STRAND 3 REACTIONS

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

Sub-strands

3.1 Chemical Equations and Calculations
3.2 Types of Reactions
3.3 Acids, Bases and Salts

3.1 Chemical Equations and Calculations

Achievement Indicators

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

- ✓ Write the names and formula of common ionic and covalent compounds.
- ✓ Work out the percentage composition from formula mass and molecular mass.
- ✓ Write balanced equations from chemical statements and vice versa.
- ✓ Define and verify the Law of Conservation of Mass by carrying out experiments.
- ✓ Explore and carry out calculations to verify the Law of Definite Composition.

Chemical formula and names of common substances

Chemical Formulas

Chemical formulas are used to express the composition of molecules and ionic substances in terms of chemical symbols.

Example

Chemical Formula	Name
NH_3	Ammonia
H_2O	Water
O_2	Oxygen
$NaCl$	Sodium chloride
$CaCO_3$	Calcium carbonate

Jons Jakob Berzelius invented a system of chemical notation in 1811.The system is based on the “law of definite proportions”, which states that all samples of a given chemical compound have the same elemental composition. In other words, this law means that every chemical compound contains fixed and constant proportions (by weight) of its constituent elements.

Source: http://crescentok.com

Writing Chemical Formula

The simplest chemical formulas are for binary compounds - compounds made up of two elements.

Note: The sum of the charges in a compound must always equal to zero.

Example 1

Write the chemical formula of calcium chloride.

Solution

Step 1: Write the cation (positive ion) first followed by the anion (negative ion).

Ca^{2+} Cl^{-1}

Step 2: Write the charges above each symbol.

+2 -1

Ca Cl

Step 3: Cross the charges.

+2 -1

Ca Cl

Step 4: Put the charges as subscript (do not wright the positive and negative signs)

$CaCl_2$

Note: Subscripts of 1 are NEVER written. They are understood.

This method is commonly called the “Crossover method”

Other Examples

Write the chemical formula of the following:

1. Calcium hydroxide
2. Magnesium chloride
3. Aluminium oxide

Solution

1. $Ca(OH)_2$
2. $MgCl_2$
3. Al_3O_2

Exercise 3.1.1

1. In the table given below, write the symbol of the positive and negative ions respectively (*The symbols of the first two ions are done for you*). Following that, write the correct formula of each compound that would form from the two ions in the appropriate box (*The first one is done for you).*

Name	Anions	chloride	oxide	hydroxide	nitrate	sulphate	carbonate
Cations	**Symbol of ions**	Cl^-					
sodium	Na^+	NaCl					
potassium							
magnesium							
calcium							
aluminium							
ammonium							
lead (II)							
iron (II)							
iron (III)							
zinc							
copper (II)							

2. Name the following substances:

   i) $AgNO_3$
   ii) $PbSO_4$
   iii) $Al(OH)_3$

3. Write the formula of the following substances:

   i) Copper sulphate
   ii) Zinc nitrate
   iii) Calcium carbonate

4. The formula : $(NH_4)_2SO_4$ has

   A. One Nitrogen atom
   B. Four hydrogen atoms
   C. One sulphur atom
   D. Eight oxygen atoms

Importance of learning the correct chemical formula

Molecular Mass (Mr)

- Used to express the mass of molecules. The mass is obtained by adding the relative atomic masses of the atoms in the molecule.
- There is no unit for molecular mass; however atomic mass unit (a.m.u) can be used.

Example

Calculate the molecular mass of CO_2.

Solution:

C: 1 x 12 = 12
O: 2 x 16 = 32
Molecular mass = 44 a.m.u

Formula Mass

- Used to express the mass of **ionic** compounds. The mass calculation is similar to Mr.
- There is no unit for formula mass; however atomic mass unit (a.m.u) can be used.

