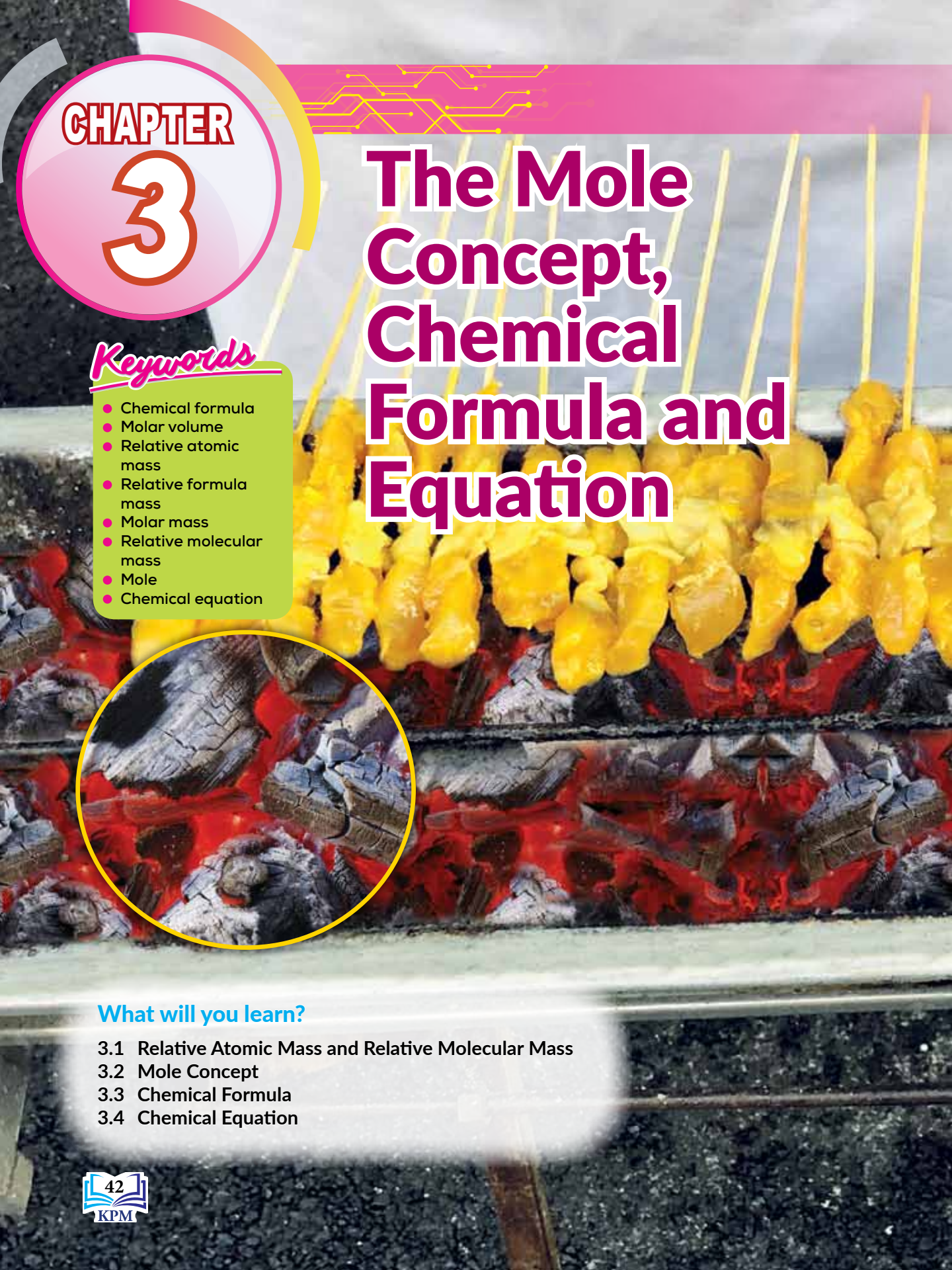
CHAPTER 3

The Mole Concept, Chemical Formula and Equation



Keywords

- Chemical formula
- Molar volume
- Relative atomic mass
- Relative formula mass
- Molar mass
- Relative molecular mass
- Mole
- Chemical equation

What will you learn?

3.1 Relative Atomic Mass and Relative Molecular Mass
3.2 Mole Concept
3.3 Chemical Formula
3.4 Chemical Equation

Bulletin

Satay is a favourite food among Malaysians. Satay is made from pieces of spiced meat skewered on coconut or bamboo skewers and grilled over burning charcoal.

Did you know that the burning of charcoal is a form of chemical reaction? This reaction can be represented by the following chemical equation.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

The symbol 'C' in the equation shows the chemical formula of carbon element in the charcoal. What is a chemical formula? What are the meanings of the other letters and numbers in the above equation?

How do you write the formula of a chemical substance?

What information is found in a chemical equation?

How do you measure the quantity of a chemical substance?

3.1 Relative Atomic Mass and Relative Molecular Mass

Photograph 3.1 Rice

Have you ever tried counting the number of rice grains in a sack of rice? Rice grains cannot be counted because their size is extremely small. Chemists face a similar problem too. As atoms are too small, it is difficult to determine their number and the mass of each atom. How do chemists overcome this problem?

Learning Standard

At the end of the lesson, pupils are able to:

3.1.1 Conceptualise relative atomic mass and relative molecular mass based on the carbon-12 scale

3.1.2 Calculate relative molecular mass and relative formula mass

Relative Atomic Mass, RAM

Chemists use the concept of 'relative atomic mass' by comparing the mass of atom of an element to the mass of atom of another element that is chosen as the standard. Therefore, we do not need to know the actual mass of an atom.

Initially, the hydrogen atom was used as the standard atom because it is the lightest atom. Masses of atoms of all other elements were compared with the hydrogen atom. For example, one carbon atom is as heavy as 12 hydrogen atoms. Hence, the relative atomic mass of carbon is 12 while the relative atomic mass of hydrogen is assigned as one as shown in Figure 3.1.

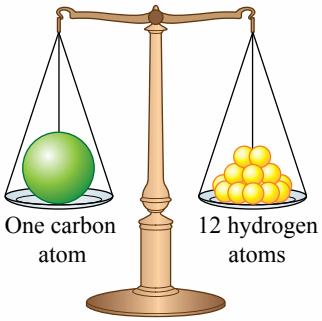


Figure 3.1 *Mass of carbon atom compared to hydrogen atom*

However, in 1961, chemists across the world agreed to use **carbon-12 atom as the standard atom** after finding that the usage of hydrogen atom as the standard atom encountered various problems. The **relative atomic mass, RAM** of an element is defined as the average mass of an atom of the element compared to $\frac{1}{12}$ of the mass of one carbon-12 atom.

$$\text{Relative atomic mass of an element} = \frac{\text{Average mass of one atom of the element}}{\frac{1}{12} \times \text{Mass of one carbon-12 atom}}$$

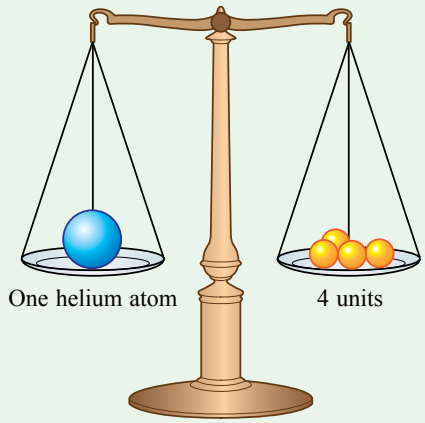
Brain Teaser

Determining the relative atomic mass of hydrogen atom as the standard atom has encountered various problems. Try to investigate what the problems are.

Explain why there is a value of $\frac{1}{12}$ in the definition of the relative atomic mass based on the carbon-12 scale.

One atom of carbon-12 is given a definite mass of 12 units.

So, $\frac{1}{12}$ of the mass of a carbon-12 atom is the same as the mass of one hydrogen atom, that is 1 unit.



The relative atomic mass of helium is 4. This means the average mass of one atom of helium is 4 times the mass of $\frac{1}{12}$ of carbon-12 atom.

Figure 3.2 Relative atomic mass of helium

Brain Teaser

One magnesium atom is twice as heavy as one atom of carbon-12. What is the RAM of magnesium?

Literacy Tips

The relative atomic masses of elements are given in the Data Table of Elements on page 276. Since the relative atomic mass is a comparative value, it has no unit.

Activity 3.1

Discussing why carbon-12 is used as the standard to determine RAM

1. Carry out the activity in groups.
2. Gather information from printed reference materials or surf the Internet and discuss why carbon-12 is used as the standard to determine the relative atomic mass.
3. Present your group discussion results in a suitable thinking map.

Carbon-12 is chosen as the standard because it is a solid at room temperature and thus can be handled easily. Carbon-12 combines easily with other elements. Therefore, this element is found in most substances. Although carbon has three isotopes, carbon-12 is the major isotope with the abundance of 99%. This makes the relative atomic mass of carbon-12 exactly 12.0.

Relative Molecular Mass, RMM

Similarly, we can compare the molecular mass of a substance with the standard carbon-12 atom. The **relative molecular mass, RMM** of a molecule is the average mass of the molecule compared to $\frac{1}{12}$ of the mass of one carbon-12 atom.

$$\text{Relative molecular mass of a substance} = \frac{\text{Average mass of one molecule}}{\frac{1}{12} \times \text{Mass of one carbon-12 atom}}$$

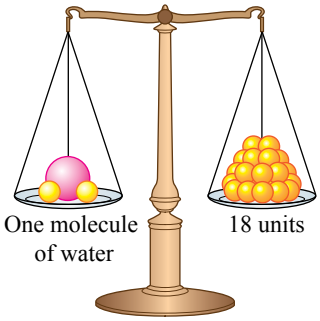


Figure 3.3 Relative molecular mass of water

Figure 3.3 shows water molecule that has a relative molecular mass of 18. This means the mass of a water molecule is 18 times the mass of $\frac{1}{12}$ of carbon-12 atom. Activity 3.2 can strengthen your understanding on the concepts of relative atomic mass and relative molecular mass based on the carbon-12 scale as an analogy.

Activity 3.2

Studying the concepts of relative atomic mass and relative molecular mass by analogy

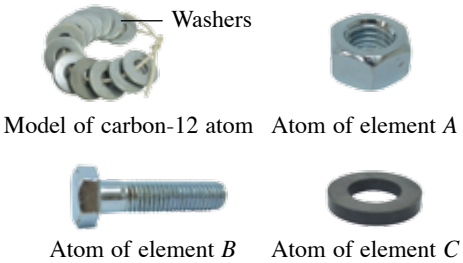
Materials: 36 washers, one 5 cm bolt, five nuts, one flat magnet and strings

Apparatus: Two-pan balance

A Relative atomic mass based on the carbon-12 scale

Procedure:

1. You are given three models of carbon-12 atom. Calculate the number of washers required to form each model of carbon-12 atom.
2. Separate the washers in each model and use them for the following steps.
3. Place an atom of element A on a two-pan balance as shown in Figure 3.4.
4. Place the washers on the other pan one by one until they are balanced.
5. Count and record the number of washers used in Table 3.1.
6. Repeat steps 3 to 5 using atom of element *B* and atom of element *C*.
7. Calculate the relative mass of each washer in the model by assuming that each atom of carbon-12 is given the accurate mass of 12 units. Then, deduce the relative atomic masses of elements *A*, *B* and *C*.



