3. Laws of Chemical Combination

3.1 Empirical and molecular formulae

While a symbol represents an atom of an element, the molecular formula of a compound represents the smallest unit of the compound that is capable of independent existence.

The empirical formula of a compound is the simplest formula that shows the ratio of the number of atoms of the different elements in the compound.

Experiment 3.1 Determination of the Empirical Formula of Magnesium Oxide

Clean a length of magnesium ribbon with sandpaper. Weigh a crucible and its lid. Cut the magnesium ribbon into tiny bits and put the bits in the crucible. Weigh the crucible, lid and magnesium. Record all these masses.

Place the crucible containing the magnesium on a pipe-clay triangle resting on a tripod stand, Figure 3.1. Heat the crucible, and occasionally lift the crucible lid with a pair of tongs to admit more air. Avoid loss of particles from the crucible while doing this. Once the magnesium starts burning, do not lift the lid anymore.

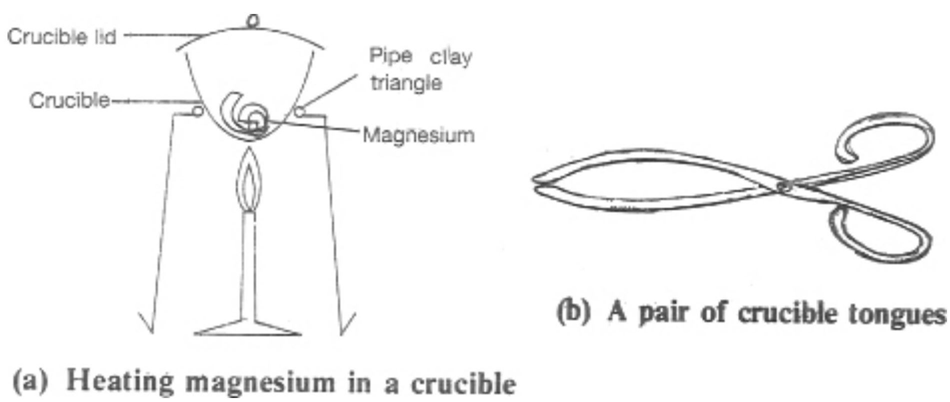

Figure 3.1

After all the magnesium has burnt, allow the crucible to cool, then weigh it with the product of combustion. Heat again, cool and weigh to ensure that burning is completed. How would you tell whether burning was complete or not?

Specimen Results and Calculations.

Mass of crucible and lid = 27.84g

Mass of crucible, lid and Magnesium = 30.34g

$\therefore$ Mass of magnesium = (30.34 â€“ 27.84) = 2.5g

Mass of crucible, lid and product of burning, magnesium oxide. = 31.95g

$\therefore$ Mass of magnesiun oxide = 31.95 â€“ 27.84 = 4.11g

$\therefore$ Mass of oxygen combining with all the magnesium during burning = 4.11 â€“ 2.5 = 1.61g

		Magnesium		Oxygen
Mass of element		2.5g		1.61g
Atomic mass		24g		16g
Number of mole		$\frac{2.5}{24}$		$\frac{1.61}{16}$
Mole ratio		$\frac{2.5}{24}$	:	$\frac{1.61}{16}$
	=	0.104	:	0.1007
Divide throughout by the smaller figure,				
	=	$\frac{0.104}{0.1007}$	:	$\frac{0.1007}{0.1007}$
	=	1.03	:	1.0
Approximating to a whole number,				
	=	1	:	1

The ratio of magnesium to oxygen is therefore 1:1. The empirical formula of magnesium oxide is therefore, MgO.

The empirical formula shows the ratio of the number of moles of each atom of the elements in combination. It may not indicate the actual number of atoms of each element in a molecule of the compound. It is thus not always the same as the molecular formula of the compound. For example, as the empirical formula of magnesium oxide is MgO the molecular formula may be MgO, Mg_2O_2, Mg_3O_3, etc. The ratio of the number of moles of the combining elements remains 1:1.