Example

Calculate the formula mass of $Ca(NO_3)_2$
Ca: 1 x 40 = 40
N: 2 x 14 = 28
O: 6 x 16 = 96

Formula Mass = 164 a.m.u

Exercise 3.1.3

Calculate the molecular/formula mass of the given substances.

a. $AgNO_3$ b. $PbSO_4$ c. N_2O_3 d. NH_3 e. PBr_3 f. B_2F_6

Percentage Composition

Percent composition gives the percentage of each element by mass present in a compound. This information is important because chemical compound is always the combination of two or more elements. Therefore, when studying a chemical compound, determining percent composition of a certain element within that chemical compound becomes necessary.

$$\text{Percent Composition} = \frac{\text{Mass of sample}}{\text{Total mass of compound}} \times 100$$

Steps in determining the percent composition of the elements in a compound.

1. Determine the formula mass
2. Find the fraction mass of each element in the compound.
3. Convert to percentage.

Example 1

Find the percent composition of red M&M's in a bag of 200 M&M's when there are 26 red M&M's altogether.

Solution

$$\frac{26\text{ red}}{200\text{ total}} \times 100 = 13\%$$

Example 2

Find the percent composition of copper and bromine in $CuBr_2$.

Solution

i. % of Copper (Cu)

$$= \frac{\text{Mass of Cu in the } CuBr_2}{\text{Total Mass of } CuBr_2} \times 100$$

$$= \frac{64}{224} \times 100$$

$$= 28.6\%$$

ii. % of Bromine (Br)

$$= \frac{\text{Mass of Br in the } CuBr_2}{\text{Total Mass of } CuBr_2} \times 100$$

$$= \frac{160}{224} \times 100$$

$$= 71.4\%$$

Hint: The % composition of each element in a substance always adds up to 100%.

Exercise 3.1.3

Calculate the % composition of each element in the following:

a. $NaOH$
b. $KMnO_4$
c. $Al_2(SO_4)_3$

Chemical Equations

- Chemical reactions convert reactants to products, whose properties differ from those of the reactants.
- They have the general form: Reactant → Product

Equations are always written in the same format.

- The left side of the equation lists all the reactants.
- The right side of the equation lists all the products.
- The "+" means "reacts with".
- The arrow means "to produce".

1. Word Equations

Chemists use word equations to describe what has happened in a chemical reaction in a sentence/literal form.

Example

Magnesium + Oxygen → Magnesium oxide

The above word equation can be read as follows: magnesium and oxygen are reactants and magnesium oxide is the product or magnesium and oxygen react to form magnesium oxide.

2. Chemical Equations

Chemical reactions are represented in a concise way by **chemical equations**.

Example 1

For the word equation in (1), the chemical equation is: $2Mg + O_2 \rightarrow 2MgO$

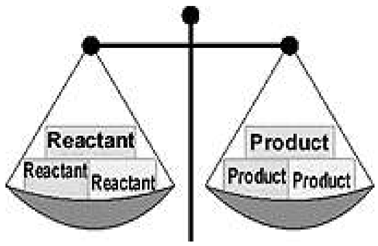
Example 2

$\mathbf{2}H_{\mathbf{2}(g)} + O_{\mathbf{2}(g)} \rightarrow \mathbf{2}H_{\mathbf{2}}O_{(l)}$

- Numbers in front of the formulas are **coefficients**. For example, the number 2 in front of H_2 and H_2O. Coefficients indicate the relative number of molecules or ions of each kind involved in the reaction. Coefficients of 1 are never written - they are understood.
- Numbers to the lower right of chemical symbols in a formula are **subscripts.** For example the number 2 in H_2 and O_2. Subscripts indicate the specific number of atoms of the element found in the substance. Subscripts of 1 are never written - they are understood.
- The **physical state** of each substance in a reaction may be shown in an equation by placing the following symbols to the right of the formula.
  - (*g*) for gas
  - (*l*) for liquid
  - (*s*) for solid
  - (*aq*) for aqueous (water) solution

A chemical equation must have the same number of atoms of each element on both sides of the arrow. When this condition is met, the equation is said to be balanced.