Photograph 3.2 *A model of carbon-12 atom and three elements A, B and C*

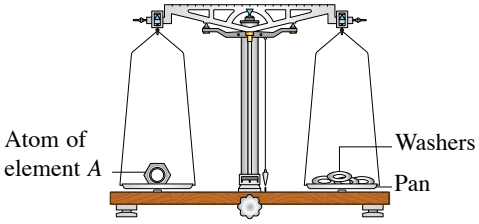


Figure 3.4 *Studying the relative mass by analogy*

Results:

Table 3.1

Atom of element	Number of washers used	Relative atomic mass
A		
B		
C		

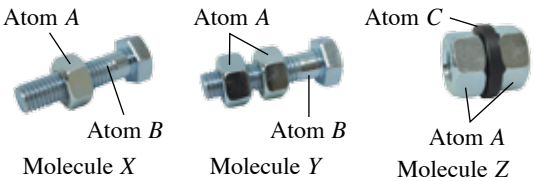
Discussion:

1. How many washers form a model of carbon-12 atom?
2. What is represented by $\frac{1}{12}$ of the mass of carbon-12 atom in this activity?
3. Define the relative atomic mass of an element based on the carbon-12 scale.

B Relative molecular mass based on the carbon-12 scale

Procedure:

1. Prepare models of molecules *X*, *Y* and *Z* as in Photograph 3.3.
2. Place molecule *X* on one of the pans of the balance.
3. Place washers on the other pan one by one until they are balanced.



Photograph 3.3 *Models of molecules X, Y and Z*

4. Count and record the number of washers used in Table 3.2.
5. Repeat steps 2 to 4 using molecule *Y* and *Z*.
6. Deduce the relative molecular masses of *X*, *Y* and *Z*.

Results:

Table 3.2

Molecule	Composition of molecule	Number of washers used	Relative molecular mass
X	1 atom *A* + 1 atom *B*		
Y			
Z			

Discussion:

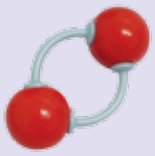
1. Based on Activity B, give the definition of relative molecular mass based on the carbon-12 scale.
2. Calculate the relative atomic masses of all the elements that form molecules *X*, *Y* and *Z*.
3. Compare the answers from question 2 with the relative molecular masses you obtained in Activity B. What inference can you make about the relationship between the relative molecular mass and the relative atomic mass?
4. Molecule *W* is formed from one atom of element *A*, one atom of element *B* and one atom of element *C*. Predict the relative molecular mass of *W*.

Prepare a complete report after carrying out this activity.

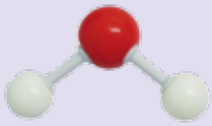
The relative molecular mass of a molecule can be calculated by **adding up the relative atomic masses of all the atoms** that form the molecule, as shown in Figure 3.5 and Example 1.

Literacy Tips

Relative molecular mass of a molecule is similar to the total of relative atomic mass of all atoms in the molecule.



RMM of oxygen gas, O_2
= 2(RAM of O)
= 2(16)
= 32



RMM of water, H_2O
= 2(RAM of H) + RAM of O
= 2(1) + 16
= 18

Figure 3.5 *Calculation of the relative molecular mass, RMM*

Example 1

Calculate the relative molecular mass of glucose, $C_6H_{12}O_6$.
[Relative atomic mass: H = 1, C = 12, O = 16]

Solution

RMM of glucose, $C_6H_{12}O_6$ = 6(RAM of C) + 12(RAM of H) + 6(RAM of O)
= 6(12) + 12(1) + 6(16)
= 72 + 12 + 96
= 180

Relative Formula Mass, RFM

The concept of relative mass is also used for ionic substances. The relative mass of an ionic substance is called the **relative formula mass, RFM**. The relative formula mass is calculated by summing up the relative atomic masses of all the atoms shown in the formula of the ionic substance. This is because the mass of an ion does not differ much from the mass of its atom that forms the ion. Check out Example 2.

Example 2

Calculate the relative formula mass of zinc chloride, $ZnCl_2$ and aluminium sulphate, $Al_2(SO_4)_3$.
[Relative atomic mass: O = 16, Al = 27, S = 32, Cl = 35.5, Zn = 65]

Solution

RFM of zinc chloride, $ZnCl_2$ = RAM of Zn + 2(RAM of Cl)
= 65 + 2(35.5)
= 65 + 71
= 136

RFM of aluminium sulphate, $Al_2(SO_4)_3$ = 2(RAM of Al) + 3[RAM of S + 4(RAM of O)]
= 2(27) + 3[32 + 4(16)]
= 54 + 3[96]
= 342

Activity 3.3

Calculating the relative molecular mass and relative formula mass CT

Determine the relative molecular mass or the relative formula mass of each of the following substances. Refer to the Data Table of Elements on page 276 to obtain the relative atomic mass.

1. H_2
2. O_3
3. CO
4. NH_3
5. N_2O_4
6. C_4H_{10}
7. $CuCl_2$
8. $Zn(OH)_2$
9. $K_2Cr_2O_7$
10. $Fe(NO_3)_3$

Activity 3.4

Tic-tac-toe with relative masses CT

Materials: 10 pieces of formula cards and a tic-tac-toe card

1. Carry out the activity in pairs.
2. Each pair is given a tic-tac-toe card and 10 pieces of formula cards. Each card has the formula of a specific substance and its relative mass.
3. Shuffle the cards and put them at the centre of the table with the written side of the cards facing down.
4. The first player will take a piece of card. Without showing it to the second player, the first player will read the formula of the substance on the card to the second player.

Material for Activity 3.4
http://bit.ly/2PeH6a5

5. Referring to the Data Table of Elements on page 276, the second player will calculate the relative mass of the substance and show the answer to the first player. If the answer is correct, the second player is allowed to mark the tic-tac-toe card. If the answer is wrong, the second player will lose the chance to mark the tic-tac-toe card.
6. Repeat steps 3 to 4 with the second player taking the card while the first player calculates the relative mass of the substance.
7. Continue to take turns until one of the players succeeds in marking a complete line vertically, horizontally or diagonally or until all spaces are filled up.

Test Yourself 3.1

1. Define relative atomic mass based on the carbon-12 scale.
2. Refer to the Data Table of Elements on page 276 to get the relative atomic masses.
   (a) How many atoms of lithium are required to equalise the mass of one atom of krypton?
   (b) How many atoms of helium are required to equalise the mass of a silver atom?
3. Calculate the relative molecular mass or the relative formula mass of each of the following substances:
   (a) Methane, CH_4
   (b) Magnesium nitrate, $Mg(NO_3)_2$
   (c) Sulphuric acid, H_2SO_4
   (d) Formic acid, HCOOH

3.2 Mole Concept

Learning Standard

At the end of the lesson, pupils are able to:
3.2.1 Define mole
3.2.2 Interrelate the Avogadro constant, N_A, the number of particles and the number of moles
3.2.3 State the meaning of molar mass
3.2.4 Interrelate the molar mass, mass and the number of moles
3.2.5 State the meaning of molar volume
3.2.6 Interrelate the molar volume, volume of gas and the number of moles
3.2.7 Solve numerical problems involving the number of particles, number of moles, mass of the substances and volume of gases

In our daily lives, we use units such as pairs and dozens to represent the quantity or number of objects. Photograph 3.4, shows the objects that can be quantified using the units pair and dozen. The unit pair represents 2 objects while the unit dozen represents 12 objects.

Photograph 3.4 *Uses of units in daily life*

In the field of chemistry, we use the unit **mole** to measure the amount of substance. What is the amount of substance represented by the unit mole?

According to the International Union of Pure and Applied Chemistry (IUPAC), the new definition of mole is as follows:

> The **mole**, with the symbol mol, is the SI unit of amount of substance. One mole of substance contains $6.02214076 \times 10^{23}$ elementary entities of the substance. This number is a fixed value known as the Avogadro constant, N_A that is expressed in mol^{-1}. The Avogadro constant, N_A is also called the Avogadro number.

The Avogadro constant, N_A is defined as the number of particles contained in one mole of substance, that is 6.02×10^{23} mol^{-1}. In other words, **1 mol of a substance contains 6.02×10^{23} particles** that form the substance. The type of particles depends on the type of substance, namely atomic substance, molecular substance or ionic substance as shown in Figure 3.6.

MATHEMATICS INTEGRATION

For calculation at this level, Avogadro constant, N_A is taken as 6.02×10^{23} mol^{-1}, to three significant figures.

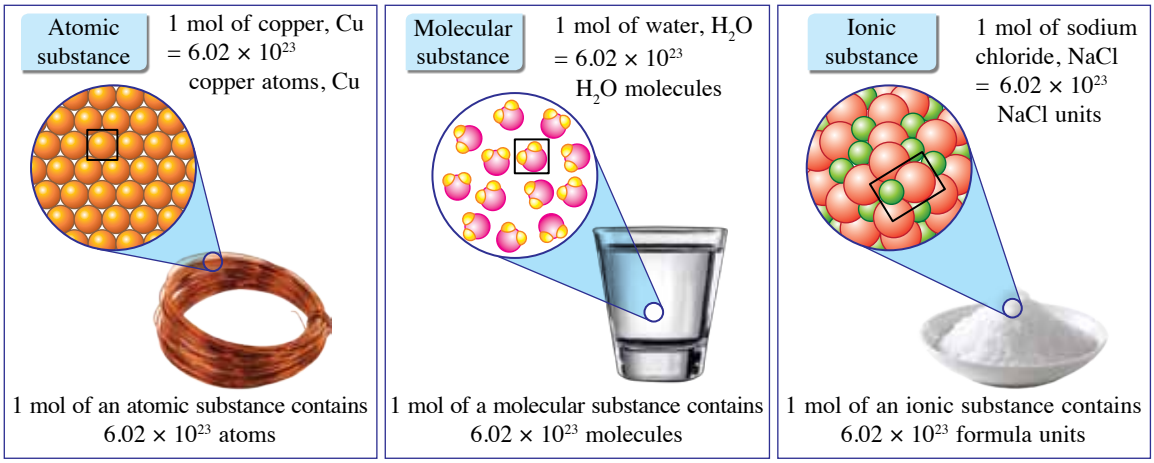


Figure 3.6 *Numbers of particles in 1 mol of substance*

Number of Moles and Number of Particles

The way to use mole is similar to the way of using the unit dozen. For example, 2 dozens of pencils represent 2 × 12 or 24 pencils. Similarly, the Avogadro constant, N_A is used as the conversion factor between the number of moles and the number of particles.

$$\text{Number of moles, } n = \frac{\text{Number of particles}}{N_A}$$

Diagrammatically, the relationship between the number of mole and the number of particles by using Avogadro constant as the conversion factor is shown below:

$$\text{Number of moles} \underset{\div N_A}{\overset{\times N_A}{\rightleftarrows}} \text{Number of particles}$$

HISTORY INTEGRATION

The Avogadro constant is named after a famous Italian scientist, Amedeo Avogadro (1776 – 1856).

The example of conversion between the number of mole and the number of particles by using Avogadro constant, N_A are shown in Examples 3, 4 and 5.