To find the molecular formula of the compound we must know the relative molecular mass of the compound in addition to the empirical formula. We can then calculate the actual number of atoms per molecule, to obtain the molecular formula.

Given the percentage composition of a compound or the masses of the atoms in a given mass of the compound the empirical formula can be calculated.

Worked Example.

A compound contains 28.8% of magnesium, 14.20% of carbon and 57.0% of oxygen. What is its empirical formula?

Solution.

	Mg		C		O
Relative atomic mass	24.3		12.		16.
Mass of atoms in 100g. of compound	28.8g.		14.2g.		57.0g.
Numbers of mole	$\frac{28.8}{24.3}$		$\frac{14.2}{12}$		$\frac{57.0}{16}$
Mole ratio =	1.185	:	1.183	:	3.563
Dividing with the lowest number	$\frac{1.185}{1.183}$		$\frac{1.183}{1.183}$		$\frac{3.563}{1.183}$
=	1	:	1	:	3

∴ Empirical Formula is MgCO .

Exercise 3 A

A compound containing carbon and hydrogen only contains 80% of carbon. Calculate its empirical formula.

(b) If 0.9g. of aluminium combined with 3.55g. of chlorine, calculate the empirical formula of the aluminium chloride formed. (A1 = 27, C1 = 35.5).

3.2 The Law of conservation of mass

The law of conservation of mass is also called the law of indestructibility of matter. The law states that **in any chemical reaction the total mass of reactants is always equal to the total mass of products.**

Experiment 3.2 Investigating the Law of Conservation of Mass

Pour about 5 cm^3 of dilute tetraoxosulphate(VI) acid into a reagent bottle that has a teat pipette. Take up some barium chloride solution into the pipette. Replace the cap of the reagent bottle carrying the teat pipette (Figure 3.2), and weigh the whole set up. Press the rubber teat to allow the barium chloride solution to mix with the tetraoxosulphate(VI) acid. Describe what is observed. Weigh the whole set up again. Compare the masses of the set-up before and after the reaction.

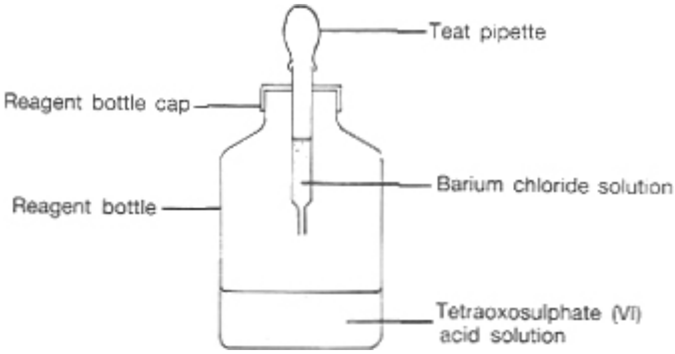

Figure 3.2 Law of conservation of mass

Repeat the experiment with silver trioxonitrate(V) solution in the teat pipette and dilute hydrochloric acid in the reagent bottle.

You will find that when the clear solutions of the tetraoxosulphate(VI) acid and barium chloride are mixed, a white **precipitate** is formed. A reaction has taken place. But the total mass of reagents and bottle before and after the reaction are the same. A similar result is obtained when the experiment is repeated with silver trioxonitrate(V) solution and dilute hydrochloric acid. This shows that mass is conserved in the course of a chemical reaction.

Dalton explained this and the other laws of chemical combination when he put up his atomic theory. A chemical reaction is the rearrangement of the particles (atoms) of matter. It involves **the breaking and making of bonds between these particles**. Since these atoms are neither created nor destroyed, mass must be conserved in the course of a chemical reaction.

Burning of wood is a chemical reaction in which it seems that mass is not conserved. However, if we collect all the smoke, carbon(IV) oxide and water vapour produced during the burning as well as the ashes left behind, and weigh these, their total mass will be equal to the mass of the wood before burning, plus the mass of oxygen used up for the burning. For ordinary chemical reactions the law holds because very little amount of energy is involved. In nuclear reactions where a large amount of energy is involved, mass is not conserved.