Balancing a Chemical Equation



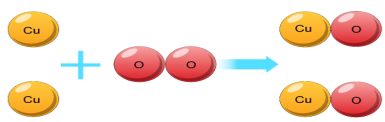
Refer to the word equation:

$\underbrace{\text{Copper + Oxygen}}_{\text{Reactants}} \rightarrow \underbrace{\text{Copper (II) oxide}}_{\text{Product}}$

In the example above, if we replace the words shown above with the correct chemical formulae, we will get an unbalanced equation, as shown below:

$Cu_{(s)} + O_{2(g)} \rightarrow CuO_{(s)}$

To make the equation balanced, we need to adjust the number of units of some of the substances until we get equal numbers of each type of atom on both sides of the arrow.



Here is the balanced chemical equation:

$2Cu_{(s)} + O_{2(g)} \rightarrow 2CuO_{(s)}$

You can now see that there are two copper atoms and two oxygen atoms on each side.

Example

Write the balanced chemical equation for the reaction between hydrogen and oxygen to produce water.

Step 1

Identify the reactants and products and write the formulae for each substance in an equation form.

$H_2 + O_2 \rightarrow H_2O$

Step 2

Check for an unbalanced element. Adjust the number of each atom or molecule needed, but never change the formula. In this case it is O is unbalanced. We need two O atoms on each side.

$H_2 + O_2 \rightarrow 2H_2O$

Step 3

Check for another unbalanced element. In this example, there are 2 H atoms on the left and 4 H atoms (2 × 2) on the right. So we need to double the number of hydrogen molecules on the left.

$2H_2 + O_2 \rightarrow 2H_2O$

Source: www.clipartsheep.com

Hint: Never change subscripts to balance chemical equations

Here are some other examples of balanced equations. Check that you understand why they are balanced.

- $Mg + Cl_2 \rightarrow MgCl_2$
- $2Na + Cl_2 \rightarrow 2NaCl$
- $4Fe + 3O_2 \rightarrow 2Fe_2O_3$
- $4Na + O_2 \rightarrow 2Na_2O$
- $2Na + 2H_2O \rightarrow 2NaOH + H_2$

Exercise 3.14

1. Write chemical equations for the following word equations:

   i. Ammonia gas plus oxygen gives nitrogen monoxide gas plus water vapour.

   ii. Calcium hydroxide solution and carbon dioxide produces solid calcium carbonate and liquid water.

   iii. Magnesium carbonate solution plus aqueous hydrochloric acid gives magnesium chloride plus liquid water and carbon dioxide gas.

2. Balance the following equations:

   i. $N_2 + H_2 \rightarrow NH_3$

   ii. $N_2 + O_2 \rightarrow N_2O$

   iii. $C_2H_5OH + O_2 \rightarrow CO_2 + H_2O$

   iv. $CaC_2 + H_2O \rightarrow C_2H_2 + Ca(OH)_2$

   v. $Fe(OH)_3 \rightarrow Fe_2O_3 + H_2O$

Ionic Equation

The ionic equation is used to describe the chemical reaction while also clearly indicating which of the reactants and/or products exist primarily as ions in aqueous solution.

Steps of writing a net ionic equation:

1. Start with a *balanced* molecular equation.
2. Break all soluble strong electrolytes (compounds with (aq) beside them) into their ions.
    - Indicate the correct formula and charge of each ion.
    - Indicate the correct number of each ion.
    - Write (aq) after each ion.
3. Bring down all compounds with (s), (*l*) or (g) unchanged.

Example

$$2Na_3PO_{4(aq)} \quad + \quad 3CaCl_{2(aq)} \quad \rightarrow \quad 6NaCl_{(aq)} \quad + \quad Ca_3(PO_4)_{2(s)}$$

For the reaction given above, the complete ionic equation is:

$$6\ Na^+_{(aq)} + 2PO_4^{3-}{}_{(aq)} + 3Ca^{2+}{}_{(aq)} + 6Cl^-{}_{(aq)} \quad \rightarrow \quad 6Na^+_{(aq)} + 6Cl^-{}_{(aq)} + Ca_3(PO_4)_{2(s)}$$

The following shows how each ion is achieved; Consider each reactant or product separately:

- 1 mole of Na_3PO_4 contains 3 moles of Na^+ and 1 mole of PO_4^{3-}. Since the balanced equation shows that two moles of Na_3PO_4 are involved in the reaction, a total of 6 moles (2 x 3) of Na^+ and 2 moles (2 x 1) of PO_4^{3-} are formed. Take note that the subscript "4" in the formula for the phosphate ion is not used when determining the number of phosphate ions present. The subscript is part of the formula for the phosphate ion itself.

$$2Na_3PO_{4(aq)} \rightarrow 6Na^+_{(aq)} + 2PO_4^{3-}{}_{(aq)}$$

- 1 mole of $CaCl_2$ contains 1 mole of Ca^{2+} and 2 moles of Cl^-. Remember, the subscript "2" indicates the number of chloride ions. You will never have a diatomic chlorine ion (i.e. Cl_2^- or Cl_2^{2-}) in aqueous solution. Since the balanced equation shows that 3 moles of $CaCl_2$ are involved in the reaction a total of 3 moles (3 x 1) of Ca^{2+} and 6 moles (3 x 2) of Cl^- are formed.

$$3CaCl_{2(aq)} \quad \rightarrow \quad 3Ca^{2+}{}_{(aq)} + 6Cl^-{}_{(aq)}$$

- 1 mole of NaCl contains 1 mole of Na^+ and 1 mole of Cl^- . Since the balanced equation shows that 6 moles of NaCl are produced by the reaction, 6 moles (6 x 1) of Na^+ and 6 moles (6 x 1) of Cl^- will be formed.

$$6NaCl_{(aq)} \rightarrow \ 6Na^+_{(aq)} + 6Cl^-{}_{(aq)}$$

- Since calcium phosphate is an insoluble solid (indicated by the (s) beside its formula), it will not form ions in water. It is brought down unchanged into the complete ionic equation.

Obtaining Net Ionic Equation

In the previous example, the complete ionic equation for the reaction between sodium phosphate and calcium chloride was:

$6Na^+{}_{(aq)} + 2PO_4^{3-}{}_{(aq)} + 3Ca^{2+}{}_{(aq)} + 6Cl^-{}_{(aq)} \rightarrow 6Na^+{}_{(aq)} + 6Cl^-{}_{(aq)} + Ca_3(PO_4)_{2(s)}$

If you look at the ions present on the reactant and product sides of the equation, you will notice that some of the ions are exactly the same.

For example, $6Na^+{}_{(aq)}$ and $6Cl^-{}_{(aq}$ are present on both sides of the equation. That means that the sodium ions and the chloride ions were present in the reaction mixture but *did not participate* in the reaction. The sodium and chloride ions in this reaction are referred to as **spectator ions.**

Spectator ions are ions that are present in the reaction mixture but do not participate in it. They "sit around and watch the reaction take place" just like a spectator at a rugby match watches the players in the game but doesn't play the game himself/herself. You can recognize spectator ions by looking for ions that are present on both sides of the equation. They will always have the same formula, charge and physical state. They will also be present in exactly the same number on both sides of the equation.

To get a net ionic equation:

1. Write the balanced molecular equation.
2. Write the balanced complete ionic equation.
3. Cross out the spectator ions that are present.
4. Write the "leftovers" as the **net ionic equation**.

For the above reaction, the net ionic equation is found by crossing out the spectator ions from the complete ionic equation:

$\cancel{6\,Na^+{}_{(aq)}} + 2PO_4^{3-}{}_{(aq)} + 3Ca^{2+}{}_{(aq)} + \cancel{6Cl^-{}_{(aq)}} \rightarrow \cancel{6Na^+{}_{(aq)}} + \cancel{6Cl^-{}_{(aq)}} + Ca_3(PO_4)_{2(s)}$

Net ionic equation: $2PO_4^{3-}{}_{(aq)} + 3Ca^{2+}{}_{(aq)} \rightarrow Ca_3(PO_4)_{2(s)}$

Exercise 3.1.7

1. Write balanced net ionic equations for the following:

   i. $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

   ii. $Zn_{(s)} + CuSO_{4(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$

   iii. $2NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

   iv. $Na_2CO_{3(aq)} + 2HNO_{3(aq)} \rightarrow 2NaNO_{3(aq)} + H_2O_{(l)} + CO_{2(g)}$

Law of Conservation of Mass

- The law of conservation of mass was established in 1789 by French Chemist Antoine Lavoisier
- This law states that mass is neither created nor destroyed in any ordinary chemical reaction. Therefore the mass of the reactants will always be equal to the mass of the products.