[Note: Assume Avogadro constant, N_A: 6.02×10^{23} mol^{-1}]

Example 3

How many atoms are there in 0.2 mol of magnesium, Mg?

Solution

Number of magnesium atoms, Mg $= 0.2 \text{ mol} \times 6.02 \times 10^{23} \text{ mol}^{-1}$
$= 1.204 \times 10^{23}$ atoms

Use the equation:
Number of particles
= Number of moles × N_A

Example 4

A sample of zinc chloride, $ZnCl_2$ contains 3.01×10^{24} $ZnCl_2$ units. Calculate the number of moles of zinc chloride, $ZnCl_2$ found in the sample.

Solution

Number of moles of zinc chloride, $ZnCl_2$ $= \dfrac{3.01 \times 10^{24}}{6.02 \times 10^{23} \text{ mol}^{-1}}$
$= 5$ mol

Use the equation:
Number of moles
$= \dfrac{\text{Number of particles}}{N_A}$

Example 5

A gas jar is filled with 2 mol of oxygen gas, O_2.
(a) How many molecules of oxygen are there in the gas jar?
(b) How many atoms of oxygen are there in the gas jar?

Solution

(a) Number of oxygen molecules, O_2 $= 2 \text{ mol} \times 6.02 \times 10^{23} \text{ mol}^{-1}$
$= 1.204 \times 10^{24}$ molecules

Use the equation:
Number of particles
= Number of moles × N_A

(b) Each oxygen molecule, O_2 has 2 oxygen atoms, O.
Hence, the number of oxygen atoms, O = Number of O_2 molecules × 2
$= 1.204 \times 10^{24} \times 2$
$= 2.408 \times 10^{24}$ atoms

Activity 3.5

Calculating the number of moles and number of particles CT

[Avogadro constant, N_A: 6.02×10^{23} mol^{-1}]

1. Calculate the number of atoms found in
   (a) 0.1 mol of carbon, C
   (b) 3.5 mol of neon gas, Ne
2. Calculate the number of molecules found in
   (a) 1.2 mol of hydrogen gas, H_2
   (b) 0.8 mol of ammonia, NH_3
3. Calculate the number of formula units found in
   (a) 3 mol of sodium chloride, NaCl
   (b) 0.25 mol of potassium nitrate, KNO_3

4. Calculate the number of moles of each of the following substances:
   (a) 6.02×10^{24} lead atoms, Pb
   (b) 3.02×10^{23} magnesium oxide units, MgO
   (c) 9.03×10^{22} bromine molecules, Br_2
   (d) 3.612×10^{24} carbon dioxide molecules, CO_2
5. A reagent bottle contains 1.806×10^{25} units of copper(II) oxide, CuO.
   (a) How many moles of copper(II) oxide, CuO are found in the bottle?
   (b) Calculate the number of ions found in that bottle.
6. A sample contains 0.2 mol of ethene gas, C_2H_4.
   (a) How many ethene molecules, C_2H_4 are found in the sample?
   (b) How many hydrogen atoms, H are found in the sample?
   (c) Calculate the total number of atoms found in the sample.

Number of Moles and Mass of Substances

The number of moles of a substance is impossible to be determined by counting the number of particles in the substance. Therefore, to get the number of moles, the mass of a substance must be measured and we also need to know its molar mass. What is molar mass?

Molar mass is the mass of one mole of substance.

The unit for molar mass is gram/mol or g mol^{-1}. Chemists found that the **value of molar mass of a substance is the same as its relative mass**. For example, the relative atomic mass of carbon, C is 12. Thus, the molar mass of carbon is 12 g mol^{-1} because 12 g of carbon, C contains 1 mol of carbon, C, that is 6.02×10^{23} atoms of carbon, C. Look at Figure 3.7 to strengthen your understanding.

- Copper consists of copper atoms.
- RAM of copper = 64
- Molar mass of copper = 64 g mol^{-1}

- Water consists of H_2O molecules.
- RMM of water = 2(1) + 16 = 18
- Molar mass of water = 18 g mol^{-1}

- Sodium chloride consists of NaCl units.
- RFM of sodium chloride = 23 + 35.5 = 58.5
- Molar mass of sodium chloride = 58.5 g mol^{-1}

Figure 3.7 *Determining the molar mass of substances*

The mass of any fraction of a mole of a substance can be weighed. For example, 12 g of carbon for 1 mol of carbon, 6 g of carbon powder for 0.5 mol of carbon and so on. The molar mass is used as the conversion factor between the number of moles and the mass of substance. The formula and the relationship between the number of moles and the mass of substance by using molar mass as the conversion factor is as follows:

$$\text{Number of moles, } n = \frac{\text{Mass (g)}}{\text{Molar mass (g mol}^{-1})}$$

Number of moles $\xrightarrow{\times \text{ Molar mass}}$ Mass (g)

Mass (g) $\xrightarrow{\div \text{ Molar mass}}$ Number of moles

PHYSICS INTEGRATION

Why is g used as the unit for mass in the formula?

Mass
= Number of moles × Molar mass
$= \cancel{\text{mol}} \times \frac{\text{g}}{\cancel{\text{mol}}}$
= g

Examples of the conversion between the number of moles and the mass of the particles using molar mass are shown in Examples 6, 7 and 8.

Example 6

What is the mass of 1.5 mol of aluminium, Al?
[Relative atomic mass: Al = 27]

Solution

Molar mass of aluminium, Al = 27 g mol^{-1}

Value of the molar mass of an atomic substance is equal to RAM.

Mass of aluminium, Al = 1.5 mol × 27 g mol^{-1}
= 40.5 g

Use the formula:
Mass = Number of moles × Molar mass

Example 7

How many moles of molecules are found in 32 g of sulphur dioxide gas, SO_2?
[Relative atomic mass: O = 16, S = 32]

Solution

Relative molecular mass of sulphur dioxide, SO_2 = 32 + 2(16)
= 64

Thus, the molar mass of sulphur dioxide, SO_2 = 64 g mol^{-1}

Value of the molar mass of a molecular substance is equal to RMM.

Number of moles of sulphur dioxide molecules, $SO_2 = \frac{32 \text{ g}}{64 \text{ g mol}^{-1}}$
= 0.5 mol

Use the formula:
Number of moles = $\frac{\text{Mass}}{\text{Molar mass}}$

Example 8

How many moles are found in 4.7 g of potassium oxide, K_2O?
[Relative atomic mass: O = 16, K = 39]

Solution

Relative formula mass of potassium oxide, K_2O = 2(39) + 16
= 94

Thus, the molar mass of potassium oxide, K_2O = 94 g mol^{-1}

Value of the molar mass of an ionic substance is equal to RFM.

Number of moles of potassium oxide, $K_2O = \frac{4.7 \text{ g}}{94 \text{ g mol}^{-1}}$
= 0.05 mol

Use the formula:
Number of moles = $\frac{\text{Mass}}{\text{Molar mass}}$

Activity 3.6

Calculating the number of moles and mass

CT

[Relative atomic mass: H = 1, C = 12, N = 14, O = 16, Mg = 24, S = 32, Fe = 56;
Avogadro constant, N_A: 6.02×10^{23} mol^{-1}]

1. Calculate the mass of each of the following substances:
   (a) 0.4 mol of iron fillings, Fe
   (b) 2.2 mol of carbon monoxide, CO
2. Calculate the number of moles in each of the following substances:
   (a) 49 g of sulphuric acid, H_2SO_4
   (b) 8.88 g of magnesium nitrate, $Mg(NO_3)_2$

3. An experiment requires 0.05 mol of ammonium sulphate crystals, $(NH_4)_2SO_4$. What is the mass of ammonium sulphate, $(NH_4)_2SO_4$ that should be used?
4. 0.2 mol of substance *Y* has the mass of 11 g. What is the molar mass of substance *Y*?

Number of Moles and Volume of Gases

Measuring the volume of a gas is easier compared to measuring its mass because gas is very light. How are the number of moles and the volume of a gas related?

From studies, chemists found that the volume of 1 mol of any gas has similar value under the same conditions of temperature and pressure. Thus, the concept of molar volume was explained.

Molar volume is the volume occupied by 1 mol of a gas. The molar volume of any gas depends on the condition, that is 22.4 dm^3 mol^{-1} at STP or 24 dm^3 mol^{-1} at room conditions.

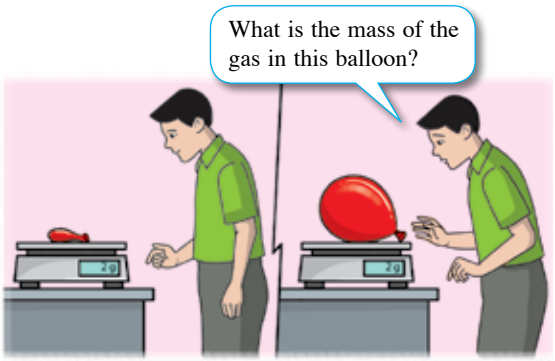
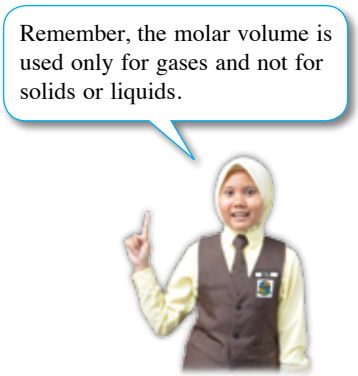


Figure 3.8 *Weighing the mass of a gas*

This means at STP,
- 1 mol of neon gas, Ne occupies 22.4 dm^3
- 1 mol of nitrogen dioxide gas, NO_2 occupies 22.4 dm^3

While at room conditions,
- 1 mol of neon gas, Ne occupies 24 dm^3
- 1 mol of nitrogen dioxide gas, NO_2 occupies 24 dm^3



How do we use the molar volume to measure the number of moles of a gas? The formula and relationship between the number of moles and the volume of gas by using molar volume as a conversion factor are as follow:

$$\text{Number of moles, } n = \frac{\text{Volume of gas}}{\text{Molar volume}}$$

Number of moles $\xrightarrow{\times \text{ Molar volume}}$ Volume of gas

Volume of gas $\xrightarrow{\div \text{ Molar volume}}$ Number of moles

Chemistry Lens

- STP is the abbreviation for standard temperature and pressure, the condition where temperature is at 0 °C and pressure of 1 atm.
- Room conditions refer to the condition where temperature is at 25 °C and pressure of 1 atm.

The conversion between the number of moles and the volume of gas using molar volume are shown in Examples 9, 10 and 11.

