 82.8% carbon, 17.2 hydrogen and has a molecular mass of 58. Determine the molecular formula of the hydrocarbon.

20. An organic compound Z reacts with hydrogen iodide to form another compound with the formula.

\[
\begin{align*}
\text{H} & \quad \text{H} & \quad \text{I} & \quad \text{I} \\
\text{H} \quad \text{C} \quad \text{C} \quad \text{C} \quad \text{C} \quad \text{H} \\
\text{H} & \quad \text{H} & \quad \text{H} & \quad \text{H}
\end{align*}
\]

(a) Name the compound Z.

(b) Write a balanced chemical equation to show how the reaction occurs.

---

**Paper 2**

1. Melinda and Eddy carried out an experiment to check the concentration of some sulphuric acid being used in cleaning the body of a car. They titrated 25.0 cm³ of the acid with sodium hydroxide solution containing 1.2 mol/dm³ NaOH. They found 35.0 cm³ of the alkali were required for neutralization.
(a) Write the equation for the reaction.
(b) What should they use to measure out:
   (i) 25.0 cm$^3$ of acid?
   (ii) the alkali required to neutralize the acid completely?
(c) State, giving a named example, what they added in order to know the end point.
(d) (i) How many moles of sodium hydroxide reacted?
   (ii) How many moles of sulphuric acid did 35.0 cm$^3$ of sodium hydroxide react with?
   (iii) What is the concentration of the sulphuric acid in mole dm$^{-3}$?

2. (a) State what is observed when
   (i) crystals of iron (II) sulphate, FeSO$_4$.7H$_2$O, are strongly heated. Write equation(s) for the reaction that occur(s).
   (ii) Concentrated nitric acid is added to a solution of iron(II) sulphate.
(b) Describe one test in each case that can be carried out to show that iron (II) sulphate crystals contain
   (i) water of crystallization
   (ii) iron (II) ions.
   (iii) sulphate ions.

3. (a) Diamond and graphite are allotropes of carbon.
   (i) Draw the structure of each allotrope.
   (ii) With the help of structure and bonding, explain why:
       I Graphite is soft and slippery while diamond is hard.
       II Graphite conducts electricity while diamond does not.

4. (a) Name the following compounds.
   (i) \[ \text{CH}_3\text{CH}_2\text{CHCH}_2\text{CH}_3 \]
   (ii) \[ \text{CH}_3\text{CH} = \text{CHCH}_3 \]

(b) Write down the structures of the following compounds:
   (i) 2, 2 – dimethylpropane
   (ii) Hex–2–ene
   (iii) 2–chloropropane
(c) (i) 1 – chloroethene, Cl\(\text{CH}=\text{CH}_2\), can form a polymer.
   (ii) What is the name of the polymer.
   (iii) Write down a general structure of the polymer.
5. The equation below represents the reaction between calcium carbonate and dilute hydrochloric acid.

\[
CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)
\]

(a) What observation would be made when dilute hydrochloric acid is added to calcium carbonate.

(b) 2.0g of CaCO$_3$ are added to excess dilute hydrochloric acid.
   (i) What volume of CO$_2$(g) would be given at s.t.p?
   (ii) What is the mass of CO$_2$ gas produced.

(c) Name two factors that would affect the rate at which carbon(IV) oxide is produced.

6. (a) Study the flow chart and answer the question that follow

   (i) Identify solids U and W.
   (ii) Write an equation to show how the colourless gas forms a white precipitate with calcium hydroxide solution.

(b) 20.0 cm$^3$ of sodium hydroxide solution containing 8.0 g dm$^{-3}$ were required for complete neutralisation of 0.18 g of a dibasic acid, H$_2$X, calculate the relative molecular mass of the acid.
   Explain why it is not suitable to prepare carbon(IV) oxide gas using sulphuric acid and calcium carbonate.

**Paper 3 (Practical)**

David and Valentine were provided with:
- A aqueous hydrochloric acid solution in a beaker, solution X.
- Solution Y containing 4.9 g of H$_2$SO$_4$ per litre.
- A aqueous sodium hydroxide, solution Z.
- Phenolphthalein indicator
They were required to:
• Standardize the sodium hydroxide solution Z.
• Use the standardised solution Z to determine the concentration of solution X.

**Procedure 1**
David and Valentine were to fill the burette with solution Y. Pipette 25.0 cm$^3$ of solution Z and transfer into a conical flask. They were to titrate using phenolphthalein indicator and record their results. Their results are shown below.

<table>
<thead>
<tr>
<th>Titration number</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading volume (cm$^3$)</td>
<td>32.20</td>
<td>25.0</td>
<td>35.10</td>
</tr>
<tr>
<td>Initial burette reading volume (cm$^3$)</td>
<td>5.00</td>
<td>0</td>
<td>10.00</td>
</tr>
<tr>
<td>Volume of solution Y used (cm$^3$)</td>
<td>27.20</td>
<td>25.0</td>
<td>25.10</td>
</tr>
</tbody>
</table>

(a) Point out the mistakes David and Valentine made and record the correct data as shown in table 1 above.
(b) They calculated their average volume of solution Y as shown below.

$$\frac{(27.20 + 25.0 + 25.10) \text{ cm}^3}{3} = 25.77\text{ cm}^3$$

Do you agree with the above average volume? Yes or no.

(i) If your answer is no, explain why?
(ii) If your answer is no, calculate the correct average volume of solution Y. (Show how you arrive at your answer)
(c) Calculate the concentration of the solution Y in mol/litre.
(d) Calculate the number of moles in Y used.
(e) Calculate the number of moles of solution Z in the above experiment that were present in 25 cm$^3$.
(f) Calculate the concentration of solution Z in mol/litre.

**Procedure II**
David and Valentine then used a measuring cylinder to measure 90 cm$^3$ of distilled water and placed it into 250 cm$^3$ beaker. They then added 10 cm$^3$ of solution X. Using a measuring cylinder, they mixed the solution well and labelled it as solution Q. They then filled the burette with solution Q and pipetted 25.0 cm$^3$ of solution Z
and transferred into a conical flask. They titrated using phenolphthalein indicator and recorded their results as shown below. Volume of pipette used 10 cm$^3$.

<table>
<thead>
<tr>
<th>Titration Number</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading (cm$^3$)</td>
<td>15.0</td>
<td>28.5</td>
<td>41.5</td>
</tr>
<tr>
<td>Initial burette reading (cm$^3$)</td>
<td>0</td>
<td>16.0</td>
<td>29.0</td>
</tr>
<tr>
<td>Volume of solution Q used (cm$^3$)</td>
<td>15</td>
<td>12.5</td>
<td>12.5</td>
</tr>
</tbody>
</table>

(a) Point out the mistakes in the data recorded above. Give reasons for every mistake you point out.
(b) Calculate the average of solution Q used in cm$^3$ (show your working)
(c) (i) Calculate the number of moles of sodium hydroxide in solution Z.
      (ii) Calculate the moles of the diluted solution and the moles that reacted completely with 25 cm$^3$ of NaOH.
(d) Calculate the concentration of the original HCl solution X in moles per litre.
Test paper two

Paper 1

1. (a) In an experiment, metal L reacts with cold water while K does not react with water at all. Metal K reduces the oxide of J. What is the order of reactivity of the three metals starting with the most reactive first?
   (b) A salt H dissolves easily in water and conducts electricity in both aqueous and molten state. What type of bonding exists in H.

2. Two liquids X and Y mix in all proportions. How can the two liquids be separated.

3. The formula of the bromide of X is XBr₄. What is the formula of the oxide of X.

4. A liquid is added dropwise to 10 cm³ of household ammonia solution. The pH value is noted after the addition of every 10 drops and plotted as on the graph below.

   From the evidence on the graph make the most reasonable deductions about the nature of the liquid added. Explain your answer.

