CHAPTER 19

Quantitative Aspects Of Chemical Change

MathU Teaching® Emporium

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1 ATOMIC MASS AND THE MOLE

An equation for a chemical reaction can provide us with a lot of useful information. It tells us what the reactants and the products are in the reaction, and it also tells us the ratio in which the reactants combine to form products. Look at the equation below:

$$\text{Fe} + \text{S} \rightarrow \text{FeS}$$

In this reaction, every atom of iron Fe will react with a single atom of sulphur (S) to form iron sulphide (FeS). However, what the equation does not tell us, is the **quantities** or the **amount** of each substance that is involved. You may for example be given a small sample of iron for the reaction. How will you know how many atoms of iron are in this sample? And how many atoms of sulphur will you need for the reaction to use up all the iron you have? Is there a way of knowing what mass of iron sulphide will be produced at the end of the reaction? These are all very important questions, especially when the reaction is an industrial one, where it is important to know the quantities of reactants that are needed, and the quantity of product that will be formed. This chapter will look at how to quantify the changes that take place in chemical reactions.

1.1 The mole

Sometimes it is important to know exactly how many particles (e.g. atoms or molecules) are in a sample of a substance, or what quantity of a substance is needed for a chemical reaction to take place.

The amount of substance is so important in chemistry that it is given its own name, which is the mole.

DEFINITION

Mole

The mole (abbreviation "mol") is the SI (Standard International) unit for "amount of substance".

The mole is a counting unit just like hours or days. We can easily count one second or one minute or one hour. If we want bigger units of time, we refer to days, months and years. Even longer time periods are centuries and millennia. The mole is even bigger than these numbers. The mole is $602\ 204\ 500\ 000\ 000\ 000\ 000\ 000$ or $6,022 \times 10^{23}$ particles. This is a very big number! We call this number Avogadro's number.

DEFINITION

Avogadro's number

The number of particles in a mole, equal to $6,022 \times 10^{23}$.

If we had this number of cold drink cans, then we could cover the surface of the earth to a depth of over 300 km If you could count atoms at a rate of 10 million per second, then it would take you 2 billion years to count the atoms in one mole!

DID YOU KNOW?

The original hypothesis that was proposed by Amadeo Avogadro was that "equal volumes of gases, at the same temperature and pressure, contain the same number of molecules". His ideas were not accepted by the scientific community and it was only four years after his death, that his original hypothesis was accepted and that it became known as "Avogadro's Law". In honour of his contribution to science, the number of particles in one mole was named *Avogadro's number*.

We use Avogadro's number and the mole in chemistry to help us quantify what happens in chemical reaction. The mole is a very special number. If we measure $12,0$ g of carbon we have one mole or $6,022 \times 10^{23}$ carbon atoms. $63,5$ g of copper is one mole of copper or $6,022 \times 10^{23}$ copper atoms. In fact, if we measure the relative atomic mass of any element on the periodic table, we have one mole of that element.

1.2 Molar mass

DEFINITION

Molar mass

Molar mass (M) is the mass of 1 mole of a chemical substance. The unit for molar mass is **grams per mole** or g·mol^{-1}.

You will remember that when the mass, in grams, of an element is equal to its relative atomic mass, the sample contains one mole of that element. This mass is called the **molar mass** of that element.

NOTE

You may sometimes see the molar mass written as M_m. We will use M in this book, but you should be aware of the alternate notation.

It is worth remembering the following: On the periodic table, the relative atomic mass that is shown can be interpreted in two ways.

1. The mass (in grams) of a *single, average atom* of that element relative to the mass of an atom of carbon.
2. The average atomic mass of all the isotopes of that element. This use is the *relative atomic mass*.
3. The mass of *one mole of the element*. This third use is the molar mass of the element.

Element	Relative atomic mass (u)	Molar mass (g·mol^{-1})	Mass of one mole of the element (g)
Magnesium	$24,3$	$24,3$	$24,3$
Lithium	$6,94$	$6,94$	$6,94$
Oxygen	$16,0$	$16,0$	$16,0$
Nitrogen	$14,0$	$14,0$	$14,0$
Iron	$55,8$	$55,8$	$55,8$

Table 1: The relationship between relative atomic mass, molar mass and the mass of one mole for a number of elements.