Example 9

Calculate the volume of 2.2 mol of hydrogen gas, H_2 in dm^3 at STP.
[Molar volume = 22.4 dm^3 mol^{-1} at STP]

Solution

Volume of hydrogen gas, H_2
= Number of moles × Molar volume at STP
= 2.2 mol × 22.4 dm^3 mol^{-1}
= 49.28 dm^3

PHYSICS INTEGRATION

Volume of gas
= Number of moles × Molar volume
$= \cancel{mol} \times \frac{dm^3}{\cancel{mol}}$
= dm^3

Example 10

What is the volume of 0.01 mol of ammonia gas, NH_3 in cm^3 at room conditions?
[Molar volume = 24 dm^3 mol^{-1} at room conditions]

Solution

Volume of ammonia gas, NH_3 = Number of moles × Molar volume at room conditions
= 0.01 mol × 24 dm^3 mol^{-1}
= 0.24 dm^3
= 0.24 × 1 000 cm^3 ← Convert unit of volume: 1 dm^3 = 1 000 cm^3
= 240 cm^3

Example 11

How many moles of oxygen gas, O_2 has the volume of 600 cm^3 at room conditions?
[Molar volume = 24 dm^3 mol^{-1} at room conditions]

Solution

Volume of oxygen gas, O_2 = 600 cm^3
$= \frac{600}{1\,000}$ dm^3 ← Convert unit of volume: 1 dm^3 = 1 000 cm^3
= 0.6 dm^3

Number of moles of oxygen gas, O_2 $= \frac{\text{Volume of gas}}{\text{Molar volume at room conditions}}$
$= \frac{0.6\ dm^3}{24\ dm^3\ mol^{-1}}$
= 0.025 mol

Alternative solution
http://bit.ly/2MyrUmi

Activity 3.7

Calculating the number of moles and volume of gases CT

[Molar volume of gas = 22.4 dm^3 mol^{-1} at STP or 24 dm^3 mol^{-1} at room conditions]

1. Calculate the volume of 0.6 mol of chlorine gas, Cl_2 at STP and room conditions.
2. Calculate the number of moles of each of the following gases:
   (a) 48 cm^3 of argon gas, Ar at room conditions (b) 39.2 dm^3 of carbon dioxide gas, CO_2 at STP
3. A sample contains 0.2 mol of methane gas, CH_4 and 0.3 mol of ethane gas, C_2H_6. What is the volume of the sample at room conditions?

Activity 3.8

Building a chart showing the relationship between the number of particles, number of moles, mass of substances and volume of gases at STP and room conditions

1. Carry out the activity in groups.
2. Discuss among the group members and build a chart on a flip chart paper that shows the relationship between the number of moles, number of particles, mass of substances and volume of gases.
3. Each member needs to copy the chart onto a small, pocket-sized card to produce a memory card.
4. Use this memory card to solve all the following numerical problems.

The relationship between the number of moles, number of particles, mass of substances and volume of gases is shown in Figure 3.9.

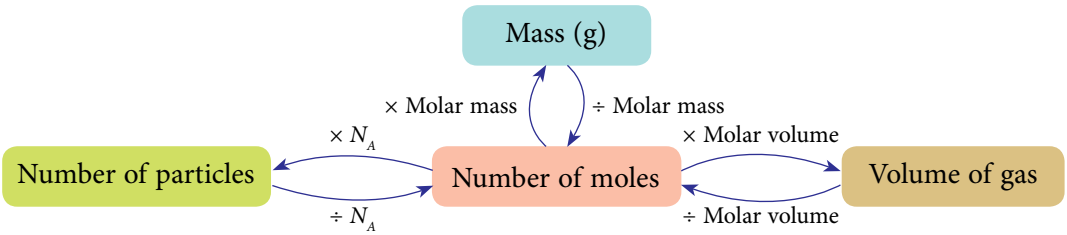


Figure 3.9 *Relationship between the number of moles, number of particles, mass and volume of gases*

Examples 12 and 13 show the function of the number of moles as a medium to convert from one quantity to another.

Example 12

What is the volume of 26.4 g of carbon dioxide, CO_2 at room conditions?
[Relative atomic mass: C = 12, O = 16; Molar volume of gas = 24 $dm^3\ mol^{-1}$ at room conditions]

Solution

Question analysis and solution plan

Information from the question: Mass = 26.4 g → Volume of gas at room conditions?

Solution plan: ÷ Molar mass → Number of moles → × Molar volume

Step 1 | Step 2

RMM of carbon dioxide, CO_2 = 12 + 2(16)
= 44

Thus, the molar mass of carbon dioxide, CO_2 = 44 g mol^{-1} ← Before carrying out step 1, the molar mass must first be determined.

Number of moles of carbon dioxide, CO_2 = $\frac{\text{Mass}}{\text{Molar mass}}$ ← Step 1: Mass → Number of moles

$= \frac{26.4\ g}{44\ g\ mol^{-1}}$

= 0.6 mol

Volume of carbon dioxide, CO_2 = Number of moles × Molar volume
= 0.6 mol × 24 dm^3 mol^{-1}
= 14.4 dm^3

Step 2:
Number of moles → Volume

Hence, 26.4 g of carbon dioxide gas, CO_2 occupies a volume of 14.4 dm^3 at room conditions.

Example 13

How many molecules are there in 672 cm^3 of hydrogen gas, H_2 at STP?
[Avogadro constant, N_A: 6.02×10^{23} mol^{-1}; Molar volume = 22.4 dm^3 mol^{-1} at STP]

Solution

Question analysis and solution plan

Information from the question: Volume of gas = 672 cm^3 → Number of molecules?

Solution plan: ÷ Molar volume → Number of moles → × N_A

Step 1 | Step 2

Number of moles of hydrogen gas, H_2 = $\frac{\text{Volume of gas}}{\text{Molar volume}}$

$= \frac{672\ cm^3}{22.4 \times 1\ 000\ cm^3\ mol^{-1}}$

= 0.03 mol

Step 1:
Volume → Number of moles

Number of hydrogen molecules, H_2 = Number of moles × N_A
= 0.03 mol × 6.02×10^{23} mol^{-1}
= 1.806×10^{22} molecules

Step 2:
Number of moles →
Number of molecules

Hence, 672 cm^3 of hydrogen gas, H_2 at STP consists of 1.806×10^{22} molecules.

Further example
http://bit.ly/2MBDA7Z

Activity 3.9

Solving problems involving the number of particles, number of moles, mass of substances and volume of gases at STP or room conditions

CT

1. Carry out this activity in groups.
2. Read and answer the following questions.
   [Relative atomic mass: H = 1, He = 4, C = 12, N = 14, O = 16, Al = 27, S = 32; Avogadro constant, N_A: 6.02×10^{23} mol^{-1}; Molar volume = 22.4 dm^3 mol^{-1} at STP or 24 dm^3 mol^{-1} at room conditions]
   (a) Calculate the number of atoms found in the following substances:
   (i) 6.75 g of aluminium, Al (ii) 5.1 g of ammonia gas, NH_3
   (b) Calculate the volume of the following gases at STP.
   (i) 5.6 g of nitrogen gas, N_2 (ii) 1.204×10^{22} helium atoms, He
   (c) What is the mass of oxygen gas, O_2 that has the same number of molecules as in 8 g of sulphur trioxide gas, SO_3?
   (d) A sample of methane gas, CH_4 occupies a volume of 9.84 dm^3 at room conditions. How many molecules are found in that sample? Calculate the mass of the sample.
   (e) A reaction releases 120 cm^3 of carbon dioxide gas, CO_2 per minute at room conditions. Calculate the total mass of carbon dioxide, CO_2 released after 10 minutes.
3. Write the calculation steps for questions (a) to (e) clearly and systematically on a piece of flip chart paper.
4. Present your group solution in front of the class.

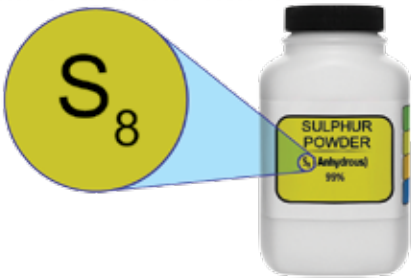
Test Yourself 3.2

[Relative atomic mass: H = 1, C = 12, N = 14, O = 16, Na = 23, Cl = 35.5, K = 39, Fe = 56, Pb = 207; Avogadro constant, N_A: 6.02×10^{23} mol^{-1}; Molar volume = 22.4 dm^3 mol^{-1} at STP or 24 dm^3 mol^{-1} at room conditions]

1. Calculate the molar mass of each of the following substances:
   (a) Lead metal, Pb (c) Sodium nitrate, $NaNO_3$
   (b) Chloroform, $CHCl_3$ (d) Iron(III) oxide, Fe_2O_3
2. Calculate the number of molecules found in 8 mol of water.
3. What is the mass of 0.5 mol of ammonia, NH_3?
4. How many moles of K_2O units are found in 14.1 g of potassium oxide, K_2O?
5. Calculate the volume of 16 g of oxygen gas, O_2 at STP.
6. The mass of 4 dm^3 of a gas is 12 g at room conditions. Calculate the molar mass of the gas.
7. 4 g of hydrogen gas, H_2 has greater number of molecules than 14 g of nitrogen gas, N_2

   Do you agree with the above statement? Give your reason.

3.3 Chemical Formula



Photograph 3.5 *A chemical formula represents a chemical substance*

Chemical formula is a representation of a chemical substance using alphabets to represent the atoms and subscript numbers to show the number of each type of atoms found in the elementary entities of the substance.

Examples of chemical formulae of elements and compounds are shown in Figure 3.10.

Elements

Substance: Magnesium Chemical formula: Mg	The chemical formula shows that magnesium consists of magnesium atoms only.
Substance: Oxygen gas Chemical formula: O_2	The chemical formula shows that oxygen gas molecule consists of two oxygen atoms.

Compounds

Substance: Water Chemical formula: H_2O	The subscript number shows that two atoms of hydrogen combine with one atom of oxygen.
Substance: Aluminium oxide Chemical formula: Al_2O_3	The subscript number shows that two atoms of aluminium combine with three atoms of oxygen.

Figure 3.10 *Chemical formulae of elements and compounds*

Learning Standard

At the end of the lesson, pupils are able to:

3.3.1 State the meaning of chemical formula, empirical formula and molecular formula

3.3.2 Determine the empirical formula of magnesium oxide, MgO through an activity

3.3.3 Determine the empirical formula of copper(II) oxide, CuO through an activity

3.3.4 Solve numerical problems involving empirical formula and molecular formula

3.3.5 Construct chemical formulae of compounds

Chemistry Lens

Elements are substances that consist of only one type of atoms. Elements like metals and inert gases are atomic substances while elements such as oxygen gas are molecular substances.

Literacy Tips

The subscript number 1 need not be written in a chemical formula.

Empirical Formula and Molecular Formula

In general, compounds can be represented by two types of chemical formulae, namely the empirical formula and the molecular formula. What are the empirical formula and the molecular formula?

Activity 3.10

CT

Gathering and interpreting information involving chemical formulae, empirical formulae and molecular formulae

1. Carry out this activity in groups.
2. Gather information on chemical formulae, empirical formulae and molecular formulae by referring to reading materials or surfing the Internet.
3. Based on the information gathered, construct a suitable thinking map to show the difference between the empirical formula and the molecular formula using a suitable computer software.
4. List out the examples of chemical formulae in a table and use this list throughout your lesson.

The **empirical formula** is the chemical formula that shows the **simplest ratio** of the number of atoms of each element in a compound. The **molecular formula**, on the other hand, is the chemical formula that shows the **actual number** of atoms of each element found in a molecule of a compound. Figure 3.11 shows the difference between the empirical formula and the molecular formula.