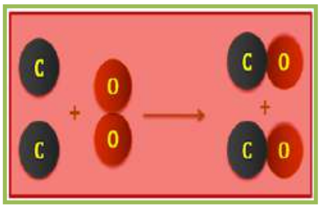
Source: www.catawiki.com

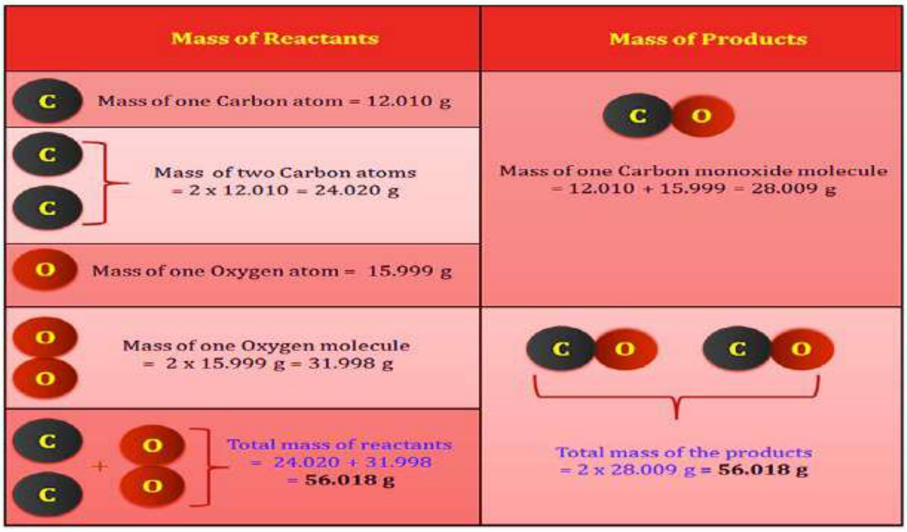
Example 1

Carbon combines with oxygen at low concentration to form carbon monoxide.

$$2C + O_2 \rightarrow 2CO$$

This equation shows that the total mass of the reactants and products are equal. The number of atoms of carbon and oxygen in the reactant and product sides are also equal.



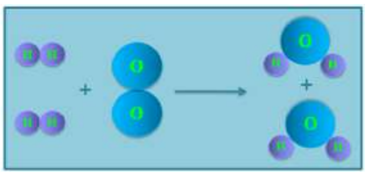


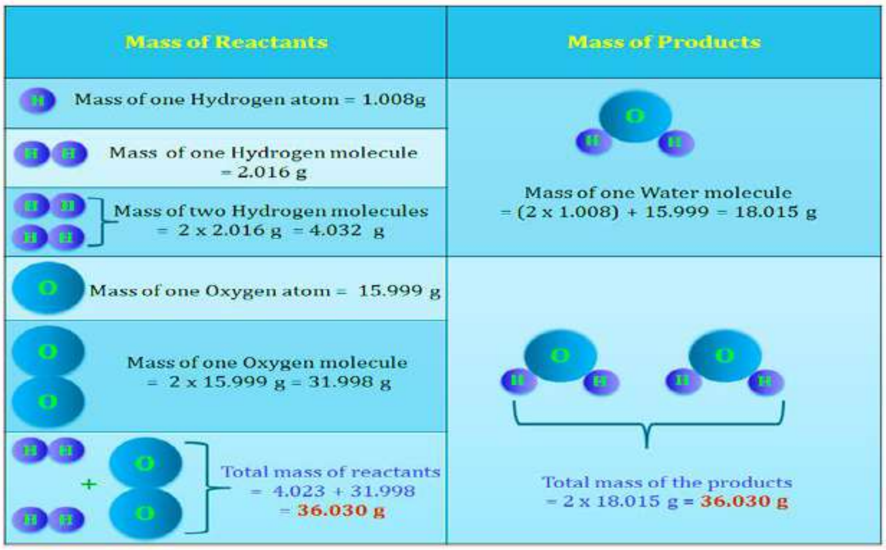
Example 2

Consider the formation of water molecule:

$2H_2 \ + O_2 \ \ \rightarrow \ \ 2H_2O$

This equation also shows that the total mass of the reactants and products are equal. The number of atoms of hydrogen and oxygen in the reactant and product sides are also equal.