3.3 The law of (definite proportion) constant composition

The constant composition of pure compounds was discovered quite early and expressed as a law known as the law of constant composition. The law of constant composition states that:

All pure samples of the same compound contain the same elements combined in a fixed proportion by mass.

Experiment 3.3 Investigating the Law of Constant Composition

Weigh a crucible and its lid accurately. Place one spatula full of copper turnings in the crucible and weigh again. Add four spatula heaps of powdered sulphur. Mix them thoroughly with the spatula and weigh. Record the masses. Support the covered crucible on a pipe-clay triangle (Figure 3.3). Heat it with a low bunsen flame for some time. Blue flames of excess sulphur burning off are seen all around the crucible lid. Continue the low heating until the blue flames are no longer noticeable. Heat more strongly until the bottom of the crucible is red-hot. Burn off the last traces of excess sulphur by spreading the flame over the top of the crucible.

Stop the heating and allow the crucible to cool down. Weigh it with the lid and contents. Heat again, cool and weigh to make sure that all excess sulphur has been burnt off. This is when two successive weighings give constant results.

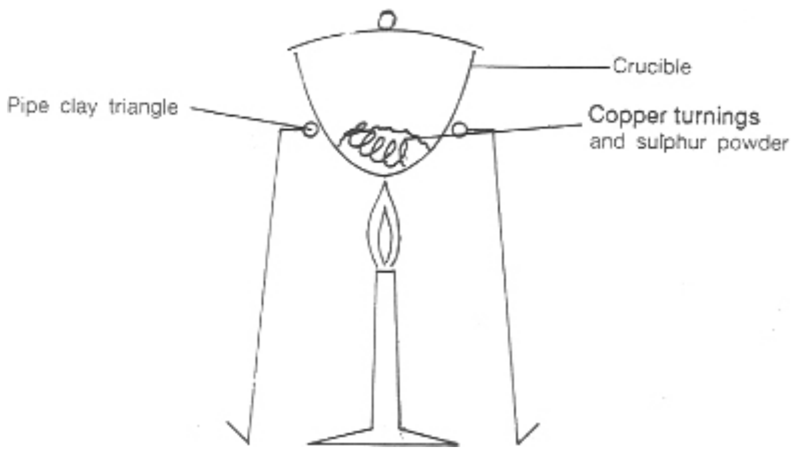

Figure 3.3 Law of definite proportion

Specimen results and calculations

Mass of crucible + lid = 27.84g

Mass of crucible, lid and copper = 30.35g

∴Mass of copper = (30.35 − 27.84)g = 2.51g

Mass of crucible, lid and contents after heating = 31.56g

∴Mass of contents [copper(II) sulphide] = (31.56 − 27.84)g = 3.72g

∴Mass of sulphur in copper(II) sulphide = (3.72 − 2.51)g = 1.21g

∴ % of sulphur in copper(II) sulphide = $\frac{1.21}{3.72} \times 100$ = 32.7%

% of copper = $\frac{2.51}{3.72} \times 100$ = 67.3%

Compare your result with those of other members of your class because your result alone is not enough to make a generalization; or compare it with the following results obtained by groups of students in a particular class.

TABLE 3.1

Groups	Mass of Copper (g)	Mass of Sulphur (g)	Mass of Copper (II) Sulphide (g)
1.	1.78	0.93	2.71
2.	1.12	0.55	1.67
3.	0.83	0.43	1.26
4.	1.21	0.58	1.79
5.	1.78	0.87	2.65
6.	0.62	0.31	0.93
7.	1.48	0.79	2.27
8.	1.83	0.84	2.67
9.	1.44	0.75	2.19
10.	1.51	0.76	2.27
11.	0.84	0.40	1.24
12.	1.48	0.73	2.21
13.	1.58	0.84	2.42
14.	1.17	0.57	1.74

Calculate the percentage of copper (or of sulphur) in copper(II) sulphide for each experiment. You will observe that the percentage of copper (or of sulphur) is a constant, within the limits of experimental error.