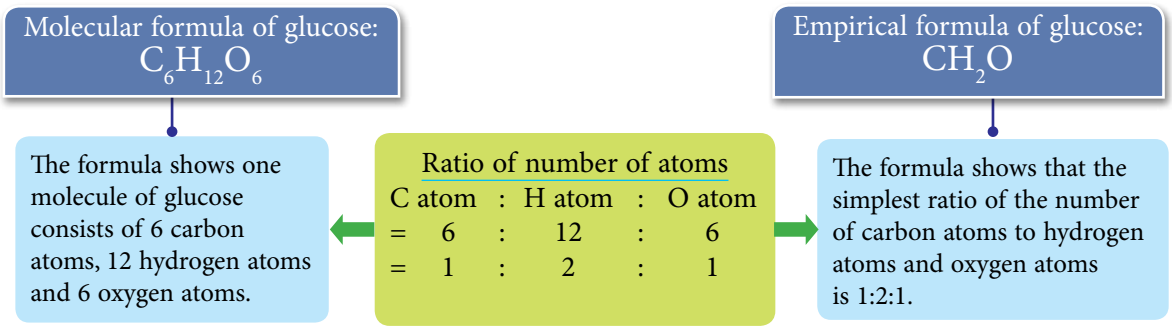


Figure 3.11 *Molecular formula and empirical formula of glucose*

Table 3.3 *Molecular formulae and empirical formulae of several substances*

Substance	Molecular formula	Empirical formula
Water	H_2O	H_2O
Ammonia	NH_3	NH_3
Hydrazine	N_2H_4	NH_2
Propene	C_3H_6	CH_2
Benzene	C_6H_6	CH

Brain Teaser

Some compounds have the same empirical formula and molecular formula. However, there are some other compounds that have different empirical formula and molecular formula. Try to think why it is so.

Determination of an Empirical Formula

The empirical formula is obtained by the analysis of percentage composition of a substance. This is done by determining the **simplest whole number ratio of atoms** of each element that combines through an experiment. Example 14 is used as a guide to solve Activity 3.11.

Brain Teaser

Hexane is an organic solvent that is widely used in the food industry. The molecular formula of hexane is C_6H_{14}. What is its empirical formula?

Example 14

1.35 g of aluminium combines with 1.2 g of oxygen to form aluminium oxide.
What is the empirical formula of aluminium oxide?
[Relative atomic mass: O = 16; Al = 27]

Solution

Element	Al	O	
Mass (g)	1.35	1.2	Determine the mass of each element.
Number of moles of atoms	$\frac{1.35}{27} = 0.05$	$\frac{1.2}{16} = 0.075$	$n = \frac{\text{Mass}}{\text{Molar mass}}$
Mole ratio	$\frac{0.05}{0.05} = 1$	$\frac{0.075}{0.05} = 1.5$	Divide each number with the smallest number, that is 0.05.
Simplest mole ratio of atom	2	3	Multiply each answer by 2 to get the simplest whole number ratio.

2 mol of aluminium atoms combine with 3 mol of oxygen atoms.
Thus, the empirical formula of aluminium oxide is Al_2O_3.

Activity 3.11

Determining the empirical formulae CT

[Relative atomic mass: H = 1, C = 12, O = 16, Cl = 35.5, K = 39, Br = 80, Sn = 119, I = 127]

1. A sample of potassium bromide contains 6.24 g of potassium and 12.8 g of bromine. What is the empirical formula of potassium bromide?
2. A sample of 26.1 g of tin chloride contains 11.9 g of tin. State the empirical formula of the tin chloride.
3. 0.03 mol of element *Y* combines with 7.62 g of iodine to produce an iodide salt. State the empirical formula of the iodide salt.
4. A chemist analysed the compound that gives smell to fully ripe bananas. He found that the compound contains 64.62% carbon, 10.77% hydrogen and 24.61% oxygen. What is the empirical formula of that compound?

Photograph 3.6 *Bananas*

Using the calculation skills learned, the determination of the empirical formulae of magnesium oxide and copper(II) oxide can be carried out through Activity 3.12 and 3.13.

Activity 3.12

Aim: To determine the empirical formula of magnesium oxide.

Materials: 10 cm magnesium ribbon and sand paper

Apparatus: Crucible with lid, tongs, Bunsen burner, tripod stand, pipeclay triangle and electronic balance

Procedure:

1. Weigh and record the mass of a crucible together with its lid.
2. Rub 10 cm magnesium ribbon with a sand paper until shiny. Coil the magnesium ribbon and put it in the crucible.
3. Weigh and record the mass of the crucible together with its lid and the coil of magnesium ribbon.
4. Set up the apparatus as shown in Figure 3.12.
5. First, heat the crucible without its lid.
6. When magnesium ribbon starts to burn, close the crucible with its lid.
7. Using a pair of tongs, lift the lid slightly from time to time and quickly place it back.
8. When the burning of magnesium ribbon is complete, take off the lid and heat the crucible with high temperature for 1 to 2 minutes.
9. Put back the lid of the crucible and allow it to cool to room temperature.
10. Weigh the mass of crucible together with its lid and its contents again.
11. Repeat the heating, cooling and weighing process until a constant mass is obtained.
12. Record the constant mass in Table 3.4.

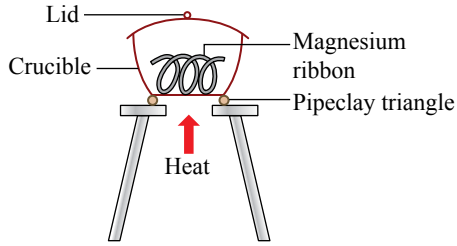


Figure 3.12

Safety Precaution

Prevent white fumes in the crucible from escaping when carrying out Step 7.

Results:

Table 3.4

Description	Mass (g)
Crucible + lid	
Crucible + lid + magnesium ribbon	
Crucible + lid + magnesium oxide	

Interpreting data:

1. Based on your results, determine the masses of magnesium and oxygen that combine.
2. Determine the empirical formula of magnesium oxide.

Discussion:

1. What is the purpose of rubbing the magnesium ribbon with a sand paper before using it?
2. Name the white fumes that are produced.
3. Why are steps 6, 7 and 11 performed?
4. What will happen if the white fumes are released into the environment?

Prepare a complete report after carrying out this activity.

Activity 3.13

Aim: To determine the empirical formula of copper(II) oxide.

Materials: Water, copper(II) oxide powder, zinc granules, 1.0 mol dm^{-3} hydrochloric acid, wooden splinter and cotton buds

Apparatus: Boiling tube, rubber stoppers, rubber tube, 12 cm glass tube, 10 cm glass tube, spirit lamp, retort stand with clamp, wooden block, electronic balance and spatula

Procedure:

1. Weigh the mass of 12 cm glass tube using an electronic balance and record its mass.
2. Put some copper(II) oxide powder into the glass tube. Use the wooden splinter to move copper(II) oxide powder to the middle of the glass tube. Weigh the mass of the glass tube together with its contents and record the mass.
3. Fill $\frac{2}{3}$ of the boiling tube with water.
4. Close the boiling tube with a rubber stopper that has a 12 cm glass tube. Clamp the boiling tube onto the retort stand.
5. Insert a few zinc granules into another boiling tube. Add 1.0 mol dm^{-3} hydrochloric acid into the boiling tube until it is $\frac{1}{3}$ full.
6. Close the boiling tube with a rubber stopper that has a 10 cm glass tube. Clamp the boiling tube onto the other retort stand.
7. Connect the glass tube that contains copper(II) oxide powder as shown in Figure 3.13.

Guideline to determine the empirical formula of copper(II) oxide

http://bit.ly/2VLQHq6

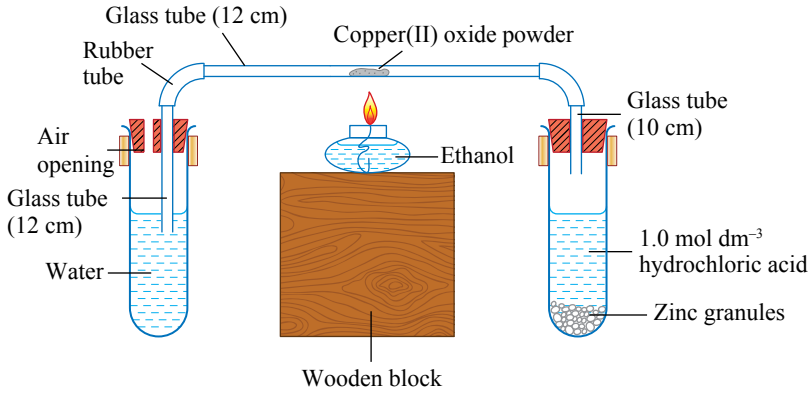


Figure 3.13

8. Let the hydrogen gas flow for 10 seconds by allowing the air bubbles to be released in the water before starting the heating process.
9. Heat copper(II) oxide using a spirit lamp with a continuous flow of hydrogen gas through the glass tube.
10. Stop the heating when the black colour of copper(II) oxide turns brown completely.
11. Keep a continuous flow of hydrogen gas until the glass tube is cooled back to room temperature.

Safety Precaution

If necessary, hold the spirit lamp by moving it under the glass tube to heat the remaining powder that is still black so that all black powder turns brown.

12. Remove the glass tube that contains brown powder. Eliminate water drops at the end of the glass tube with a cotton bud.
13. Weigh the mass of the glass tube together with its contents and record its mass.
14. Repeat the heating, cooling and weighing processes from steps 9 to 13 until a constant mass reading is obtained.
15. Record the constant mass in Table 3.5.

Another method of determining the empirical formula of copper(II) oxide

http://bit.ly/2BeHBbY

Results:

Table 3.5

Description	Mass (g)
Glass tube	
Glass tube + copper(II) oxide	
Glass tube + copper	
Copper	
Oxygen	

Interpreting data:

1. Determine the empirical formula of copper(II) oxide in this activity.

Discussion:

1. What is the purpose of using zinc granules and hydrochloric acid in this activity?
2. Why does the hydrogen gas need to flow continuously for a while before starting the heating process?
3. The hydrogen gas is allowed to flow until the product of heating is at room temperature in step 11. Why?
4. Why do the heating, cooling and weighing processes need to be repeated until a constant mass is obtained?

Prepare a complete report after carrying out this activity.

For reactive metals like magnesium, the metal needs to be heated only slightly before it can react with the oxygen in the air. Figure 3.14 shows how the mass of magnesium and oxygen that combine are calculated to determine the simplest mole ratio of atom.

Can you name another reactive metal oxide which empirical formula can be determined using the same method?

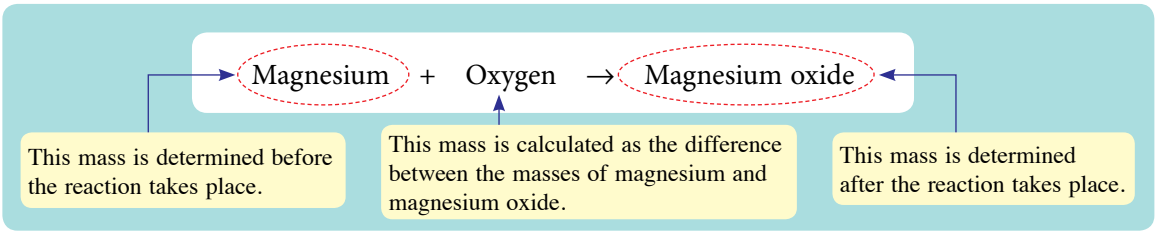


Figure 3.14 *Calculation of the mass of magnesium and oxygen in magnesium oxide*

However, this method is not suitable in determining the empirical formula of copper(II) oxide because copper is less reactive towards oxygen. Hence, copper(II) oxide is heated in a stream of hydrogen gas so that hydrogen can remove oxygen from the oxide as shown in Figure 3.15.