WORKED EXAMPLE 1: CALCULATING THE NUMBER OF MOLES FROM MASS

QUESTION

Calculate the number of moles of iron (Fe) in an $11,7$ g sample.

SOLUTION

Step 1: Find the molar mass of iron If we look at the periodic table, we see that the molar mass of iron is $55,8$ g·mol^{-1}. This means that 1 mole of iron will have a mass of $55,8$ g.

Step 2: Find the mass of iron

If 1 mole of iron has a mass of $55,8$ g, then: the number of moles of iron in $111,7$ g must be:

$$\begin{aligned} n &= \frac{111,7 \text{ g}}{55,8 \text{ g·mol}^{-1}} \\ &= \frac{111,7 \text{ g·mol}}{55,8 \text{ g}} \\ &= 2 \text{ mol} \end{aligned}$$

There are 2 moles of iron in the sample.

WORKED EXAMPLE 2: CALCULATING MASS FROM MOLES

QUESTION

You have a sample that contains 5 moles of zinc.

1. What is the mass of the zinc in the sample?
2. How many atoms of zinc are in the sample?

SOLUTION

Step 1: Find the molar mass of zinc

Molar mass of zinc is $65,4$ g·mol^{-1}, meaning that 1 mole of zinc has a mass of $65,4$ g.

WORKED EXAMPLE 2: CALCULATING MASS FROM MOLES (continued)

Step 2: Find the mass

If 1 mole of zinc has a mass of $65,4$ g, then 5 moles of zinc has a mass of: $65,4$ g $\times$ 5 mol $=$ 327 g (answer to 1)

Step 3: Find the number of atoms

5 mol $\times$ $6,022 \times 10^{23}$ atoms·mol^{-1} $= 3,011 \times 10^{23}$ atoms (answer to 2)

1.3 An equation to calculate moles and mass

We can calculate molar mass as follows: molar mass $(M) = \frac{\text{mass (g)}}{\text{mole (mol)}}$

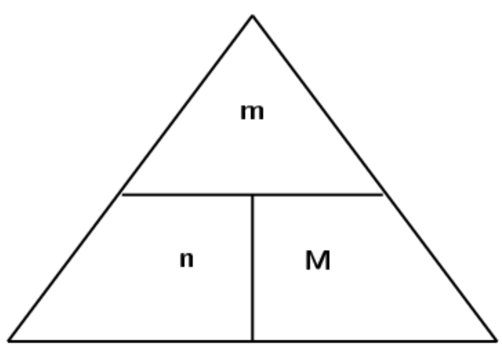
This can be rearranged to give the number of moles:

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

The following diagram may help to remember the relationship between these three variables. You need to imagine that the horizontal line is like a division sign and that the vertical line is like a multiplication sign. So, for example, if you want to calculate M, then the remaining two letters in the triangle are m and n and m is above n with a division sign between them. Your calculation will then be $M = \frac{m}{n}$

TIP

Remember that when you use the equation $n = \frac{m}{M}$, the mass is always in grams (g) and molar mass is in grams per mol (g·mol^{-1}). Always write the units next to any number you use in a formula or sum.



WORKED EXAMPLE 3: CALCULATING MOLES FROM MASS

QUESTION

Calculate the number of moles of copper there are in a sample that with a mass of 127 g.

WORKED EXAMPLE 3: CALCULATING MOLES FROM MASS (continued)

SOLUTION

Step 1: Write down the equation

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

Step 2: Find the moles

$$n = \frac{127 \text{ g}}{63,5 \text{ g·mol}^{-1}}$$
$$= 2 \text{ mol}$$

There are 2 moles of copper in the sample.