An alternative method of investigating the law of constant composition is to reduce known masses of two or more samples of copper(II) oxide prepared by different methods, and then weigh the products of reduction. Copper(II) oxide can be prepared by the following different methods:

1. heating copper in air;
2. thermal decomposition of copper(II) trioxocarbonate(IV); and
3. thermal decomposition of copper(II) trioxonitrate(V).

Experiment 3.4 investigating the Law of Constant Composition (Second Method)

Weigh two samples of copper(II) oxide which were prepared by different methods, into separate porcelain boats, labelled A & B. Put the porcelain boats into a long combustion tube. Pass dry hydrogen into the heated combustion tube and burn off excess hydrogen at the other end of the tube (Figure 3.4).

When reduction appears to be complete, cool and weigh the porcelain boats. Repeat the process of heating and cooling till constant weights are obtained.

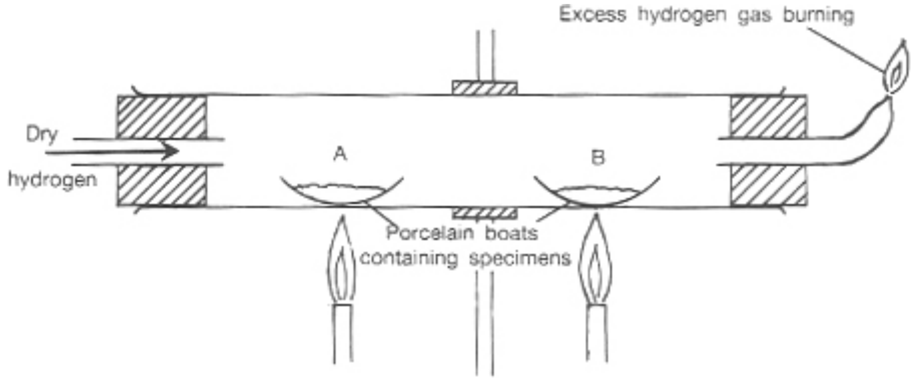

Figure 3.4 Reduction of copper(II) oxide

Specimen results and calculations

	Specimen A	Specimen B
Mass of copper(II) oxide	12.3g	13.8g
Mass of copper (reduction product)	9.8g	11.2g
∴ Mass of oxygen	2.5g	2.6g

% of copper in A = $\frac{9.8}{12.3} \times 100 = 79.7$

% of copper in B = $\frac{11.2}{13.8} \times 100 = 79.8$

The proportions of copper by mass in the two samples are the same, thus upholding the law of constant composition. In general all pure chemical compounds have a constant composition.

Worked Example

Which of the statements of Daltonâ€™s atomic theory explain the law of constant composition.

Solution

Two of Daltonâ€™s statements explain the law of constant composition. These are:

1. Atoms of the same element are alike especially in mass, and
2. Atoms combine in small whole number ratios to form compounds (molecules).

3.4 The law of multiple proportions

This law deals with compounds formed by atoms that have variable combining powers. We shall start by investigating the combination of an atom with variable combining power with another atom.

Experiment 3.5 Investigating the Law of Multiple Proportions

Weigh two porcelain boats. rut two spatula-fulls of copper(II) oxide into boat A and two spatula-fulls of copper(I) oxide into boat B and reweigh the boats. Note the colour of each oxide sample. Put the boats into the combustion tube, pass dry hydrogen through the tube and start heating as in Figure 3.4 to reduce the oxides to metallic copper.

When reduction has taken place, the black copper(II) oxide and the red copper(I) oxide will both become brown, the colour of elemental copper. Cool the boats and reweigh them. Repeat the process of heating, cooling and weighing till the weight of each boat becomes constant. The passing of dry hydrogen should continue throughout the period of heating and cooling to prevent the re-oxidation of the reduced copper.

Specimen results and calculations.