Reactivity series of metals

http://bit.ly/2pFVTQb

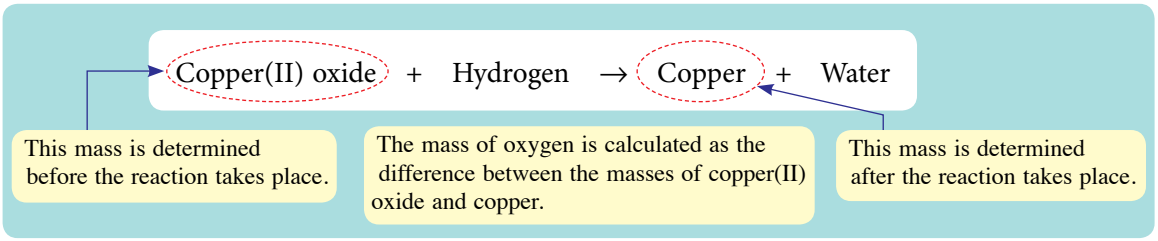


Figure 3.15 *Calculation of the masses of copper and oxygen in copper(II) oxide*

Determination of a Molecular Formula

The molecular formula of a compound is its multiplied empirical formula.

$$\text{Molecular formula} = (\text{Empirical formula})_n$$

The value of n is a positive integer. Table 3.6 shows several examples.

Table 3.6 *Relationship between the molecular formula and the empirical formula*

Substance	Water	Hydrazine	Propene	Benzene
Empirical formula	H_2O	NH_2	CH_2	CH
Molecular formula	$(H_2O)_1 = H_2O$	$(NH_2)_2 = N_2H_4$	$(CH_2)_3 = C_3H_6$	$(CH)_6 = C_6H_6$
n	1	2	3	6

Therefore, to determine the molecular formula of a compound, we first need to know its empirical formula. Examples 15 and 16 show solutions regarding chemical formula.

Example 15

A compound has the empirical formula CH_2. Its relative molecular mass is 56. What is the molecular formula of the compound? [Relative atomic mass: H = 1; C = 12]

Solution

Assume that the molecular formula of the compound is $(CH_2)_n$.

Based on its molecular formula, the RMM of compound $= n[12 + 2(1)]$

$= 14n$

Given the RMM of compound, $14n = 56$

$n = \frac{56}{14}$

$= 4$

Equate the calculated RMM with the given one.

Hence, the molecular formula of the compound is C_4H_8.

Example 16

1.2 g of element Y reacts with bromine to form 6 g of a compound with the empirical formula YBr_2. Determine the relative atomic mass of Y. [Relative atomic mass: Br = 80]

Solution

The compound consists of element Y and bromine.

Therefore, the mass of element Y + mass of bromine = mass of compound formed

1.2 g + mass of bromine = 6 g

Mass of bromine = (6 – 1.2) g

= 4.8 g

Assume that the RAM of element Y is x.

Element	Y	Br
Mass (g)	1.2	4.8
Number of moles of atoms	$\frac{1.2}{x} = ?$	$\frac{4.8}{80} = 0.06$

Further example

http://bit.ly/32BiQ5J

Based on the empirical formula YBr_2,

2 mol of Br atoms combine with 1 mol of atom Y or

1 mol of Br atoms combine with 0.5 mol of atom Y or

0.06 mol of Br atoms combine with 0.03 mol of atom Y.

Based on the empirical formula, calculate using the right ratio.

Hence, the number of moles of atom Y that reacts = 0.03 mol

$$\frac{1.2}{x} = 0.03$$

$$x = \frac{1.2}{0.03}$$

$$= 40$$

The RAM of element Y is 40.

Activity 3.14

Solving numerical problems involving empirical formulae and molecular formulae

CT

[Relative atomic mass: H = 1, C = 12, N = 14, O = 16, Ca = 40, Zn = 65]

1. Ethanoic acid has a molar mass of 60 g mol^{-1}. If its empirical formula is CH_2O, determine the molecular formula of ethanoic acid.
2. Hydrocarbons consist of carbon and hydrogen. 5.7 g of a hydrocarbon contains 4.8 g of carbon. If the relative molecular mass of the hydrocarbon is 114, determine its molecular formula.
3. What is the mass of zinc required to combine with 0.5 mol of chlorine to produce zinc chloride, $ZnCl_2$?
4. Assume you are a farmer. You want to choose a fertiliser with a high nitrogen content for your plants. Three types of commonly used fertilisers are as follows.

   Ammonium nitrate, NH_4NO_3 | Urea, $CO(NH_2)_2$ | Nitrosol, $Ca(NO_3)_2$

   Which fertiliser would you choose? Give reasons for your choice. Show the steps used in the calculation.

Chemical Formulae of Ionic Compounds

Ionic compounds are made up of cations (positively-charged ions) and anions (negatively-charged ions). In order to write the chemical formulae of ionic compounds, you need to know the formulae of cations and anions. Table 3.7 shows the examples of formulae for cations and anions that are commonly used. Figure 3.16 explains how the chemical formula of an ionic compound is constructed.

Table 3.7 *Formulae of common cations and anions*

Cation	Formula of cation	Anion	Formula of anion
Sodium ion	Na^+	Oxide ion	O^{2-}
Potassium ion	K^+	Chloride ion	Cl^-
Aluminium ion	Al^{3+}	Bromide ion	Br^-
Zinc ion	Zn^{2+}	Iodide ion	I^-
Magnesium ion	Mg^{2+}	Hydroxide ion	OH^-
Iron(II) ion	Fe^{2+}	Carbonate ion	CO_3^{2-}
Iron(III) ion	Fe^{3+}	Nitrate ion	NO_3^-
Copper(II) ion	Cu^{2+}	Sulphate ion	SO_4^{2-}
Calcium ion	Ca^{2+}	Phosphate ion	PO_4^{3-}
Silver ion	Ag^+	Manganate(VII) ion	MnO_4^-
Lead(II) ion	Pb^{2+}	Thiosulphate ion	$S_2O_3^{2-}$
Ammonium ion	NH_4^+	Dichromate(VI) ion	$Cr_2O_7^{2-}$

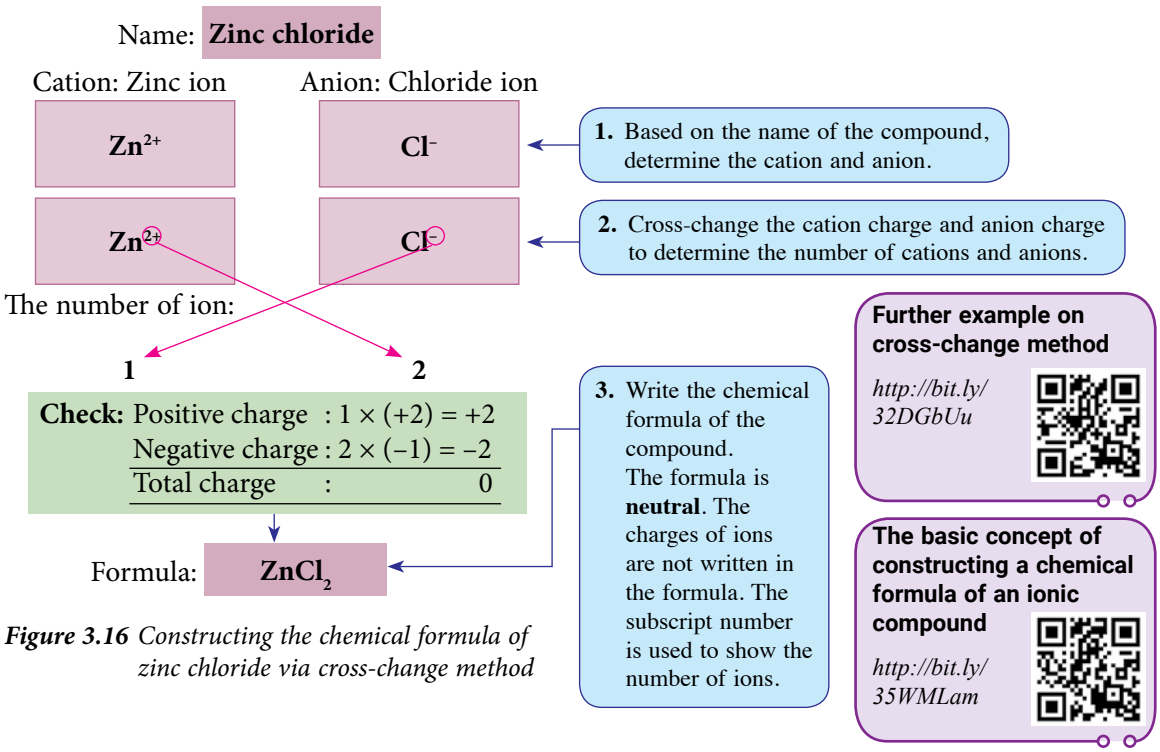


Figure 3.16 *Constructing the chemical formula of zinc chloride via cross-change method*

Activity 3.15

Constructing the chemical formulae of ionic compounds

CT

1. Carry out this activity individually.
2. Scan the QR code and download the diagram of ionic formula cards.
3. Print and cut out the ionic formula cards.
4. Use the ion formula cards to help you determine the chemical formula of each of the following ionic compounds:

Diagram of ionic formula cards
http://bit.ly/2N4JVaG

Potassium oxide	Sodium hydroxide	Magnesium nitrate	Calcium nitrate
Sodium chloride	Aluminium oxide	Potassium carbonate	Aluminium chloride
Calcium bromide	Zinc sulphate	Copper(II) sulphate	Sodium carbonate

5. Record your answers systematically in a table.

Naming of Chemical Compounds

For ionic compounds, the name of the cation is written first followed by the name of the anion as in Table 3.8.

Table 3.8 *Examples in the naming of ionic compounds*

Cation	Anion	Name of ionic compound
Sodium ion	Chloride ion	Sodium chloride
Zinc ion	Bromide ion	Zinc bromide
Magnesium ion	Nitrate ion	Magnesium nitrate

Chemistry Lens

Chemical compounds are named systematically as recommended by the International Union of Pure and Applied Chemistry (IUPAC).

Some metals form more than one type of ions. In order to distinguish these ions, Roman numerals are used in their naming. For example, iron forms two types of cations, namely Fe^{2+} and Fe^{3+}. Fe^{2+} ion is named as iron(II) ion while Fe^{3+} ion is named as iron(III) ion. Take a look at the names of the following compounds:

Shows iron(II) ion, Fe^{2+} — Iron(II) oxide

Iron(III) oxide — Shows iron(III) ion, Fe^{3+}

When naming simple molecular compounds, the more electropositive element is named first followed by the name of the more electronegative element. The name of the first element remains the same while the second element ends with 'ide'. Greek prefixes are used to represent the number of atoms of each element in simple molecular compounds. Look at the examples below.

CO – Carbon monoxide
NO_2 – Nitrogen dioxide
SO_3 – Sulphur trioxide

Greek prefixes like 'mono', 'di' and 'tri' show the numbers one, two and three respectively.