5. (i) Describe the nature of solid X.
   (ii) Give a chemical test for the colourless liquid.
6. Both graphite and molten sodium chloride conduct electricity. State how each of these substances conduct electricity.
   (a) graphite
   (b) molten sodium chloride

7. A dry gas X was passed over heated lead (II) oxide. A grey residue, and gas Z were formed. Gas Z has no effect on red litmus paper and does not support combustion. Which gas is (a) X (b) Z


9. Below is part of the periodic table with elements shown with letters. The letters used are not the actual chemical symbols. Use the letters to answer the following questions.

   (a) Which one of the element will form a divalent anion?
   (b) Which one of the metallic elements will react most vigorously with water?
   (c) Write the formula of the compound formed when G reacts with D.

10. (a) Briefly state what you understand by the term ‘drug abuse’.
     (b) How does alcoholism affect family members and the community at large?

11. When chlorine gas is bubbled through a solution of potassium iodide, a reaction occurs. Explain the observations made and write the ionic equation.

12. [Diagram of a reaction setup with labels: copper(II) oxide, anhydrous copper(II) sulphate, hydrogen flame, heat, and reaction products.]
(i) What changes are observed above in copper (II) oxide and anhydrous copper (II) sulphate. State and explain these changes.

(ii) With the help of chemical equations, show how hydrogen and copper (II) oxide have changed.

13. (a) Name two factors that accelerate rusting.
(b) What is the chemical name of rust?
(c) Why does a car exhaust pipe rust more quickly than other parts of the car?

14. 20 cm³ of hydrochloric acid of unknown molarity reacted with 25 cm³ of 0.05M sodium carbonate. What is the molarity of the acid.

15. Ammonia gas is manufactured on large scale by reacting nitrogen gas and hydrogen gas as shown in the equation below.

\[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g}) + \text{heat} \]

What effect would:
(a) Increase in temperature have on the yield of ammonia?
(b) A catalyst have on the yield of ammonia.

16. (a) What harmful acid is produced when sulphur(IV) oxide gets into the atmosphere?
(b) List three effects of acid rain.

17. (a) Drug abuse is a very serious problem affecting both young and old people in our country and the world at large. Assume you are a peer counsellor. What piece of advice would you give your friend to prevent him/her from drug abuse.
(b) What is the effect of smoking tobacco on non-smokers?

18. Oxygen is prepared in the laboratory using hydrogen peroxide and manganese(IV) oxide. The chemical equation for the reaction is as follows:

\[ 2\text{H}_2\text{O}_2(\text{l}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) \]

(a) Explain briefly why manganese(IV) oxide is omitted in the equation above.
(b) List 3 uses of oxygen.

**Paper 2**

1. Below is a table showing some elements in the periodic table and their atomic numbers. The letters are not chemical symbols but should be used to answer the questions that follow.

<table>
<thead>
<tr>
<th></th>
<th>R</th>
<th>S</th>
<th>T</th>
<th>U</th>
<th>V</th>
<th>W</th>
<th>X</th>
<th>Y</th>
<th>Z</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td>17</td>
<td>11</td>
<td>4</td>
<td>6</td>
<td>12</td>
<td>13</td>
<td>5</td>
<td>20</td>
<td></td>
</tr>
</tbody>
</table>
(a) (i) Write the electronic arrangement of elements T and V.
(ii) To which period do T and V belong?
(b) (i) Write the formulae for the chlorides of R and X
(ii) Identify the groups in which elements R and X belong.
(c) Write the formulae of oxide of U and Y.
(d) (i) Write the formulae of the compound formed when T and S react.
(ii) Write an equation for the reaction in (d)(i) above.

2. In the preparation of metal carbonate, magnesium was burned in air and the product collected. Dilute sulphuric acid was then added to the product and the mixture filtered. Sodium carbonate was added to the filtrate and the contents filtered. The residue was washed and dried to give a white powder.
(a) Give the chemical name of the product.
(b) Write a chemical equation for the formation of the product.
(c) (i) Name the filtrate collected after sodium carbonate was added.
(ii) Name the white powder.
(d) Write a chemical equation for the reaction between the product in (a) and the acid.
(e) Name the ions present in the filtrate after adding of sodium carbonate.
(f) Write an equation to show the formation of the white powder.

3. Study the table below and answer the questions that follow.

<table>
<thead>
<tr>
<th>Formula of hydrocarbon</th>
<th>Boiling Point (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C₂H₄</td>
<td>169.4</td>
</tr>
<tr>
<td>C₃H₆</td>
<td>225.5</td>
</tr>
<tr>
<td>C₄H₈</td>
<td>288.9</td>
</tr>
<tr>
<td>C₅H₁₀</td>
<td>303.0</td>
</tr>
<tr>
<td>C₆H₁₂</td>
<td>336.5</td>
</tr>
</tbody>
</table>

(a) (i) What name is given to a series of organic compounds like the ones in the table above.
(ii) To what class of organic compounds do the above belong? Explain why?
(iii) Select the hydrocarbons that would be gases at room temperature. Give a reason for your answer.
(iv) Give the formula for the seventh member of the above series.
(v) What is the relationship between the boiling point and the relative molecular masses of the hydrocarbons above? Explain your answer.

(b) What would be observed if acidified potassium manganate(VII) was reacted with any of the above hydrocarbons. Explain.
4. The flow chart below represents the main steps in the preparation of sodium carbonate.

(a) Name the substances labelled A, B, C, D
(b) Cold water is made to circulate around X. What does this suggest about the reaction between A and brine.
(c) What process is represented by Y.
(d) Name two by-products that are recycled in this process.
(e) Why is recycling important?
(f) Write the equation for the reaction that takes place in the upper part of Solvay tower.

(g) Assuming that there was no recycling in this process, two moles of ammonia would be required for producing one mole of anhydrous sodium carbonate. Calculate the volume of ammonia at s.t.p that would be used to produce 10.6kg of sodium carbonate by a factory operating at 80% efficiency. (C = 12, O = 16, H = 1, Na = 23, N = 14 and 1 mole of gas occupies 22.4 dm$^3$ at s.t.p.
(h) Give two industrial uses of sodium carbonate.

5 (a) With the aid of a fully labelled diagram, describe briefly how dry sulphur(IV) gas can be prepared in the laboratory using copper turnings.
(b) Write the equation for the reaction in (a)
(c) Describe one test that can be used in the laboratory to identify sulphur(IV) oxide.
(d) (i) What would be observed if sulphur(IV) oxide was bubbled through aqueous hydrogen sulphide? (ii) Write an equation for the reaction.
(e) The reaction in (d) above is a reduction–oxidation reaction. Name the oxidizing agent in the reaction. Explain how it oxidizes hydrogen sulphide.

6. (a) Hydrated sodium carbonate can be used to neutralise acid HX. Write an equation for the reaction.

(b) The formula of these crystals is \( \text{Na}_2\text{CO}_3 \cdot x \text{H}_2\text{O} \), where \( x \) is a whole number. An experiment was carried out by some students to find the value of \( x \). They obtained the following results:

- Mass of dry evaporation dish = 35.47 g
- Mass of dry evaporating dish + crystals before heating = 41.24 g
- Mass of dry evaporating dish + product after heating = 37.59 g

(i) How can you show that all of the water of crystallisation is lost when the crystals were heated?

(ii) Calculate the value of \( x \)

7. (a) (i) When a hydrocarbon fuel is burned, one of the main products is a slightly acidic gas \( Q \). What is the name of the gas \( Q \)?

(ii) Give two ways in which gas \( Q \) is released into the air.

(b) Fuels often contain an impurity \( Y \). \( Y \) burns in air to form a pollutant \( Z \).

\[
\text{Y} \quad \begin{array}{c} \text{burn in} \\ \text{air} \end{array} \quad \text{Pollutant} \quad \begin{array}{c} \text{Rain} \\ \text{air} \end{array} \quad \text{Acid rain}
\]

(i) Give the name of one fuel which can cause acid rain.