WORKED EXAMPLE 4: CALCULATING ATOMS FROM MASS

QUESTION

Calculate the number of atoms there are in a sample of aluminium that with a mass of 81 g

SOLUTION

Step 1: Find the number of moles

$$n = \frac{m}{M}$$
$$= \frac{81 \text{ g}}{27,0 \text{ g·mol}^{-1}}$$
$$= 3 \text{ mol}$$

Step 2: Find the number of atoms

Number of atoms in 3 mol aluminium $3 \times 6,022 \times 10^{23}$

There are $1,8069 \times 10^{24}$ aluminium atoms in a sample of 81 g

1.4 Compounds

So far, we have only discussed moles, mass and molar mass in relation to *elements*. But what happens if we are dealing with a compound? Do the same concepts and rules apply? The answer is *yes*. However, you need to remember that all your calculations will apply to the *whole compound*. So, when you calculate the **molar mass** of a covalent compound, you will need to add the molar mass of each atom in that compound. The number of moles will also apply to the whole molecule. For example, if you have one mole of nitric acid (HNO_3) the molar mass is $63,01$ g·mol^{-1} and there are $6,022 \times 10^{23}$ molecules of nitric acid. For network structures we have to use the **formula mass**. This is the mass of all the atoms in **one formula unit** of the compound. For example, one mole of sodium chloride (NaCl) has a formula mass of $63,01$ g·mol^{-1} and there are $6,022 \times 10^{23}$ molecules of sodium chloride in one formula unit.

In a balanced chemical equation, the number that is written in front of the element or compound, shows the **mole ratio** in which the reactants combine to form a product. If there are no numbers in front of the element symbol, this means the number is '1'.

e.g. $N_2 + 3H_2 \rightarrow 2NH_3$

In this reaction, 1 mole of nitrogen molecules reacts with 3 moles of hydrogen molecules to produce 2 moles of ammonia molecules.

WORKED EXAMPLE 5: CALCULATING MOLAR MASS

QUESTION

Calculate the molar mass of H_2SO_4.

SOLUTION

Step 1: Give the molar mass for each element

Hydrogen = $1,01$ g·mol^{-1}

Sulphur = $32,1$ g·mol^{-1}

Oxygen = $16,0$ g·mol^{-1}

Step 2: Work out the molar mass of the compound

$$M_{\text{H}_2\text{SO}_4} = 2(1,01 \text{ g·mol}^{-1}) + (32,1 \text{ g·mol}^{-1}) + 4(16,0 \text{ g·mol}^{-1}) = 98,12 \text{ g·mol}^{-1}$$

WORKED EXAMPLE 6: CALCULATING MOLES FROM MASS

QUESTION

Calculate the number of moles in 1 kg of $MgCl_2$.

SOLUTION

Step 1: Convert mass into grams

$$m = 1 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 1000 \text{ g}$$

WORKED EXAMPLE 6: CALCULATING MOLES FROM MASS (continued)

Step 2: Calculate the molar mass

$$M_{\text{MgCl}_2} = 24,3\ \text{g·mol}^{-1} + 2(35,45\ \text{g·mol}^{-1}) = 95,2\ \text{g·mol}^{-1}$$

Step 3: Find the number of moles

$$\begin{aligned} n &= \frac{1000\ \text{g}}{95,2\ \text{g·mol}^{-1}} \\ &= 10,5\ \text{mol} \end{aligned}$$

There are $10,5$ moles of magnesium chloride in a 1 kg sample.

DISCUSSION

Understanding moles, molecules and Avogadro's number

Divide into groups of three and spend about 20 minutes answering the following questions together:

1. What are the units of the mole? Hint: Check the definition of the mole.
2. You have a 46 g sample of nitrogen dioxide (NO_2)
   - 2.1 How many **moles** of NO_2 are there in the sample?
   - 2.2 How many moles of nitrogen atoms are there in the sample?
   - 2.3 How many moles of oxygen atoms are there in the sample?
   - 2.4 How many **molecules** of NO_2 are there in the sample?
   - 2.5 What is the difference between a mole and a molecule
3. The exact size of **Avogadro's number** is sometimes difficult to imagine.
   - 3.1 Write down Avogadro's number without using scientific notation.
   - 3.2 How long would it take to count to Avogadro's number? You can assume that you can count two numbers in each second.

2 COMPOSITION

Knowing either the empirical or molecular formula of a compound, can help to determine its composition in more detail. The opposite is also true. Knowing the *composition* of a substance can help you to determine its formula. There are four different types of composition problems that you might come across:

1. Problems where you will be **given the formula** of the substance and asked to **calculate** the **percentage**

by mass of each element in the substance.