Literacy Tips

Other Greek prefixes are as follows:

tetra – 4	hex – 6
pent – 5	hept – 7

Activity 3.16

Naming compounds CT

1. Name the ionic compounds with the following formulae:
   (a) $CaCl_2$ (c) $Mg(NO_3)_2$ (e) Na_2SO_4
   (b) KBr (d) $ZnCO_3$ (f) NH_4Cl
2. Name the molecular compounds with the following formulae:
   (a) NO (c) SO_3 (e) BF_3
   (b) CO_2 (d) CCl_4 (f) CS_2
3. The molecule of a compound consists of two nitrogen atoms and three oxygen atoms. Name the compound.

Test Yourself 3.3

1. What is meant by empirical formula and molecular formula?
2. Caffeine, $C_8H_{10}N_4O_2$ is a natural stimulant found in coffee, tea and cocoa. What is the empirical formula of caffeine?
3. Calcium carbonate and sodium fluoride are two compounds found in toothpaste. Write the chemical formulae of both compounds.
4. A sample of 5.04 g of oxide for phosphorus contains 2.48 g of phosphorus. [Relative atomic mass: O = 16, F = 31]
   (a) If the relative molecular mass of the oxide is 126, determine its empirical formula and molecular formula.
   (b) Name the oxide of the phosphorus.

3.4 Chemical Equation

Learning Standard

At the end of the lesson, pupils are able to:
3.4.1 Write balanced chemical equations
3.4.2 Interpret chemical equations quantitatively and qualitatively
3.4.3 Solve stoichiometry numerical problems

Photograph 3.7 *Burning of an oil lamp*

Did you know that the burning of fuel and the digestion of food in our bodies are all chemical reactions? Chemists have a simple and accurate way to describe chemical reactions, that is through **chemical equations**.

How to Write Chemical Equations

Chemical equations can be written in the form of words or using chemical formulae. The starting substances or **reactants** are written on the left-hand side of the equation while the new substances formed or **products** are written on the right-hand side of the equation. The arrow '→' means 'produces'. The physical state of each substance, whether solid(s), liquid(l), gas(g) or aqueous solution(aq) is usually indicated in a chemical equation. Figure 3.17 shows the examples of writing the chemical equation for the reaction between hydrogen and oxygen.

Reactants				Product
Hydrogen	+	Oxygen	→	Water
H_2	+	O_2	→	H_2O
H_2 (2 H atoms)	+	O_2 (2 O atoms)	→	H_2O (2 H atoms, 1 O atom)
Equation is not balanced				
$2H_2$ (4 H atoms)	+	O_2 (2 O atoms)	→	$2H_2O$ (4 H atoms, 2 O atoms)
$2H_2(g)$	+	$O_2(g)$	→	$2H_2O(l)$

1. Write the equation in words.
2. Write down the chemical formula of each reactant and product.
3. Check whether the equation is balanced.
4. **Balance** the equation by adjusting the **coefficient** in front of the chemical formula.
5. Write the physical state of each reactant and product.

Figure 3.17 *Writing the chemical equation for the reaction between hydrogen and oxygen*

Chemical equations need to be **balanced**. Based on the law of conservation of mass, matter can neither be created nor destroyed. Therefore, the number of atoms of each element on both sides of the equation must be the same.

Simulation on balancing chemical equation

http://bit.ly/33vr5QQ

Activity 3.17

Balancing chemical equations

CT

1. Write a balanced chemical equation for each of the following reactions:
   (a) Nitrogen gas + Hydrogen gas → Ammonia gas
   (b) Sodium metal + Water → Aqueous solution of sodium hydroxide + Hydrogen gas
   (c) Solid copper(II) carbonate decomposes into solid copper(II) oxide and carbon dioxide gas when heated.
   (d) Burning of aluminium powder in excess oxygen produces white aluminium oxide powder.
2. Balance the following chemical equations:
   (a) $KI(aq) + Br_2(aq) \rightarrow I_2(s) + KBr(aq)$
   (b) $Zn(s) + AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + Ag(s)$
   (c) $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
   (d) $AgNO_3(s) \xrightarrow{\Delta} Ag(s) + NO_2(g) + O_2(g)$

Chemistry Lens

Sometimes, chemical equations also show the condition of the reactions. For example, the Greek letter delta (Δ) below the arrow shows that heating is required in the chemical reaction.

Activity 3.18

Aim: To write balanced chemical equations.

Materials: Copper(II) carbonate powder, $CuCO_3$, concentrated hydrochloric acid, HCl, concentrated ammonia solution, NH_3, lead(II) nitrate solution, $Pb(NO_3)_2$, potassium iodide solution, KI and limewater

Apparatus: Test tubes, delivery tube and rubber stopper, test tube holder, Bunsen burner, 10 cm^3 measuring cyclinder, test tube stoppers and glass tube

CAUTION

Concentrated hydrochloric acid and concentrated ammonia are corrosive. Handle them with care and carry out Activity 3.18 in the fume chamber.

Procedure:

A Heating of copper(II) carbonate, $CuCO_3$

1. Fill a spatula of copper(II) carbonate powder, $CuCO_3$ into a test tube. Observe the colour of the powder.
2. Set up the apparatus as shown in Figure 3.18.
3. Heat copper(II) carbonate, $CuCO_3$ and let the gas produced flow into the test tube filled with limewater. Observe the changes that take place in both test tubes.
4. When the reaction is completed, remove the test tube of limewater. Then, stop the heating.
5. Record your observations.

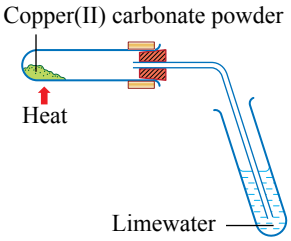


Figure 3.18

B Formation of ammonium chloride, NH_4Cl

1. Using a glass tube, put 3 or 4 drops of concentrated hydrochloric acid, HCl into a test tube. Close the test tube with a stopper and leave it for a few minutes.
2. Repeat step 1 using concentrated ammonia solution, NH_3 in another test tube.
3. Remove the stoppers from both test tubes. Quickly bring the mouths of both test tubes together as shown in Figure 3.19.
4. Observe and record the changes that take place.

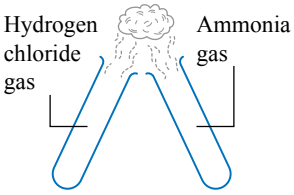


Figure 3.19

C Precipitation of lead(II) iodide, PbI_2

1. Pour 2 cm^3 of lead(II) nitrate solution, $Pb(NO_3)_2$ into a test tube.
2. Pour 2 cm^3 of potassium iodide solution, KI into another test tube.
3. Pour potassium iodide solution, KI into lead(II) nitrate solution, $Pb(NO_3)_2$ as shown in Figure 3.20. Shake the mixture.
4. Observe and record the changes that take place.

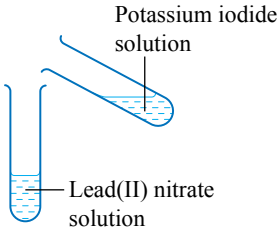


Figure 3.20

Discussion:

1. For each reaction in experiments A, B and C, state:
   (a) The reactants and products
   (b) The physical state of each reactant and product
   (c) The chemical formula of each reactant and product
2. Write a balanced chemical equation for each of the reactions.

Prepare a complete report after carrying out this activity.

Using Chemical Equations

Chemical equation can be interpreted qualitatively and quantitatively. From the **qualitative aspect**, chemical equations enable us to identify the reactants and products as well as their physical states.

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

Reactants: Sodium metal and chlorine gas

Product: Solid sodium chloride

From the **quantitative aspect**, we can study the stoichiometry of chemical equations. Stoichiometry is the quantitative study of the composition of substances involved in a chemical reaction. **Coefficients** in chemical equations show the ratio of substances involved, either as the ratio of elementary entities of substance or the mole ratio. Take a look at the following example:

$2Na(s)$	+	$Cl_2(g)$	$\rightarrow$	$2NaCl(s)$
(2 atoms)		(1 molecule)		(2 formula units)
or		or		or
(2 mol)		(1 mol)		(2 mol)

Ratio of basic entities (particles):
Two sodium atoms react with one molecule of chlorine to produce two NaCl units.

Mole ratio:
2 mol of sodium react with 1 mol of chlorine gas to produce 2 mol of sodium chloride.

Activity 3.19

Interpreting chemical equations qualitatively and quantitatively

21st Century Skills

1. Carry out the Think-Pair-Share activity.
2. Based on the chemical equations obtained from Activity 3.18, interpret each equation qualitatively and quantitatively, from the aspects of ratio of elementary entities and mole ratio.
3. Discuss with your partner.
4. Share the results of your discussion with the class.

Based on the mole ratio of substances from a balanced chemical equation, we can solve various numerical problems by calculating the number of moles of substances required in the right ratio.

$2Na(s)$	+	$Cl_2(g)$	$\rightarrow$	$2NaCl(s)$	
(2 mol)		(1 mol)		(2 mol)	Initial mole ratio from the stoichiometry
(1 mol)		(0.5 mol)		(1 mol)	All values are divided by 2
⋮		⋮		⋮	Calculated in the right ratio for other values

The number of moles determined can be converted to mass, number of particles or volume of gas using the molar mass, Avogadro constant or molar volume like all the relationships you have learned before.

Example 17

Burning of aluminium in air is as follows:

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

What is the mass of aluminium oxide produced if 5.4 g of aluminium is burnt completely in air? [Relative atomic mass: O = 16, Al = 27]

Solution

Question analysis and solution plan

Equation:	$4Al(s)$	+	$3O_2(g)$	→	$2Al_2O_3(s)$
Information from the equation:	(4 mol)				(2 mol)
Information from the question:	(5.4 g)				(? g – question to be answered)

Step 1 (How many moles?) → Step 2 (Mole ratio?) → Step 3 (What is the mass?)

Number of moles in 5.4 g of aluminium, Al $= \dfrac{\text{Mass}}{\text{Molar mass}}$

$= \dfrac{5.4\text{ g}}{27\text{ g mol}^{-1}}$

$= 0.2$ mol

Step 1: Mass of Al → Number of moles of Al

Based on the equation, 4 mol of aluminium, Al produces 2 mol of aluminium oxide, Al_2O_3. Therefore, 0.2 mol of aluminium, Al produces 0.1 mol of aluminium oxide, Al_2O_3.

Step 2: Calculate the mole ratio of Al_2O_3.

Hence, the mass of aluminium oxide, Al_2O_3 produced

= Number of moles × Molar mass

$= 0.1\text{ mol} \times [2(27) + 3(16)]\text{ g mol}^{-1}$

$= 0.1\text{ mol} \times 102\text{ g mol}^{-1}$

$= 10.2$ g

Step 3: Number of moles of Al_2O_3 → Mass of Al_2O_3

Activity 3.20

Solving numerical stoichiometry problems

CT

[Relative atomic mass: H = 1, C = 12, O = 16, Cl = 35.5, Ca = 40, Fe = 56, Zn = 65; Avogadro constant, N_A: 6.02×10^{23} mol^{-1}; Molar volume = 22.4 dm^3 mol^{-1} at STP or 24 dm^3 mol^{-1} at room conditions]

1. Decomposition of calcium carbonate by heating is as follows:

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

What is the mass of calcium carbonate required to produce 1.2 dm^3 of carbon dioxide gas, CO_2 at room conditions?