(ii) What is the name of the element \( Y \).

(iii) What is the name of the compound \( Z \).

(iv) Write a balanced chemical equation for the combustion of \( Y \), showing state symbols

(v) Name the acid in acid rain.

Chemistry practical paper 3

1. You are provided with

- solution A ethane dioic acid, \( \text{H}_2\text{C}_2\text{O}_4 \cdot n\text{H}_2\text{O} \) containing 6.4 g/dm³
- solution B sodium hydroxide containing 4.0 g dm⁻³.
- phenolphthalein indicator
- All necessary apparatus for titration.

\( (H = 1.0; \ C = 12.0; \ O = 16.0; \ Na = 23.0) \)

You are required to
1. • find the concentration of solution B in mol dm$^{-3}$. Moles of the acid that reacted with sodium hydroxide.
• calculate the relative formula mass of ethane dioic acid.
• find the number of moles of water of crystalisation in one mole of ethane dioic acid i.e. the value of n in H$_2$C$_2$O$_4$. n H$_2$O.

**Procedure**

1. Fill the burette with solution A.
2. Pipette 25.0 cm$^3$ of solution B and transfer into a conical flask. Add 2 – 3 drops of the indicator.
3. Titrate solution B with solution A until the pink colour disappears.
4. Repeat the procedure two more times to obtain accurate readings and complete the table.

<table>
<thead>
<tr>
<th></th>
<th>I</th>
<th>II</th>
<th>III</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial burette reading</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Final burette reading</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Volume of solution A (cm$^3$)</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

(a) Calculate the average volume of solution A used. (show the readings you used to get your answer).
(b) Calculate the concentration of solution B in mol dm$^{-3}$
(c) Calculate the moles of the acid that reacted with the alkali, given the equation for the reaction is

\[
2\text{NaOH(aq)} + \text{H}_2\text{C}_2\text{O}_4(aq) \rightarrow \text{Na}_2\text{C}_2\text{O}_4(aq) + 2\text{H}_2\text{O(l)}
\]

(d) Find the concentration of acid, solution A, in mol dm$^{-3}$
(e) Calculate the relative formula mass of ethane dioic acid H$_2$C$_2$O$_4$. n H$_2$O
(f) What is the value of n in H$_2$C$_2$O$_4$. nH$_2$O.

2. You are provided with a solid C. You are required to carry out the tests below. Write your observations and inferences in the spaces provided. You must identify any gas produced.

**Hint:** The compound for analysis in this question is a green powder.
(a) Place a spatula endful of C in an ignition tube and heat gently then strongly.

(b) Place 2 spatula endfuls of solid C in a boiling tube. Add dilute nitric acid, a little at a time until the solid dissolves. Divide the resulting solution into four portions.

(i) To the first portion add sodium hydroxide solution a little at a time, until it is in excess

(iii) To the second portion add aqueous ammonia, a little at a time until it is in excess

(iii) To the third portion add sodium sulphate solution

(iv) To the fourth portion add barium chloride solution.

Test paper Three  Paper 2

Attempt all questions

1. Petrol (octane) is a long chain alkane and can be converted to ethene as follows:

\[
\text{C}_8\text{H}_{18}(l) \rightarrow 4\text{C}_2\text{H}_4(g) + \text{H}_2(g)
\]

octane  ethene  hydrogen

(a) (i) What is the name of the process by which the products are obtained from octane?

(ii) Give two conditions needed for this reaction.
(b) Unleaded fuel is now widely used and has to be used in modern cars fitted with catalytic converters.
(i) State the advantage of unleaded petrol over ‘leaded’ petrol.
(ii) What is the role of catalytic converters? Why wouldn’t the converters work with ‘leaded petrol’?

2. A natural element represented by letter X has two types of atoms. The composition of the particles is summarised below.

<table>
<thead>
<tr>
<th>Type of atom</th>
<th>Particles present</th>
<th>% composition</th>
</tr>
</thead>
<tbody>
<tr>
<td>63 (\text{X}) (\text{29})</td>
<td>29, 34</td>
<td>69.1</td>
</tr>
<tr>
<td>65 (\text{X}) (\text{29})</td>
<td>29, __</td>
<td>__</td>
</tr>
</tbody>
</table>

(a) Copy the table and complete the missing numbers.
(b) What is the name assigned to these two types of atoms?
(c) Calculate the relative atomic mass of X.

3. (a) State the Graham’s Law of diffusion.
(b) Hydrogen diffuses through a porous plug in just 80 seconds. The same volume of a gas Y diffuses through the same plug in 300 seconds. Calculate the molecular mass of gas Y. (H = 1)

4. Ammonia is prepared in the laboratory by reacting an ammonium salt and an alkali.
(a) (i) Give an example of an ammonium salt and an alkali that we can use to prepare ammonia.
(ii) Name the drying agent used.
(b) Explain briefly the following properties of ammonia
(i) It is collected by upward delivery.
(ii) It is not dried by concentrated sulphuric acid.
(c) A student collected 120 cm\(^3\) of ammonia at r.t.p.
(i) How many moles did the student collect?
(ii) How many grams of the gas were collected? (Molar gas volume at r.t.p = 24 dm\(^3\), N = 14, H = 1)

5. A hydrocarbon contains 92.3% carbon. Find its molecular formula if its relative molecular mass is 26.
APPENDICES

Appendix 1: Qualitative analysis

Introduction
This will be part of the practical paper. The examiner expects you to be able to apply chemistry knowledge, utilise skills in carrying out experiments and use experimental observations to make logical conclusions.

- If successful analyses are to be performed, the procedure must be adhered to strictly. Omissions or alterations of instructions will lead to wrong observations and hence wrong conclusions.
- All the chemicals which will be supplied to you by your teacher during the normal teaching or examination are pure, unless they have been deliberately mixed with something else. Care must be taken to keep them pure.
- Contamination of a reagent may spoil an experiment. For this reason, reagents should not be returned to a bottle, once they have been removed from it. Droppers, once used should be returned to their original reagent bottles or beakers.
- If you feel, during normal learning that you can make a logical improvement on the method or procedure suggested, suggest to the teacher.
- When working in small groups, ensure that you are involved in performing the experiment. Participate actively to learn.
Test for negative ions (anions)

<table>
<thead>
<tr>
<th>Negative ion</th>
<th>Tests</th>
</tr>
</thead>
</table>
| Carbonate (CO$_3^{2-}$) | When dilute acids, for example, hydrochloric acid is added to a carbonate, carbon(IV) oxide gas is produced. Carbon(IV) oxide forms a white precipitate with calcium hydroxide solution.  
**Note:** The action of heat on some carbonates will produce carbon(IV) oxide which forms a white precipitate with calcium hydroxide solution. |
| Hydrogen carbonate (HCO$_3^-$) | When a hydrogen carbonate is heated, carbon(IV) oxide and water are produced. It also produces carbon(IV) oxide when reacted with dilute acids. |
| Chloride (Cl$^-$) | When a solid chloride is heated with concentrated sulphuric acid, a colourless gas is produced. The gas, hydrogen chloride forms misty fumes in moist air and forms white fumes when it is mixed with ammonia gas.  

\[
\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s)
\]
When a solution of a chloride e.g. sodium chloride (NaCl) is acidified with nitric acid and silver nitrate solution is added, a white precipitate of silver chloride is formed  

\[
\text{NaCl(aq)} + \text{AgNO}_3(aq) \rightarrow \text{AgCl(s)} + \text{NaNO}_3(aq)
\]

| Bromide (Br$^-$) | When a solution of a bromide, e.g sodium bromide (NaBr) is acidified with dilute nitric acid and silver nitrate solution added, a yellow precipitate of silver bromide is formed.  

\[
\text{NaBr(aq)} + \text{AgNO}_3(aq) \rightarrow \text{AgBr(s)} + \text{NaNO}_3(aq)
\]

| Iodide (I$^-$) | When a solution of an iodide e.g. sodium iodide (NaI) is acidified with nitric acid and silver nitrate solution added, a yellow precipitate of silver iodide is formed.  

\[
\text{NaI(aq)} + \text{AgNO}_3(aq) \rightarrow \text{AgI(s)} + \text{NaNO}_3(aq)
\]

| Sulphate (SO$_4^{2-}$) | When dilute nitric or hydrochloric acid and barium chloride or barium nitrate solution are added to a solution of a sulphate, a white precipitate of barium sulphate is formed immediately.  