2. Problems where you will be **given the percentage composition** and asked to **calculate the formula**.
3. Problems where you will be **given the products** of a chemical reaction and asked to **calculate the formula** of one of the reactants. These are often referred to as combustion analysis problems.
4. Problems where you will be asked to **find** number of moles of **waters of crystallisation**.

The following worked examples will show you how to do each of these.

WORKED EXAMPLE 7: CALCULATING THE PERCENTAGE BY MASS OF ELEMENTS IN A COMPOUND

QUESTION

Calculate the percentage that each element contributes to the overall mass of sulphuric acid (H_2SO_4).

SOLUTION

Step 1: Calculate the molar masses

$$\begin{aligned}\text{Hydrogen} &= 2 \times 1,01 = 2,02\ \text{g·mol}^{-1}\\ \text{Sulfur} &= 32,1\ \text{g·mol}^{-1}\\ \text{Oxygen} &= 4 \times 16,0 = 64,0\ \text{g·mol}^{-1}\end{aligned}$$

Step 2: Use the calculations in the previous step to calculate the molecular mass of sulphuric acid.

$$\text{Mass} = 2,02\ \text{g·mol}^{-1} + 32,1\ \text{g·mol}^{-1} + 64,0\ \text{g·mol}^{-1} = 98,12\ \text{g·mol}^{-1}$$

Step 3: Use the equation

$$\text{Percentage by mass} = \frac{\text{atomic mass}}{\text{molecular mass of } H_2SO_4} \times 100$$

Hydrogen

$$\frac{2,02\ \text{g·mol}^{-1}}{98,12\ \text{g·mol}^{-1}} \times 100\% = 2,06\%$$

Sulfur

$$\frac{32,1\ \text{g·mol}^{-1}}{98,12\ \text{g·mol}^{-1}} \times 100\% = 32,72\%$$

Oxygen

$$\frac{64,0\ \text{g·mol}^{-1}}{98,12\ \text{g·mol}^{-1}} \times 100\% = 65,23\%$$

(You should check at the end that these percentages add up to 100%)

In other words, in one molecule of sulphuric acid, hydrogen makes up $2,06\%$ of the mass of the compound, sulfur makes up $2,06\%$ and oxygen makes up $65,23\%$.

WORKED EXAMPLE 8: DETERMINING THE EMPIRICAL FORMULA OF A COMPOUND

SOLUTION

Step 1: Give the masses

Carbon = $52,2$ g, hydrogen = $13,0$ g and oxygen = $34,8$ g

Step 2: Calculate the number of moles

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

Therefore:

$$n_{\text{carbon}} = \frac{52,2\text{ g}}{12,0\text{ g}\cdot\text{mol}^{-1}} = 4,35\text{ mol}$$

$$n_{\text{hydrogen}} = \frac{13,0\text{ g}}{1,01\text{ g}\cdot\text{mol}^{-1}} = 12,87\text{ mol}$$

$$n_{\text{oxygen}} = \frac{34,8\text{ g}}{16,0\text{ g}\cdot\text{mol}^{-1}} = 2,18\text{ mol}$$

Step 3: Find the smallest number of moles

Use the ratios of molar numbers calculated above to find the empirical formula.

$$\text{units in empirical formula} = \frac{\text{moles of this element}}{\text{smallest number of moles}}$$

In this case, the smallest number of moles is $2,18$. Therefore:

Carbon

$$\frac{4,35}{2,18} = 2$$

Hydrogen

$$\frac{12,87}{2,18} = 6$$

Oxygen

$$\frac{2,18}{2,18} = 1$$

Therefore the empirical formula of this substance is: C_2H_6O.

WORKED EXAMPLE 9: DETERMINING THE FORMULA OF A COMPOUND

QUESTION

207 g of lead combines with oxygen to form 239 g of a lead oxide. Use this information to work out the formula of the lead oxide (Relative atomic masses: Pb $= 202,7$ u and O $= 16,0$ u).