Examples showing the Law of Conservation of Mass

1. In an experiment where 50 g of sodium was reacted with 76 g of chlorine to form sodium chloride salt, it was found that 126 g of the salt was formed.

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

50 g   76 g   126 g

2. When 24.8 g of copper carbonate is strongly heated, it produces 16 g of copper oxide and 8.8 g of carbon dioxide gas.

$$CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2\ g)}$$

24.8 g   16 g   8.8 g

3. When a 10 g sample of iron reacts with oxygen to form 18.2 g of ferric oxide, 8.2 g of oxygen was needed.

$$4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$$

10 g   8.2 g   18.2 g

Exercise 3.1.8

1. Consider the reaction of methane with oxygen: $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$

   According to the law of conservation of mass:

   A. The mass of oxygen equals the mass of methane.
   B. The mass of oxygen equals the mass of carbon dioxide.
   C. The combined mass of methane and oxygen will be equal to the combined mass of carbon dioxide and water.
   D. The mass of the reactants remains the same.

2. If 178.8 g of water is separated into hydrogen and oxygen gas and the hydrogen gas has a mass of 20.0 g, what is the mass of the oxygen gas produced?

$$H_2O_{(l)} \rightarrow H_{2(g)} + O_{2(g)}$$

3. From a laboratory process, a student collects 28.0 g of hydrogen and 224.0 g of oxygen. How much water was originally involved in the process?

4. Consider the reaction equation below and answer the question that follows.

$$Fe_2O_{3(s)} + 3CO \rightarrow 2Fe_{(l)} + 3CO_{2(g)}$$

If 160 g of Fe_2O_3 reacted with CO to form 112 g of Fe and 132 g of CO_2, how many grams of CO was used in this reaction?

5.
   i. Write a balanced equation for the reaction of carbon with oxygen to form carbon dioxide.
   ii. If 6 g of carbon reacts with oxygen to make 22 g of carbon dioxide, what mass of oxygen was used in the reaction?