**Example:**  

\[
\text{K}_2\text{SO}_4(aq) + \text{BaCl}_2(aq) \rightarrow \text{BaSO}_4(s) + 2\text{KCl(aq)}
\]
<table>
<thead>
<tr>
<th><strong>Nitrate (NO₃⁻)</strong></th>
<th>When a solution of a nitrate is added to an equal volume of fresh iron(II) sulphate solution, and with the test tube slopping concentrated sulphuric acid is carefully poured down its side, a “brown ring” forms where the two layers meet. The formula of the brown substance is FeSO₄·NO.</th>
</tr>
</thead>
</table>
| **Sulphide (S²⁻)** | When dilute hydrochloric acid is added to a sulphide, colourless hydrogen sulphide gas is produced. Hydrogen sulphide gas smells like rotten eggs and turns filter paper soaked in lead nitrate solution or lead ethanoate solution black.  
**Example:**  
\[
\text{FeS(s)} + 2\text{HCl(aq)} \rightarrow \text{FeCl}_2(\text{aq}) + \text{H}_2\text{S(g)}
\] |
| **Sulphite (SO₃²⁻)** | When dilute hydrochloric acid is added to a sulphite and the mixture is heated, colourless sulphur(IV) oxide gas is produced. The sulphur(IV) oxide has a pungent irritating smell. It turns orange potassium dichromate(VI) solution green. It also decolourises potassium manganate(VII) solution.  
**Example**  
\[
\text{Na}_2\text{SO}_3(\text{s}) + 2\text{HCl(aq)} \rightarrow 2\text{NaCl(aq)} + \text{SO}_2(\text{g}) + \text{H}_2\text{O(l)}
\]  
The sulphite ion can also be tested with barium nitrate or barium chloride solution. Formation of a white precipitate which is soluble in dilute nitric acid or hydrochloric acid confirms presence of SO₃²⁻.  
\[
\text{Ba}^{2⁺}(\text{aq}) + \text{SO}_3^{2⁻}(\text{aq}) \rightarrow \text{BaSO}_3(\text{s}), \text{white precipitate}
\]  
\[
\text{BaSO}_3(\text{s}) + 2\text{H}^⁺(\text{aq}) \rightarrow \text{Ba}^{2⁺}(\text{aq}) + \text{H}_2\text{O(l)} + \text{SO}_2(\text{g})
\] |
Table 2: Test for common gases

<table>
<thead>
<tr>
<th>Gas</th>
<th>Formula</th>
<th>Colour</th>
<th>Smell</th>
<th>Test with moist litmus paper</th>
<th>Test with lighted splint</th>
<th>Other tests</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>H₂</td>
<td>Colourless</td>
<td>Odourless</td>
<td>No change</td>
<td>• Pop sound • Splint ex-tinguished</td>
<td>• Burns with a blue flame to form water</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O₂</td>
<td>Colourless</td>
<td>Odourless</td>
<td>No change</td>
<td>• Relights a glowing splint</td>
<td>• Forms NO₂ with NO.</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N₂</td>
<td>Colourless</td>
<td>Odourless</td>
<td>no change</td>
<td>Extinguished</td>
<td>• Forms magnesium nitride with burning magnesium</td>
</tr>
<tr>
<td>Chlorine</td>
<td>Cl₂</td>
<td>Green yellow</td>
<td>Pungent irritating</td>
<td>Blue to red then bleaches</td>
<td>Extinguished</td>
<td>—</td>
</tr>
<tr>
<td>Hydrogen chloride</td>
<td>HCl</td>
<td>Colourless</td>
<td>Pungent irritating</td>
<td>blue to red</td>
<td>Extinguished</td>
<td>• Forms white fumes of ammonium chloride with ammonia</td>
</tr>
<tr>
<td>Carbon(IV) oxide</td>
<td>CO₂</td>
<td>Colourless</td>
<td>Odourless</td>
<td>From blue to slightly red</td>
<td>Extinguished</td>
<td>• Forms a white precipitate with aqueous calcium hydroxide</td>
</tr>
<tr>
<td>Carbon(II) oxide</td>
<td>CO</td>
<td>Colourless</td>
<td>Odourless</td>
<td>No change</td>
<td>Burns with blue flame</td>
<td>Burns with a blue flame to form carbon(IV) oxide</td>
</tr>
<tr>
<td>Ammonia</td>
<td>NH₃</td>
<td>Colourless</td>
<td>Pungent choking</td>
<td>Turns red litmus blue.</td>
<td>Extinguished</td>
<td>Forms white fumes of ammonium chloride with hydrogen chloride</td>
</tr>
<tr>
<td>Sulphur(IV) oxide</td>
<td>SO₂</td>
<td>Colourless</td>
<td>Pungent irritating</td>
<td>Turns blue litmus red</td>
<td>Extinguished</td>
<td>• Turns potassium dichromate(VI) to a green solution • It decolourises potassium manganate(VII) solution.</td>
</tr>
<tr>
<td>Hydrogen sulphide</td>
<td>H₂S</td>
<td>Colourless</td>
<td>Rotten eggs smell</td>
<td>Blue to slightly red</td>
<td>Burns producing sulphur(IV) oxide.</td>
<td>• Lead nitrate paper turns black • Lead ethanoate paper turns black.</td>
</tr>
<tr>
<td></td>
<td>NO₂</td>
<td>Red brown</td>
<td>Pungent irritating</td>
<td>Blue to red</td>
<td>Extinguishes</td>
<td>—</td>
</tr>
<tr>
<td>----------------</td>
<td>-----</td>
<td>-----------</td>
<td>--------------------</td>
<td>-------------</td>
<td>--------------</td>
<td>-------</td>
</tr>
<tr>
<td>Nitrogen (VI) oxide</td>
<td>N₂O₃</td>
<td>Colourless</td>
<td>slightly sweet</td>
<td>No change</td>
<td>Relights a brightly glowing splint</td>
<td>Fairly soluble in water</td>
</tr>
<tr>
<td>Nitrogen (I) oxide</td>
<td>NO</td>
<td>Colourless</td>
<td>—</td>
<td>—</td>
<td>—</td>
<td>Forms red brown fumes in air</td>
</tr>
<tr>
<td>Nitrogen (II) oxide</td>
<td>NO</td>
<td>Colourless</td>
<td>—</td>
<td>—</td>
<td>—</td>
<td>—</td>
</tr>
<tr>
<td>Methane</td>
<td>CH₄</td>
<td>Colourless</td>
<td>odourless</td>
<td>No change</td>
<td>Burns with luminous flame</td>
<td>—</td>
</tr>
</tbody>
</table>
| Ethene | C₂H₄ | Colourless | odourless | No change | Burns with sooty flame | • Decolourises bromine water and acidified potassium manganate(VII) solution  
• Turns acidified potassium dichromate(VI) solution from orange to green. |
| Ethyne | C₂H₂ | Colourless | odourless | No change | Burns with very sooty flame | • Decolourises bromine water and acidified potassium manganate(VII) solution  
• Turns acidified potassium dichromate(VI) solution from orange to green. |