SOLUTION

Step 1: Find the mass of oxygen

$$239 \text{ g} - 207 \text{ g} = 32 \text{ g}$$

Step 2: Find the moles of oxygen

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

Lead

$$n = \frac{207 \text{ g}}{207,2 \text{ g·mol}^{-1}} = 1 \text{ mol}$$

Oxygen

$$n = \frac{32 \text{ g}}{16,0 \text{ g·mol}^{-1}} = 2 \text{ mol}$$

Step 3: Find the mole ratio

The mole ratio of Pb : O in the product is $1 : 2$, which means that for every atom of lead, there will be two atoms of oxygen. The formula of the compound is PbO_2.

WORKED EXAMPLE 10: EMPIRICAL AND MOLECULAR FORMULA

QUESTION

Vinegar, which is used in our homes, is a dilute form of acetic acid. A sample of acetic acid has the following percentage composition: $39,9$ % carbon, $6,7$ % hydrogen and $53,4$ % oxygen.

1. Determine the empirical formula of acetic acid.
2. Determine the molecular formula of acetic acid if the molar mass of acetic acid is $60,06$ g·mol^{-1}.

SOLUTION

Step 1: Find the mass

In 100 g of acetic acid, there is $39,9$ g C, $6,7$ g H and $53,4$ g O.

Step 2: Find the moles

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

WORKED EXAMPLE 10: EMPIRICAL AND MOLECULAR FORMULA (continued)

$$n_{\text{C}} = \frac{39,9 \text{ g}}{12,0 \text{ g·mol}^{-1}}$$
$$= 3,325 \text{ mol}$$

$$n_{\text{H}} = \frac{6,7 \text{ g}}{1,01 \text{ g·mol}^{-1}}$$
$$= 6,634 \text{ mol}$$

$$n_{\text{O}} = \frac{53,4 \text{ g}}{16,0 \text{ g·mol}^{-1}}$$
$$= 3,338 \text{ mol}$$

Step 3: Find the empirical formula

C	H	O
3,325	6,634	3,338
1	2	1

Empirical formula is CH_2O

Step 4: Find the molecular formula

The molar mass of acetic acid using the empirical formula is $30,02$ g·mol^{-1} . However the question gives the molar mass as $60,06$ g·mol^{-1}. Therefore the actual number of moles of each element must be double what it is in the empirical formula ($\frac{60,06}{30,02} = 2$). The molecular formula is therefore $C_2H_4O_2$ or CH_3COOH

WORKED EXAMPLE 11: WATERS OF CRYSTALLISATION

QUESTION

Aluminium trichloride ($AlCl_3$) is an ionic substance that forms crystals in the solid phase. Water molecules may be trapped inside the crystal lattice. We represent this as: $AlCl_3.nH_2O$. Carine heated some aluminium trichloride crystals until all the water had evaporated and found that the mass after heating was $2,8$ g. The mass before heating was 5 g. What is the number of moles of water molecules in the aluminium trichloride before heating?

SOLUTION

Step 1: Find the number of water molecules

We first need to find n, the number of water molecules that are present in the crystal. To do this we first

WORKED EXAMPLE 11: WATERS OF CRYSTALLISATION (continued)

note that the mass of water lost is 5 g -$2,8$ g = $2,2$ g.

Step 2: Find the mass ratio

The mass ratio is:

$AlCl_3$	H_2O
$2,8$	$2,2$

Step 3: Find the mole ratio

To work out the mole ratio we divide the mass ratio by the molecular mass of each species:

$AlCl_3$	H_2O
$\frac{2,8 \text{ g}}{133,35 \text{ g·mol}^{-1}}$	$\frac{2,2 \text{ g}}{18,02 \text{ g·mol}^{-1}}$
$0,02099...$	$0,12208...$

Next we convert the ratio to whole numbers by dividing both sides by the smaller amount:

$AlCl_3$	H_2O
0,020997375	0,12208657
$\frac{0,021}{0,021}$	$\frac{0,122}{0,021}$
1	6

The mole ratio of aluminium trichloride to water is: $1 : 6$

Step 4: Write the final answer

And now we know that there are 6 moles of water molecules in the crystal. The formula is $AlCl_3.6H_2O$.

We can perform experiments to determine the composition of substances. For example, blue copper sulphate ($CuSO_4$) crystals contain water. On heating the waters of crystallisation evaporate and the blue crystals become white. By weighing the starting and ending products, we can determine the amount of water that is in copper sulphate. Another example is reducing copper oxide to copper.