		Boat A		**Boat B**
Mass of oxide		2.5g		4.5g
Mass of residue (copper)		2.0g		4.0g
∴ Mass of oxygen		0.5g		0.5g
Ratio of copper to oxygen		2:0.5		4:0.5
	=	4:1	=	8:1

The ratio of the masses of copper that combine with lg of oxygen in each oxide is 4:8 or 1:2

It is a simple whole number ratio.

The ratio of the masses of oxygen that combine with 4g of copper is 1:0.5 or 2:1, also a simple whole number ratio.

The result of this experiment is summarized in the statement of the law of multiple proportions which states that:

> **When two elements combine to form more than one compound, the different masses of one element which combine with a fixed mass of the second element are in a simple ratio.**

Worked Examples

1. Carbon forms two different compounds with oxygen. In one compound 1.2g of carbon combine with 3.2g of oxygen. In another 1.8g of carbon combine with 2.4g of oxygen. Show that these figures illustrate the law of multiple proportions.

Solution

		1st Compound		**2nd Compound**
Mass of carbon		1.2g		1.8g
Mass of oxygen		3.2g		2.4g
Mass of oxygen in combination with lg of carbon		$\frac{3.2g}{1.2}$		$\frac{2.4g}{1.8}$
	=	2.67g	=	l.33g

Ratio of the masses of oxygen that combine with lg of carbon in the two compounds is 2.67:1.33

i.e. 2:1

This is a simple ratio, so the figures are in agreement with the law of multiple proportions.

2. Hydrogen and oxygen combine to form water. In the presence of electrical discharge they also form hydrogen peroxide. Water contains 11.2% of hydrogen and 88.8% of oxygen while hydrogen peroxide contains 5.8% of hydrogen and 94.2% of oxygen by mass. What law do these figures illustrate?

Solution

In 100g of water, mass of hydrogen = 11.2g.

Mass of oxygen = 88.8g

$\therefore$ 1 g of hydrogen combines with 88.8/11.2 g of oxygen $\simeq$ 8g.

Similarly, in 100g hydrogen peroxide;

Mass of hydrogen = 5.8g

Mass of oxygen = 94.2g

$\therefore$ 1g of hydrogen combines with 94.2g/5.8 oxygen = 16g.

Thus, the ratio of the masses of oxygen mat combine with lg of hydrogen in the two compounds

is 8:16

or 1:2

The figures therefore illustrate the law of multiple proportions.

Chapter Summary

1. The empirical formula represents the simplest ratio of the number of atoms of the elements in a compound, while the molecular formula of a compound represents the smallest unit of it that is capable of independent existence.
2. Daltonâ€™s atomic theory states that matter is made up of small indivisible particles called atoms which are alike in all respects; can neither be created nor destroyed; combine in small whole number ratios.
3. Law of conservation of mass states that in a chemical reaction the total mass of reactants is always equal to the total mass of the products.
4. Law of constant composition states that all pure samples of the same compound contain the same elements combined in a fixed proportion by mass.
5. Law of multiple proportions states that when two elements combine to form more than one compound, the different masses of one element that combine with a fixed mass of the second element are in

a simple whole number ratio.

Assessment

1. If an oxide of copper contains 88.8% of copper by mass, what is the empirical formula of the oxide?
2. 6.1g of a metal, A, reacts completely with 22.9g of chlorine to give 29.0g of the metallic chloride. Calculate the empirical formula of the chloride (A = 27).

(WAEC)

3. 2.51g of a compound contains 1.30g of potassium, 0.5g of sulphur, and 0.765g of oxygen. Calculate the empirical formula of the compound.
4. The mass of carbon in 11.0g of a sample of carbon(IV) oxide is 3.00g. The mass of carbon in 4.40g of another sample of carbon(IV) oxide is 1.20g.
   (a) Find the percentage of oxygen by mass in each sample of carbon(IV) oxide.
   (b) Name the chemical law which is illustrated by the analysis in (a) above.

(WAEC)

5. A metal forms two oxides containing 20.0% and 11.1% of oxygen respectively. Show that these figures agree with the law of multiple proportions.

(WAEC)