**Test for positive ions (cations)**

There are two tests which can be used to identify positive ions (cations)

1. **Test with sodium hydroxide solution**

   When sodium hydroxide solution is added to a small quantity of the compound in solution form drop by drop until in excess, the following can be observed:
   - A precipitate may be observed which does not dissolve in excess solution of sodium hydroxide.
   - A precipitate may be observed which is soluble in excess sodium hydroxide solution.
In some compounds there is no effect when sodium hydroxide solution is added. These are summarised in the table below:

Table 3: Test for positive ions (cations)

<table>
<thead>
<tr>
<th>Cation</th>
<th>Few drops</th>
<th>Excess</th>
<th>Ionic equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium, K⁺</td>
<td>No precipitate</td>
<td>No precipitate</td>
<td>—</td>
</tr>
<tr>
<td>Sodium, Na⁺</td>
<td>No precipitate</td>
<td>No precipitate</td>
<td>—</td>
</tr>
<tr>
<td>Calcium, Ca²⁺</td>
<td>White precipitate</td>
<td>Precipitate</td>
<td>( \text{Ca}^{2+}(aq) + 2\text{OH}^–(aq) \rightarrow \text{Ca(OH)}_2(s) )</td>
</tr>
<tr>
<td>Magnesium, Mg²⁺</td>
<td>White precipitate</td>
<td>Precipitate</td>
<td>( \text{Mg}^{2+}(aq) + 2\text{OH}^–(aq) \rightarrow \text{Mg(OH)}_2(s) )</td>
</tr>
<tr>
<td>Aluminium, Al³⁺</td>
<td>White precipitate</td>
<td>Soluble to form a colourless solutions</td>
<td>( \text{Al}^{3+}(aq) + 3\text{OH}^–(aq) \rightarrow \text{Al(OH)}_3(s) )</td>
</tr>
<tr>
<td>Iron(II), Fe²⁺</td>
<td>Green precipitate</td>
<td>Precipitate</td>
<td>( \text{Fe}^{2+}(aq) + 2\text{OH}^–(aq) \rightarrow \text{Fe(OH)}_2(s) )</td>
</tr>
<tr>
<td>Lead(II), Pb²⁺</td>
<td>White precipitate</td>
<td>Soluble to form a colourless solution</td>
<td>( \text{Pb}^{2+}(aq) + 2\text{OH}^–(aq) \rightarrow \text{Pb(OH)}_2(s) )</td>
</tr>
<tr>
<td>Copper(II), Cu²⁺</td>
<td>Blue precipitate</td>
<td>Precipitate</td>
<td>( \text{Cu}^{2+}(aq) + 2\text{OH}^–(aq) \rightarrow \text{Cu(OH)}_2(s) )</td>
</tr>
<tr>
<td>Zinc, Zn²⁺</td>
<td>White precipitate</td>
<td>Soluble to form a colourless solution</td>
<td>( \text{Zn}^{2+}(aq) + 2\text{OH}^– \rightarrow \text{Zn(OH)}_2(s) )</td>
</tr>
<tr>
<td>Iron(III), Fe³⁺</td>
<td>Red-brown precipitate</td>
<td>Precipitate</td>
<td>( \text{Fe}^{3+}(aq) + 3\text{OH}^–(aq) \rightarrow \text{Fe(OH)}_3(s) )</td>
</tr>
</tbody>
</table>

2. Test with ammonium hydroxide solution
This solution is also called aqueous ammonia or ammonia solution.
To the solution of the compound add small quantities of aqueous ammonia until it is present in excess. The following can be observed:
- A precipitate maybe formed which does not dissolve in excess solution of aqueous ammonia.

360
A precipitate may be formed which is soluble in excess aqueous ammonia. In some compounds, there is no effect when aqueous ammonium is added. These are summarised in the table below.

**Table 4: Test for positive ions (cations)**

<table>
<thead>
<tr>
<th>Cation</th>
<th>Few drops</th>
<th>Excess aqueous ammonia</th>
<th>Ionic equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Potassium, K⁺</td>
<td>No precipitate</td>
<td>No precipitate</td>
<td>—</td>
</tr>
<tr>
<td>Sodium, Na⁺</td>
<td>No precipitate</td>
<td>No precipitate</td>
<td>—</td>
</tr>
<tr>
<td>Calcium, Ca²⁺</td>
<td>No precipitate</td>
<td>Ca²⁺(aq) + 2OH⁻(aq) → Ca(OH)₂(aq)</td>
<td></td>
</tr>
<tr>
<td>Magnesium, Mg²⁺</td>
<td>White precipitate</td>
<td>Precipitate insoluble</td>
<td>Mg²⁺(aq) + 2OH⁻(aq) → Mg(OH)₂(s)</td>
</tr>
<tr>
<td>Aluminium, Al³⁺</td>
<td>White precipitate</td>
<td>Precipitate insoluble</td>
<td>Al³⁺(aq) + 3OH⁻(aq) → Al(OH)₃(s)</td>
</tr>
<tr>
<td>Iron(II), Fe²⁺</td>
<td>Green precipitate</td>
<td>Precipitate insoluble</td>
<td>Fe²⁺(aq) + 2OH⁻(aq) → Fe(OH)₂(s)</td>
</tr>
<tr>
<td>Iron(III), Fe³⁺</td>
<td>Red-brown precipitate</td>
<td>Precipitate insoluble</td>
<td>Fe³⁺(aq) + 3OH⁻(aq) → Fe(OH)₃(s)</td>
</tr>
<tr>
<td>Lead(II), Pb²⁺</td>
<td>White precipitate</td>
<td>Precipitate insoluble</td>
<td>Pb²⁺(aq) + 2OH⁻(aq) → Pb(OH)₂(s)</td>
</tr>
<tr>
<td>Copper(II), Cu²⁺</td>
<td>Blue precipitate</td>
<td>soluble in excess to form a deep blue solution</td>
<td>Cu²⁺(aq) + 2OH⁻(aq) → Cu(OH)₂(s)</td>
</tr>
<tr>
<td>Zinc, Zn²⁺</td>
<td>White precipitate</td>
<td>Soluble to form a colourless solution.</td>
<td>Zn²⁺(aq) + 2OH⁻(aq) → Zn(OH)₂(s)</td>
</tr>
</tbody>
</table>
In the above reactions, there are compounds in solutions that behave differently with sodium hydroxide solution and aqueous ammonia solution. These are summarised as shown in the table below.

Table 5: Summary of the test for cations

<table>
<thead>
<tr>
<th>Cation</th>
<th>Addition of sodium hydroxide</th>
<th>Addition of aqueous ammonia</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calcium, Ca²⁺</td>
<td>White precipitate insoluble in excess</td>
<td>No precipitate</td>
</tr>
<tr>
<td>aluminium, Al³⁺</td>
<td>White precipitate soluble in excess</td>
<td>White precipitate insoluble in excess</td>
</tr>
<tr>
<td>Lead(II), Pb²⁺</td>
<td>Blue precipitate insoluble in excess</td>
<td>Blue precipitate soluble in excess to form a deep blue solution.</td>
</tr>
<tr>
<td>Copper, Cu²⁺</td>
<td>Blue precipitate insoluble in excess</td>
<td>Blue precipitate soluble in excess</td>
</tr>
<tr>
<td>Zinc, Zn²⁺</td>
<td>White precipitate soluble in excess</td>
<td>White precipitate soluble in excess</td>
</tr>
</tbody>
</table>

1. **What to suspect when a substance is heated in a test tube or boiling tube and you observe the following:**

- **Sublimation**
  (i) It can be an ammonium compound e.g. ammonium chloride
  (ii) It can be iodine.
  (iii) It can be aluminium chloride or iron(III) chloride.