3 AMOUNT OF SUBSTANCE

3.1 Molar Volumes of Gases

DEFINITON

Molar volume of gases

One mole of gas occupies $22,4$ dm^3 at standard temperature and pressure.

This applies to any gas that is at standard temperature and pressure. In grade 11 you will learn more about this and the gas laws.

NOTE

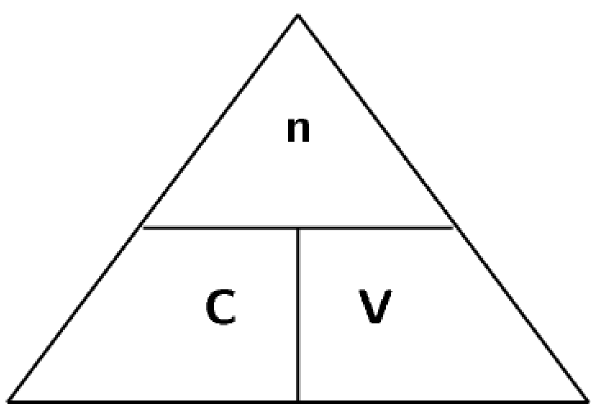
Standard temperature and pressure (S.T.P.) is defined as a temperature of $273,15$ K and a pressure of $0,986$ atm.

For example, 2 mol of H_2 gas will occupy a volume of $44,8$ dm^3 at standard temperature and pressure (S.T.P.). and $67,2$ dm^3 of ammonia gas (NH_3) contains 3 mol of ammonia.

3.2 Molar concentrations of liquids

A typical solution is made by dissolving some solid substance in a liquid. The amount of substance that is dissolved in a given volume of liquid is known as the **concentration** of the liquid. Mathematically, concentration (C) is defined as moles of solute (n) per unit volume (V) of solution.

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



For this equation, the units for volume are dm^3 (which is equal to litres). Therefore, the unit of concentration is mol·dm^{-3}.

DEFINITON

Concentration

Concentration is a measure of the amount of solute that is dissolved in a given volume of liquid. It is measured in mol·dm^{-3}.

WORKED EXAMPLE 12: CONCENTRATION CALCULATIONS 1

QUESTION

If $3,5$ g of sodium hydroxide (NaOH) is dissolved in $2,5$ dm^3 of water, what is the concentration of the solution in mol·dm^{-3}?

SOLUTION

Step 1: Find the number of moles of sodium hydroxide

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{3,5 \text{ g}}{40,01 \text{ g·mol}^{-1}} \\ &= 0,0875 \text{ mol} \end{aligned}$$

Step 2: Calculate the concentration

$$\begin{aligned} C &= \frac{n}{V} \\ &= \frac{0,0875 \text{ mol}}{2,5 \text{ dm}^3} \\ &= 0,035 \text{ mol·dm}^{-3} \end{aligned}$$

The concentration of the solution is $0,035$ mol·dm^{-3}.

WORKED EXAMPLE 13: CONCENTRATION CALCULATIONS 2

QUESTION

You have a 1 dm^3 container in which to prepare a solution of potassium permanganate ($KMnO_4$). What mass of $KMnO_4$ is needed to make a solution with a concentration of $0,2$ mol·dm^{-3}?

SOLUTION

Step 1: Calculate the number of moles

$C = \frac{n}{V}$ therefore:

$$\begin{aligned} n &= C \times V \\ &= 0,2 \text{ mol·dm}^{-3} \times 1 \text{ dm}^3 \\ &= 0,2 \text{ mol} \end{aligned}$$

WORKED EXAMPLE 13: CONCENTRATION CALCULATIONS 2 (continued)

Step 2: Find the mass

$$\begin{aligned} m &= nM \\ &= (0,2 \text{ mol})(158 \text{ g·mol}^{-1}) \\ &= 31,6 \text{ g} \end{aligned}$$

The mass of $KMnO_4$ that is needed is $31,6$ g.

WORKED EXAMPLE 14: CONCENTRATION CALCULATIONS 3

QUESTION

How much sodium chloride (in g) will one need to prepare 500 cm^3 of solution with a concentration of $0,01$ mol·dm^{-3}?