Law of Definite Composition

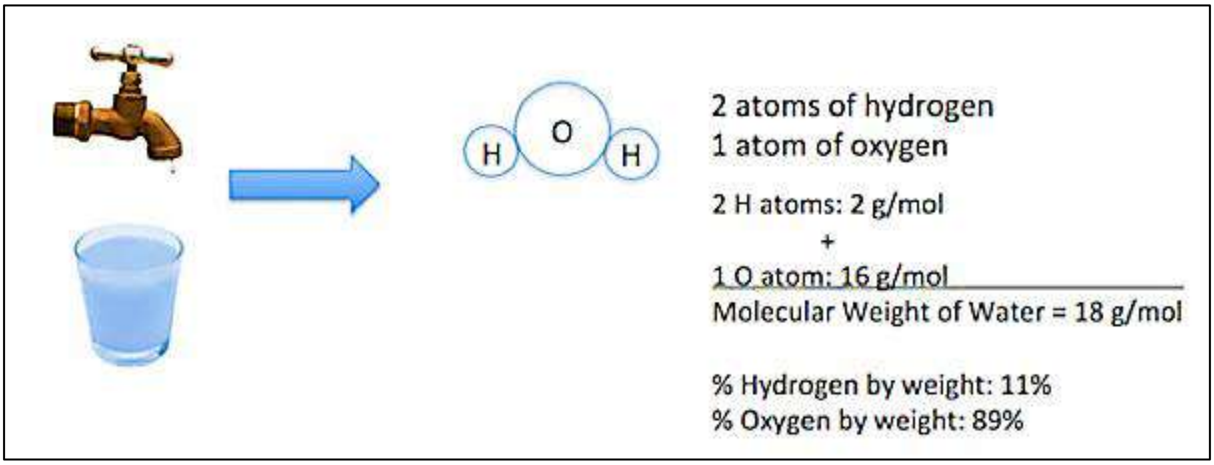
Source: www.daviddarling.info

- The law of definite composition was proposed by Joseph Proust based on his observations on the fixed composition of chemical compounds.
- This law states that chemical compounds are composed of a fixed ratio of elements as determined by mass.
- Elements combine in whole numbers: it is not possible to have a compound with portion an atom.

Example 1

The compound water (H_2O) is always a chemical combination of hydrogen and oxygen in a 1:8 ratio by mass. This means that whatever the source of water, its composition is "two atoms of hydrogen and one atom of oxygen". If a mixture of hydrogen and oxygen in some other ratio, say 1:2, were reacted, there would be water formed, but there would also be some un- reacted oxygen, because water always forms in the 1:8 ratio by mass.

Look at the illustration below:



Source:http://study.com/

Therefore by weight, there are 11% hydrogen and 89% oxygen in one molecule of water. This gives a ratio of 1:8 of hydrogen to oxygen in water.

Example 2

Glucose has the chemical formula $C_6H_{12}O_6$. This means that for glucose to be formed, it must be comprised of 6 atoms of carbon, 12 atoms of hydrogen, and 6 atoms of oxygen. The carbon makes up 40 % of glucose; the hydrogen makes up 6.7% of glucose and the oxygen makes up 53.3% of glucose. This means that if the carbon, hydrogen and oxygen are combined in any other proportion than stated, glucose will not be formed.

Example 3

Ammonia, commonly used in household cleaning agents is made up of the elements hydrogen and nitrogen. It has the chemical formula; NH_3 indicating that there is one atom of nitrogen combined with 3 atoms of hydrogen. This means that ammonia contains 82% nitrogen and 18% hydrogen by mass. Therefore, any other combination of hydrogen and nitrogen would result in an entirely different chemical compound.

Calculations on Law of Definite Proportion

$$\% \text{ of element} = \frac{\text{Mass of the element}}{\text{Total mass of the compound}} \times 100$$

Example 1

What is the experimental percentage of oxygen in CO_2 if 42.0 g of carbon reacted completely with 112.0 g of oxygen?

$$C_{(s)} + O_{2\,(g)} \rightarrow CO_{2\,(g)}$$

$$\% \text{ O} = \frac{\text{Mass of oxygen}}{\text{Mass of carbon dioxide}} \times 100$$

$$= \frac{112 \text{ g}}{112 \text{ g} + 42 \text{ g}} \times 100$$

$$= 72.7\%$$

Example 2

In a 80 gram sample of sodium hydroxide (NaOH), calculate the mass of each element.

Solution

First find the % composition of each element in NaOH (Mr = 40)

Na: $\frac{23}{40} \times 100 = 57.5\%$ **O:** $\frac{16}{40} \times 100 = 40\%$ **H:** $\frac{1}{40} \times 100 = 2.5\%$

Find the mass from the % composition.