- **Water vapour**
  Water vapour is colourless and forms colourless droplets on the cool part of the tube. The substance can be:
  (i) Hydrated salt i.e a salt with water of crystallisation.
  (ii) A hydrogen carbonate

- **Crackling sound**
  The substance can be:
  (i) Sodium chloride
  (ii) Lead(II) nitrate, but this will also produce a red-brown gas (nitrogen(IV) oxide and oxygen).
• **A solid residue**
  You are expected to observe the residue when you decompose a compound when hot and after it has cooled.
  (i) Zinc oxide is yellow when hot and white on cooling.
  (ii) Lead(II) oxide is red-brown when hot and yellow on cooling.
  (iii) Copper(II) oxide is black and no change on cooling.
  (iv) Iron(III) oxide is red-brown and no change on cooling.

2. **What to suspect when you add dilute hydrochloric acid, nitric acid or sulphuric acid on a solid and you get the following:**

• **Hydrogen gas is evolved**
  The substance is likely to be a metal e.g. zinc, iron or magnesium.

• **Carbon(IV) oxide gas evolved**
  The substance is likely to be a metal carbonate or hydrogen carbonate.

• **Chlorine gas evolved**
  The substance is likely to be calcium hypochlorite.

• **Sulphur(IV) oxide gas evolved**
  The substance is likely to be a metal sulphite e.g. sodium sulphite.

3. **What to suspect when a metal is added to water or solution and you observe the following:**

• **Bubbles (effervescence)**
  The metal reacts with water to produce hydrogen. This metal is higher in the reactivity series than aluminium.

• **A solid is deposited**
  The metal added is higher in the reactivity series than the metal whose ions are in the solution. It is a displacement reaction. For example:

\[
\text{CuSO}_4(\text{aq}) + \text{Zn(s)} \rightarrow \text{Cu(s)} + \text{ZnSO}_4(\text{aq})
\]

  red brown
  solid

4. **What to suspect when you are given the following coloured substances.**

• **Blue**
  These are usually copper(II) compounds.
• **Pale green**
  These are usually iron(II) compounds or copper(II) compounds.

• **Black**
  This is usually copper(II) oxide, manganese(IV) oxide or carbon. Carbon does not react with dilute sulphuric acid but copper(II) oxide forms a blue solution of copper(II) sulphate with dilute sulphuric acid. Manganese(IV) oxide reacts with hydrogen peroxide to form oxygen and water.

**Worked examples of qualitative analysis**

**Introduction**

- In a practical examination, you are given time to read the question paper and make sure that you have all the chemicals and the apparatus you may require.
- You are advised to write neat and accurate notes of your observations as soon as you have made them. Write what you observe but not what you would like to observe. In most practical examinations, it is very easy to raise your neck, and see what your fellow student has written and copy. You are strongly advised to be independent and honest and write what you observe and make your own conclusions. This attitude will help you to grow up into a self-disciplined, self-reliant, honest and trustworthy citizen. These values are socially, morally, spiritually and politically acceptable. Your motto should be “**hard work and an honest living**”.
- Follow the instructions strictly.

**Examples**

1. A student carried out the following experiments with solid X. All gases evolved were identified. Observations and conclusions were recorded as in Table 5. Substance X was to be identified.

   Table 5 shows how the student recorded the results of the qualitative analysis.
Table 5: Procedures, Observations and Conclusions

<table>
<thead>
<tr>
<th>Procedures</th>
<th>Observations</th>
<th>Conclusions</th>
</tr>
</thead>
<tbody>
<tr>
<td><em>(a)</em> Put a spatula endful of X in a dry test tube. Heat gently and then strongly.</td>
<td>• Colourless gas which turns moist blue litmus slightly red is produced. Forms a white precipitate with aqueous calcium hydroxide. • Yellow residue when hot. Residue turns white on cooling.</td>
<td>• Carbon (IV) oxide produced. • Carbonate ions present in X. • Zinc oxide formed. Zn$^{2+}$ ions present in X.</td>
</tr>
<tr>
<td><em>(b)</em> Put half spatula endful of X into a clean test tube. Add about 2 cm$^3$ of dilute hydrochloric acid. Keep the resulting solution.</td>
<td>• Effervescence (bubbles). • Colourless gas which turns moist blue litmus paper slightly red Forms a white precipitate with aqueous calcium hydroxide.</td>
<td>• Carbon(IV) oxide produced • CO$_3^{2-}$ ions present in X.</td>
</tr>
<tr>
<td><em>(c)</em> To the solution in <em>(b)</em> add about 5 cm$^3$ of distilled water. Divide this solution into two portions.</td>
<td>• Colourless solution formed.</td>
<td>• Coloured ions absent e.g. Cu$^{2+}$, Fe$^{2+}$, Fe$^{3+}$.</td>
</tr>
<tr>
<td><em>(i)</em> To the first portion, add aqueous sodium hydroxide drop by drop until in excess.</td>
<td>• White precipitate, soluble in excess to form a colourless solution.</td>
<td>• Zn$^{2+}$ ions present in X.</td>
</tr>
<tr>
<td><em>(ii)</em> To the second portion, add aqueous ammonia drop by drop until in excess.</td>
<td>• White precipitate soluble in excess to form a colourless solution.</td>
<td>• Zn$^{2+}$ ions in X confirmed</td>
</tr>
</tbody>
</table>

The student can summarise the results as follows:

• The cation present in solid X is Zn$^{2+}$ ion.
• The anion present in solid X is CO$_3^{2-}$ ion.
• Solid X is zinc carbonate.
2. Carry out the following experiments with solid Y. You should identify any gas evolved. Record your observations and conclusions in the table below.

### Table 6: Procedure, observations and conclusions

<table>
<thead>
<tr>
<th>Procedure</th>
<th>Observations</th>
<th>Conclusions</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a) Heat one spatula endful of Y in a test-tube</td>
<td>• Colourless gas with a pungent choking smell which turns moist red litmus paper blue. • It forms white fumes with concentrated hydrochloric acid on a glass rod. A colourless liquid formed on the cool part of the test-tube.</td>
<td>• Ammonia produced • Ammonium ion present in Y • Y is hydrated or Y has water of crystallisation.</td>
</tr>
<tr>
<td>b) Put half a spatula endful of Y in a test tube. Add about 4 cm³ of distilled water to dissolve it. Add a little aqueous sodium hydroxide and heat.</td>
<td>Y dissolves to form a colourless solution. Produced a colourless gas with a pungent choking smell. Turns moist red litmus paper blue. It gives white fumes with concentrated hydrochloric acid on a glass rod.</td>
<td>• Y is soluble in water. Coloured ions eg Cu²⁺; Fe²⁺, Fe³⁺ absent. • NH⁴⁺ gas is formed. • NH⁴⁺ ions present in Y.</td>
</tr>
<tr>
<td>c) Put one spatula endful of Y in a boiling tube. Dissolve it using 25 cm³ of distilled water. Divide the solution into four portions. (i) To the first portion, add aqueous sodium hydroxide drop by drop till excess. (ii) To the second portion, add aqueous ammonia drop by drop till in excess. (iii) To the third portion, add dilute sulphuric acid. (iv) To the fourth portion, add dilute nitric acid, followed by barium nitrate solution.</td>
<td>• White precipitate soluble in excess. • White precipitate insoluble in excess. No observable change reaction/ no precipitate White precipitate</td>
<td>— Zn²⁺ ions, Al³⁺ ions, Pb²⁺ ions. Al³⁺ ions, Pb²⁺ ions. Al³⁺ ions present in Y Barium sulphate formed. SO₄²⁻ ions present in Y</td>
</tr>
</tbody>
</table>
• The cations present in Y are aluminium, ($\text{Al}^{3+}$) ions and ammonium, ($\text{NH}_4^+$) ions.
• The anion present in Y is sulphate ion ($\text{SO}_4^{2-}$)

3. Carry out the following experiments with P which is a mixture of two compounds. You should identify any gas evolved. Record your observations and conclusions in the table below.