2. Zinc reacts with hydrochloric acid as follows:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

What is the mass of zinc that should be used to produce 0.5 mol of hydrogen gas, H_2?

3. A sample of iron(III) oxide, Fe_2O_3 is heated in a stream of excess hydrogen gas, H_2 to produce 5.6 g of iron metal according to the following equation:

$$Fe_2O_3(g) + 3H_2(g) \rightarrow 2Fe(s) + 3H_2O(l)$$

Calculate the mass of the iron(III) oxide sample.

4. Nitrogen and hydrogen gases react according to the following equation:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

How many molecules of ammonia, NH_3 are produced if 6.72 dm^3 of nitrogen gas at STP reacts completely with hydrogen gas?

Activity 3.21

Creating a computer worksheet

CT

Decomposition of potassium chlorate(V), $KClO_3$ by heat is often used to produce oxygen gas in the laboratory.

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

Assume you are a laboratory assistant. You are required to prepare different amounts of oxygen gas from time to time. Repeated calculations using chemical equations can be simplified using a computer worksheet. Use Microsoft Excel or other suitable programmes to prepare a computer worksheet involving the equation above to solve the following problems:
[Relative atomic mass: O = 16, Cl = 35.5, K = 39; Molar volume = 24 dm^3 mol^{-1} at room conditions]

1. What are the masses of potassium chlorate(V), $KClO_3$ needed to produce 1 dm^3, 5 dm^3, 10 dm^3, 20 dm^3 and 50 dm^3 of oxygen gas?

2. What are the volumes of oxygen gas produced if 0.25 kg, 0.5 kg, 1 kg, 1.5 kg and 2 kg of potassium chlorate(V), $KClO_3$ are used?

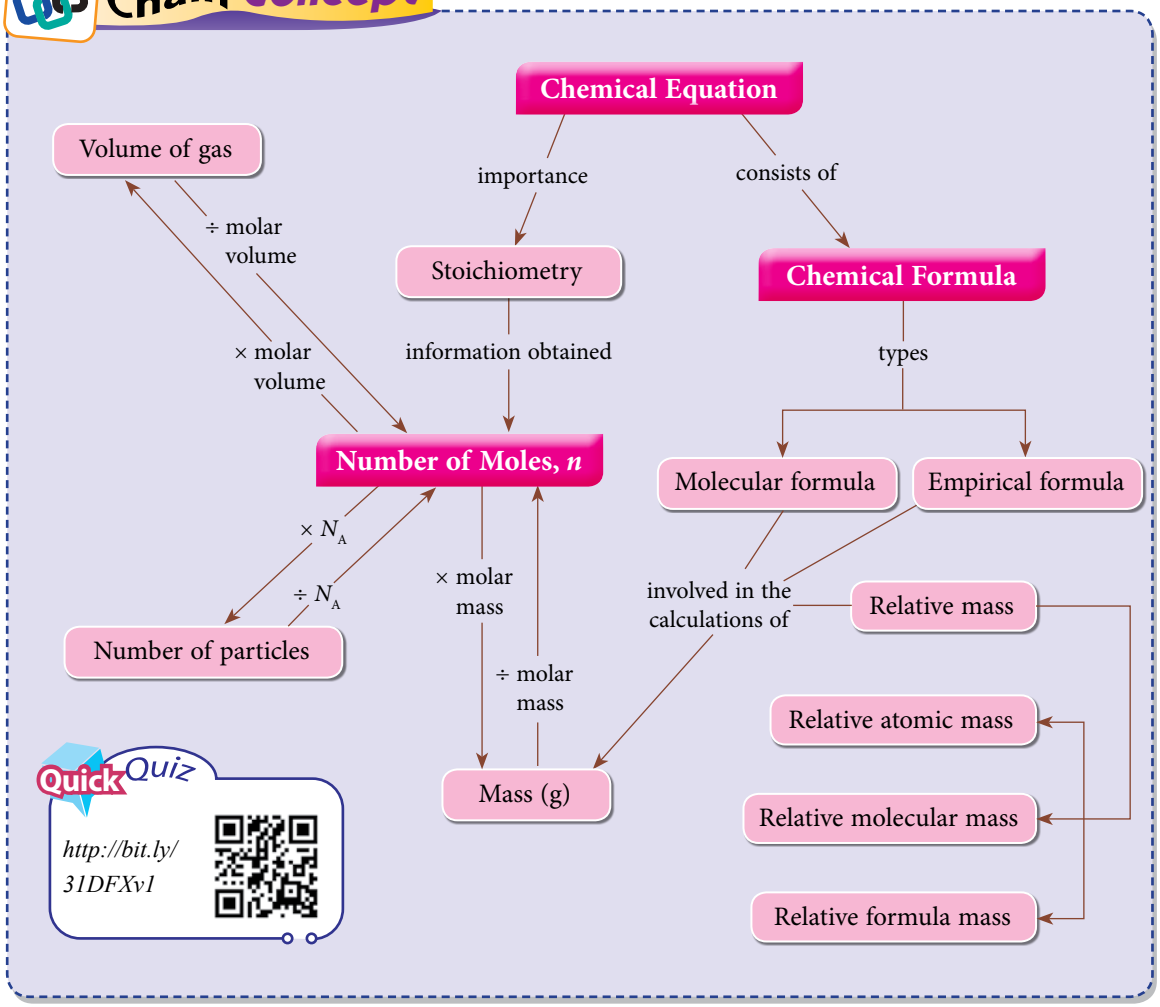
Test Yourself 3.4

1. Write the chemical equations for the following reactions:
   (a) Copper + Silver nitrate solution → Copper(II) nitrate solution + Silver
   (b) Hot zinc metal will react with chlorine gas to produce solid zinc chloride
2. Decomposition of hydrogen peroxide, H_2O_2 occurs according to the following equation:

$$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$$

   (a) What are the products of the decomposition of hydrogen peroxide, H_2O_2?
   (b) Calculate the volume of oxygen produced at STP from the decomposition of 30.6 g of hydrogen peroxide, H_2O_2.

Chain Concept



Quick Quiz

http://bit.ly/31DFXv1

SELF Reflection

1. What is interesting about **The Mole Concept, Chemical Formula and Equation**?
2. Why is the learning of **The Mole Concept, Chemical Formula and Equation** important in the next chemistry lesson?
3. Rate your performance in **The Mole Concept, Chemical Formula and Equation** on a scale of 1 to 10; 1 being the lowest and 10 the highest. Why would you rate yourself at that level?
4. What can you do to improve your mastery in **The Mole Concept, Chemical Formula and Equation**?
5. What else would you like to know about **The Mole Concept, Chemical Formula and Equation**?

http://bit.ly/2MiTOlY

Achievement Test 3

Refer to the Data Table of Elements on page 276.
[Avogadro constant, N_A: 6.02×10^{23} mol^{-1}; Molar volume = 22.4 dm^3 mol^{-1} at STP or 24 dm^3 mol^{-1} at room conditions]

1. What is meant by molar mass and molar volume?

2. What is the relationship between Avogadro constant, number of particles and number of moles?

3. Relative atomic mass of nitrogen is 14

   State the meaning of the above statement based on the carbon-12 scale.

4. Vitamin C or ascorbic acid is an important antioxidant required for our health. Vitamin C has the molecular formula $C_6H_8O_6$.
   (a) What is the empirical formula of vitamin C?
   (b) What is the relative molecular mass of vitamin C?

5. Antacid functions to relieve gastric problems. Figure 1 shows the label on a bottle of antacid.

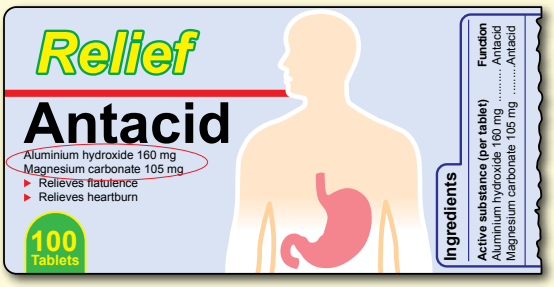


Figure 1

   Give the chemical formulae of the two active ingredients in the antacid.

6. Figure 2 shows the aerobic respiration in our body cells to produce energy from glucose, $C_6H_{12}O_6$. Write a balanced chemical equation for the process of aerobic respiration.

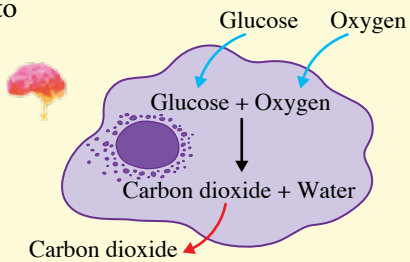


Figure 2

7. Iron(II) sulphate heptahydrate, $FeSO_4.7H_2O$ is often used to treat anaemic patients suffering from the lack of iron mineral.
   (a) What is the molar mass of iron(II) sulphate heptahydrate?
   (b) Calculate the percentage of iron in iron(II) sulphate heptahydrate.

8. Figure 3 shows the weighing steps taken in the determination of the empirical formula of the oxide of metal *Y*.

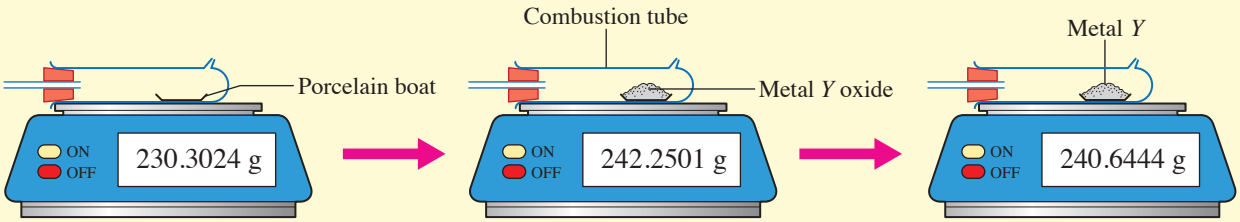


Figure 3

Determine the empirical formula of metal *Y* oxide.
[Relative atomic mass: O = 16, *Y* = 207]

9. *P*, *Q* and *R* are three samples of chemical substances.

P – 0.2 mol of calcium chloride
Q – 12 dm^3 of nitrogen monoxide gas at room conditions
R – 2.408×10^{23} carbon dioxide molecules

Arrange the three samples in ascending order of mass.

10. In your opinion, between the empirical formula and the molecular formula, which formula is more suitable to be used when writing chemical equations? Give your reasons.

Enrichment Corner

1. When steam is passed over a hot iron metal, hydrogen gas and iron(III) oxide are formed. What is the mass of steam required to react completely with 100 g of iron?
[Relative atomic mass: H = 1, O = 16, Fe = 56]

2. Lithium hydroxide, LiOH is used to remove carbon dioxide from the exhaled air in the cabin of a spaceship. [Relative atomic mass: H = 1, Li = 7, C = 12, O = 16]

$$2LiOH + CO_2 \rightarrow Li_2CO_3 + H_2O$$

An outer space mission is carried out for a period of 18 days involving five people on board. If each person is expected to exhale on the average of 42 g of carbon dioxide per hour and each absorption tube can contain 750 g of LiOH, calculate the number of absorption tubes that should be loaded into the spaceship.

Check Answers
https://bit.ly/32OHQGV