i. Na = 57.5%

$= \frac{57.5}{100} \times 80$ g

= <u>46 g</u>

ii. O = 40%

$= \frac{40}{100} \times 80$ g

= <u>32 g</u>

iii. H = 2.5%

$= \frac{2.5}{100} \times 80$ g

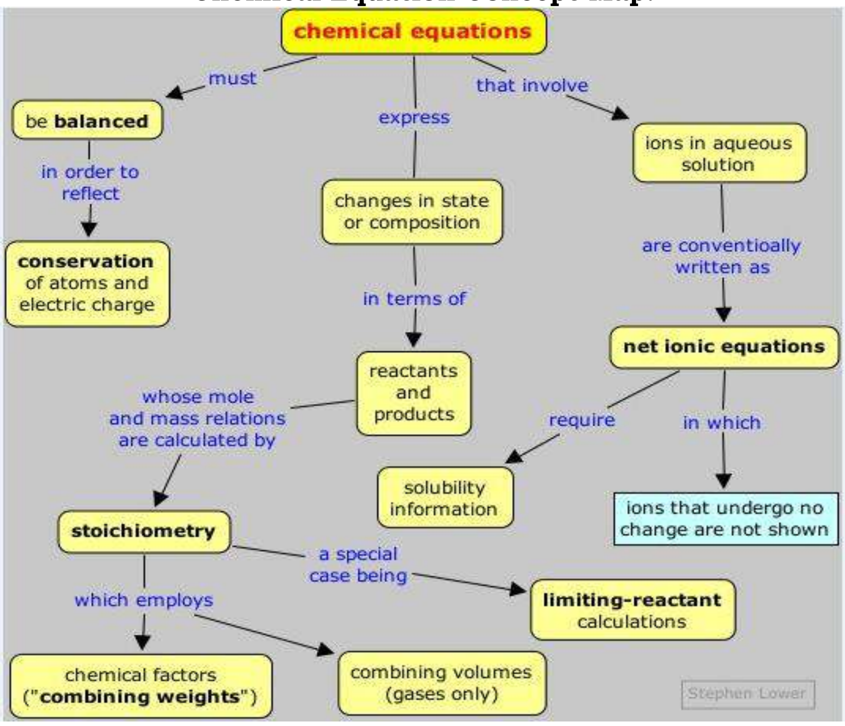
= <u>2 g</u>

Check: The total mass adds up to 80 g. This indicates that the calculations are correctly done.

Exercise 3.1.8

1. Find the mass of each element in 50 g sample of calcium hydroxide ($Ca(OH)_2$).
2. Calculate the mass of copper and bromine in 112 g of $CuBr_2$.

Chemical Equation Concept Map:



Who will win the race below?

Balancing Equations Race

1) ___ C_3H_8 + ___ O_2 → ___ CO_2 + ___ H_2O
2) ___ Al + ___ Fe_3N_2 → ___ AlN + ___ Fe
3) ___ Na + ___ Cl_2 → ___ NaCl
4) ___ H_2O_2 → ___ H_2O + ___ O_2
5) ___ $C_6H_{12}O_6$ + ___ O_2 → ___ H_2O + ___ CO_2
6) ___ H_2O + ___ CO_2 → ___ C_7H_8 + ___ O_2
7) ___ $NaClO_3$ → ___ NaCl + ___ O_2
8) ___ $(NH_4)_3PO_4$ + ___ $Pb(NO_3)_4$ → ___ $Pb_3(PO_4)_4$ + ___ NH_4NO_3
9) ___ BF_3 + ___ Li_2SO_3 → ___ $B_2(SO_3)_3$ + ___ LiF
10) ___ C_7H_{17} + ___ O_2 → ___ CO_2 + ___ H_2O
11) ___ $CaCO_3$ + ___ H_3PO_4 → ___ $Ca_3(PO_4)_2$ + ___ H_2CO_3
12) ___ Ag_2S → ___ Ag + ___ S_8
13) ___ KBr + ___ $Fe(OH)_3$ → ___ KOH + ___ $FeBr_3$
14) ___ KNO_3 + ___ H_2CO_3 → ___ K_2CO_3 + ___ HNO_3
15) ___ $Pb(OH)_4$ + ___ Cu_2O → ___ PbO_2 + ___ CuOH
16) ___ $Cr(NO_2)_2$ + ___ $(NH_4)_2SO_4$ → ___ $CrSO_4$ + ___ NH_4NO_2
17) ___ KOH + ___ $Co_3(PO_4)_2$ → ___ K_3PO_4 + ___ $Co(OH)_2$
18) ___ $Sn(NO_2)_4$ + ___ Pt_3N_4 → ___ Sn_3N_4 + ___ $Pt(NO_2)_4$
19) ___ B_2Br_6 + ___ HNO_3 → ___ $B(NO_3)_3$ + ___ HBr
20) ___ ZnS + ___ AlP → ___ Zn_3P_2 + ___ Al_2S_3