**Table 7: Procedure, Observations and conclusions**

<table>
<thead>
<tr>
<th>Procedure</th>
<th>Observations</th>
<th>Conclusions</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a) Put a spatula endful of P into a boiling tube. Add about 15cm$^3$ of distilled water, shake well and filter. Keep both the residue and the filtrate.</td>
<td>• A green residue and blue filtrate.</td>
<td>• One compound of P is insoluble. • Cu$^{2+}$ ions may be present.</td>
</tr>
<tr>
<td>(b) Using narrow spatula put the residue into a clean test tube. Add about 2cm$^3$ of dilute nitric acid. Keep the solution.</td>
<td>• Effervescence; a colourless gas which turns moist blue litmus paper slightly red. • Forms a white precipitate with calcium hydroxide solution</td>
<td>• Carbon(IV) oxide produced. • CO$_3^{2-}$ ions present in the insoluble compound of P (residue).</td>
</tr>
<tr>
<td>(c) To solution from (b) add aqueous ammonia drop by drop until in excess.</td>
<td>• Blue precipitate which dissolves in excess to form a deep blue solution.</td>
<td>• Cu$^{2+}$ ions present in the residue</td>
</tr>
<tr>
<td>(d) Divide the filtrate in (a) into three portions</td>
<td>—</td>
<td></td>
</tr>
<tr>
<td>(i) To the first portion add aqueous sodium hydroxide drop by drop until in excess.</td>
<td>• Blue precipitate insoluble in excess. • Blue precipitate soluble in excess.</td>
<td>• Cu$^{2+}$ ions present in the filtrate.</td>
</tr>
</tbody>
</table>
(ii) To the second portion add aqueous ammonia drop by drop until in excess.

(iii) Evaporate the third portion to dryness and then heat the remaining solid strongly.

- A red-brown gas which turns moist blue litmus paper red evolved.
- A gas which relights a glowing splint also evolved. A black residue formed.
- Nitrogen(IV) oxide NO$_3^-$ ions present in the soluble compound of P Oxide evolved. Copper (II) oxide formed.
- Cu$^{2+}$ ions confirmed in the filtrate.

The ions present in mixture P are copper (II) ions, carbonate ions and nitrate ions. Explain how you can confirm the presence of the anion in the filtrate in experiment (a).

**Exercises**

1. (i) Give two colourless gases which have no smell.
   (ii) Give two coloured gases and two colourless gases which have a smell.

2. How can you identify:
   (i) Ammonia gas.
   (ii) Solid chloride salt
   (iii) Zinc ions

3. (i) Name one salt which is blue in colour.
   (ii) Name one salt which is green in colour.
   (iii) Name one salt which is pale-green in colour.

4. (i) Why do we add nitric acid in the test of chloride when using silver nitrate solution?
   (ii) Why do we add nitric acid in the test of sulphate when using barium nitrate or barium chloride solutions?
### APPENDIX II

**Periodic table**

**Groups**

- **I**
- **II**

**Periods**

1. **1**
   - **H** Hydrogen
   - **He** Helium

2. **2**
   - **Li** Lithium
   - **Be** Beryllium
   - **B** Boron
   - **C** Carbon
   - **N** Nitrogen
   - **O** Oxygen

3. **3**
   - **F** Fluorine
   - **Ne** Neon
   - **Na** Sodium
   - **Mg** Magnesium
   - **Al** Aluminium
   - **Si** Silicon
   - **P** Phosphorus
   - **S** Sulfur
   - **Cl** Chlorine
   - **Ar** Argon

4. **4**
   - **K** Potassium
   - **Ca** Calcium
   - **Sc** Scandium
   - **Ti** Titanium
   - **V** Vanadium
   - **Cr** Chromium
   - **Mn** Manganese
   - **Fe** Iron
   - **Co** Cobalt
   - **Ni** Nickel
   - **Cu** Copper
   - **Zn** Zinc
   - **Ga** Gallium
   - **Ge** Germanium
   - **As** Arsenic
   - **Se** Selenium
   - **Br** Bromine
   - **Kr** Krypton

5. **5**
   - **Rb** Rubidium
   - **Sr** Strontium
   - **Y** Yttrium
   - **Zr** Zirconium
   - **Nb** Niobium
   - **Mo** Molybdenum
   - **Tc** Technetium
   - **Ru** Ruthenium
   - **Rh** Rhodium
   - **Pd** Palladium
   - **Ag** Silver
   - **Cd** Cadmium
   - **In** Indium
   - **Sn** Tin
   - **Sb** Antimony
   - **Te** Tellurium
   - **I** Iodine
   - **Xe** Xenon

6. **6**
   - **Cs** Cesium
   - **Ba** Barium
   - **La** Lanthanum
   - **Ce** Cerium
   - **Pr** Praseodymium
   - **Nd** Neodymium
   - **Pm** Promethium
   - **Sm** Samarium
   - **Eu** Europium
   - **Gd** Gadolinium
   - **Tb** Terbium
   - **Dy** Dysprosium
   - **Ho** Holmium
   - **Er** Erbium
   - **Tm** Thulium
   - **Yb** Ytterbium
   - **Lu** Lutetium

7. **7**
   - **Fr** Francium
   - **Ra** Radium
   - **Ac** Actinium
   - **Th** Thorium
   - **Pa** Protactinium
   - **U** Uranium
   - **Np** Neptunium
   - **Pu** Plutonium
   - **Am** Americium
   - **Cm** Curium
   - **Bk** Berkelium
   - **Cf** Californium
   - **Es** Eka-mercury
   - **Fm** Flerovium
   - **Md** Magmatic
   - **No** Nihonium

**Lanthanides**

- 257 La Lanthanum
- 140 Ce Cerium
- 144 Nd Neodymium
- 147 Pr Promethium
- 150 Sm Samarium
- 152 Eu Europium
- 157 Gd Gadolinium
- 159 Tb Terbium
- 163 Dy Dysprosium
- 165 Ho Holmium
- 167 Er Erbium
- 169 Tm Thulium

**Actinides**

- 227 Ac Actinium
- 232 Th Thorium
- 233 U Uranium
- 235 Np Neptunium
- 237 Pu Plutonium
- 241 Am Americium
- 244 Cm Curium
- 247 Bk Berkelium
- 248 Cf Californium
- 250 Es Eka-mercury
- 253 Fm Flerovium
- 256 Md Magmatic
- 254 No Nihonium
# Appendix

Atomic numbers and relative atomic masses of some elements

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Atomic number</th>
<th>Relative atomic mass</th>
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<tr>
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Glossary

A

Absolute temperature:
It is the temperature measured with respect to absolute zero on the Kelvin scale.

Absolute zero:
This is the lowest possible temperature that can be achieved. It is the starting point of absolute temperature scale (Kelvin scale). It is –273˚C.

Addition polymer:
A polymer formed by an addition reaction—the monomer molecules must contain C = C double bond or C ≡ C triple bond; for example polythene is formed from ethene.

Alkane
A series of hydrocarbons with the general formula \((C_nH_{2n+2})\). They are saturated compounds as they have only single bonds between carbon atoms in their structure.

Alkenes:
A series of hydrocarbons with the general formula \((C_nH_{2n})\). They are unsaturated molecules as they have a C = C (double bond) somewhere in the chain.

Alkynes:
A series of hydrocarbons with the general formula \((C_nH_{2n-2})\). They are unsaturated molecules as have a C ≡ C triple bond somewhere in the chain.

Artificial fertiliser:
A substance added to soil to increase the amount of elements such as nitrogen, potassium and phosphorus (NPK fertilisers). This enables crops to grow more healthy and produces higher yields.

Avogadro number or constant:
The number of particles (atoms, molecules, ions or electrons) in one mole of a substance. It is usually \(6.023 \times 10^{23}\).

B

Bleaching:
The process by which colour is partially or wholly removed from a coloured substance.
Bleaching powder:
A chlorine-containing substance used in bleaching textiles or sterilising water. It is made from solid calcium hydroxide and chlorine. Its chemical name is calcium hypochlorite.

Boyle’s Law:
The volume of a given mass of a gas is inversely proportional to its pressure provided that the temperature remains constant.

Bromination:
The addition of bromine to an unsaturated organic compound e.g. ethene plus bromine

C
Charles Law:
The volume of a given mass of a gas at constant pressure is directly proportional to its absolute temperature.

Chloride test:
A chloride ion is present in a substance if it forms a white precipitate when dilute nitric acid and silver nitrate solution are added to a solution of the substance.

Chlorination:
Addition of chlorine to an unsaturated organic compound or the treatment of water with chlorine to destroy germs.

Constant
A magnitude that does not change in value e.g Avogadro’s constant (6.023 \times 10^{23}).

Contact process
The industrial manufacture of sulphuric acid in which the sulphur(IV) oxide and oxygen come in contact with vanadium(V) oxide catalyst.

Cracking:
The breaking down of long hydrocarbon molecules with heat and/or a catalyst to produce short hydrocarbon molecule and alkene.

D
Diffusion:
The process by which particles spread out as a result of the random motion.
Displacement reactions:
A reaction in which a more reactive element displaces a less reactive element from a solution of its salt.

E
Empirical formula
A chemical formula for a compound which shows the simplest ratio of atoms present. It does not give the actual number of each atoms in most cases e.g., the empirical formula of ethane (C₂H₆) is CH₃.

Excess:
In a chemical reaction there is a connection between the quantities of substances reacting. In practice if one of the reactants is present in larger quantities than is required for the reaction it is said to be in excess.

Exothermic reaction:
A reaction that gives out heat e.g. burning of magnesium.

F
Frasch process:
The process of obtaining sulphur from sulphur beds below the earth’s surface. Superheated water is pumped down a shaft to liquefy the sulphur, which is then forced to the surface.

Free radicals:
Atoms or groups of atoms with unpaired electrons and therefore highly reactive. They can be produced in high-energy radiation such as ultra-violet light in photo-chemical reactions.

Functional group:
Refers to the site of a chemical reaction in an organic compound.

H
Homologous series.
The name given to a family of organic compounds e.g. alkane. Members of the series conform to a general formula and show similar chemical reactions, varying only in their reactivity.

Hydrocarbons:
Compounds which contain carbon and hydrogen only, e.g. methane (CH₄)

Hydrogenation:
The chemical addition of hydrogen to a substance, e.g. ethene (C2H4) to give ethane
(C2H6). Oil plus hydrogen to give margarine is also called **hydrogenation**.

**I**

**Inert substance:**
A substance that reacts passively with other substances, e.g. nitrogen reacts with air only at very high temperatures.

**Isomers:**
Compounds which have the same molecular formula but different structural arrangements of the atoms i.e. they have different structural formulae.

**Isomerism**
it is the occurrence of two or more compounds with the same molecular formula but different structural formula

**K**

**Kelvin scale**
Is a scale of temperature with absolute zero as the starting point.

**L**

**Law:**
A statement of scientific fact concerning natural phenomena, e.g. Boyle’s law.

**Liquefaction:**
The process of converting a gas to a liquid.

**M**

**Molar gas volume**
The volume of one mole of a gas at specified temperature and pressure e.g. the molar
gas volume of any gas at s.t.p is 22.4 dm³

**Morality:**
The concentration of a solution expressed in moles of solute per one dm³ of solution.

**Mole:**
The measure of amount of substance in chemistry. One mole of a substance has a mass equal to its relative formula mass in grams. This amount of substance contain 6.023 × 10²³ (Avogadro’s constant) of atoms, molecules, ions or formula unit depending on the substance considered.

**Molecular formula**
A formula which shows the actual number of atoms of element present in a molecule of a compound.

**Monomer:**
A small molecule, such as ethene, which can be polymerised to make a polymer. Many monomers are joined together to make a polymer. For example ethene can be polymerised to make polythene.

**N**
**Nitride:**
A compound of nitrogen and an electropositive element e.g. Magnesium nitride (Mg$_3$N$_2$)

**O**
**Organic chemistry:**
The study of carbon compounds, both natural and synthetic.

**Organic matter:**
Plant and animal material containing carbon as one of the elements, e.g. starch

**Ozone: (trioxygen)**
A colourless gas (O$_3$). It is produced in the upper atmosphere (10 to 50 km) above earth’s surface by the action of high-energy ultraviolet radiation on oxygen gas, producing oxygen atoms. These oxygen atoms then react to produce the ozone gas.

**Ozone depletion:**
This is the production of a “hole” in the ozone layer by the action of certain compounds, mostly CFCs.

**Ozone layer:**
A layer of ozone in the atmosphere which prevents harmful ultra-violet radiation reaching the earth’s surface.

**P**
**Petroleum:**
A liquid mixture of hydrocarbons obtained from beneath the surface of the ground.

**Phenomenon:**
A situation or an event of scientific interest susceptible of scientific description and explanation.
Polymer:
A long chain molecule built up of a number of repeating units or monomers e.g (polyethene) is a polymer made up from ethene molecules linked together

Polymerisation:
The chemical reaction in which molecules (monomers) join together to form a long-chain polymer.

Polar solvent:
The molecules of some solvent, such as water contain slight positive and negative charges due to different strengths of attraction of shared electrons making bonds.. A polar solved dissolves substances containing ionic bonds. Solvents without these charges are called non-polar solvents, e.g methylbenzene.

Q
Qualitative:
A qualitative study or analysis is one which depends upon changes in appearance only.

Quantitative:
This involves a study or analysis of quantities e.g. mass, volume among others.

R
Relative atomic mass (A_r):
The average mass of an atom of an element, taking account of the isotopes of the element on a scale where a carbon – 12 atoms has a mass of exactly 12.00.

Relative molecular mass (RMM):
The sum, of the relative atomic mass of atoms present in the molecule or one mole of a compound. It is also called relative formula mass (R.F.M)

S
Saturated organic compound:
A compound which have all the carbon atoms in its molecules linked with another by a single covalent bond, e.g. ethane, propane.

Standard pressure:
The pressure exerted by a column of mercury exactly 760 mm (76 cm) at O°C (273 K).
Standard solution
A solution whose concentration is known

Standard temperature:
O °C or 273 K

Stoichiometric equation:
A normal balanced chemical equation which shows the reactants and products and their physical states.

Substitution reaction
A reaction in which an atom (or atoms) of a hydrocarbon molecule is (are) replaced by different atoms (s), without changing the molecules general structure.

Sulphate test:
If a white precipitate is produced when dilute hydrochloric acid or nitric acid and barium chloride or nitrate solution are added to a solution of a substance, sulphate ions are present.

Systematic naming
Describes a way of naming compounds based on a special international agreed system, the IUPAC system. For example C$_2$H$_6$ is called ethane. It has two carbons. The -ane shows the absence of double or triple bond i.e.. It is a saturated hydrocarbon.

T
Titration
A method of quantitative analysis using solutions. A volume of one solution is slowly added to known volume of another solution using a burette until an end-point is reached. If an acid and alkali are used then an indicator is used to show that the end point has been reached.

Titre: The volume in cm$^3$ of solution run from the burette during titration. It is also called titration value.

U
Unsaturated organic compound.
A compound which has the carbon atoms linked by two (double) or three (triple) covalent bonds e.g. ethene (H$_2$C = CH$_2$), ethyne (HC ≡ CH)

V
Vulcanisation
The process of converting weak plastic inelastic raw rubber into a strong elastic non-plastic material. For example heating of rubber with sulphur.