SOLUTION

Step 1: Convert the given volume to the correct units

$$V = 500 \text{ cm}^3 \times \frac{1 \text{ dm}^3}{1000 \text{ cm}^3} = 0,5 \text{ dm}^3$$

Step 2: Find the number of moles

$$\begin{aligned} n &= C \times V \\ &= 0,01 \text{ mol·dm}^{-3} \times 0,5 \text{ dm}^3 \\ &= 0,005 \text{ mol} \end{aligned}$$

Step 3: Find the mass

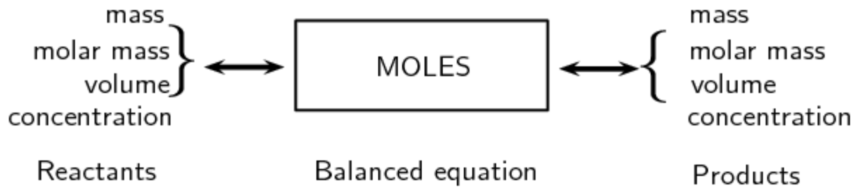
$$\begin{aligned} m &= nM \\ &= (0,005 \text{ mol})(58,45 \text{ g·mol}^{-1}) \\ &= 0,29 \text{ g} \end{aligned}$$

The mass of sodium chloride needed is $0,29$ g

4 STOICHIOMETRIC CALCULATIONS

Stoichiometry is the calculation of the quantities of reactants and products in chemical reactions. It is important to know how much product will be formed in a chemical reaction, or how much of a reactant is needed to make a specific product.

The following diagram shows how the concepts that we have learnt in this chapter relate to each other and to the balanced chemical equation:



WORKED EXAMPLE 15: STOICHIOMETRIC CALCULATION 1

QUESTION

What volume of oxygen at S.T.P. is needed for the complete combustion of 2 dm^3 of propane (C_3H_8)? (Hint: CO_2 and H_2O are the products in this reaction (and in all combustion reactions))

SOLUTION

Step 1: Write the balanced equation

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O\ (g)$$

Step 2: Find the ratio

Because all the reactants are gases, we can use the mole ratios to do a comparison. From the balanced equation, the ratio of oxygen to propane in the reactants is $5 : 1$.

Step 3: Find the answer

One volume of propane needs five volumes of oxygen, therefore 2 dm^3 of propane will need 10 dm^3 of oxygen for the reaction to proceed to completion.

WORKED EXAMPLE 16: STOICHIOMETRIC CALCULATION 2

QUESTION

What mass of iron (II) sulphide is formed when $5,6$ g of iron is completely reacted with sulphur?

SOLUTION

Step 1: Write the balanced equation

$$Fe\ (s) + S\ (s) \rightarrow FeS\ (s)$$

WORKED EXAMPLE 16: STOICHIOMETRIC CALCULATION 2 (continued)

Step 2: Calculate the number of moles

We find the number of moles of the given substance:

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{5,6 \text{ g}}{55,8 \text{ g·mol}^{-1}} \\ &= 0,1 \text{ mol} \end{aligned}$$

Step 3: Find the mole ratio

We find the mole ratio between what was given and what you are looking for. From the equation 1 mol of Fe gives 1 mol of FeS. Therefore, $0,1$ mol of iron in the reactants will give $0,1$ mol of iron sulphide in the product.

Step 4: Find the mass of iron sulphide

$$\begin{aligned} m &= nM \\ &= (0,1 \text{ mol})(87,9 \text{ g·mol}^{-1}) \\ &= 8,79 \text{ g} \end{aligned}$$

The mass of iron (II) sulphide that is produced during this reaction is $8,79$ g.

4.1 Theoretical yield

When we are given a known mass of a reactant and are asked to work out how much product is formed, we are working out the theoretical yield of the reaction. In the laboratory, chemists almost never get this amount of product. In each step of a reaction a small amount of product and reactants is "lost" either because a reactant did not completely react or some other unwanted products are formed. This amount of product that you actually got is called the actual yield. You can calculate the percentage yield with the following equation:

$$\text{\% yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

WORKED EXAMPLE 17: INDUSTRIAL REACTION TO PRODUCE FERTILISER

QUESTION

Sulphuric acid (H_2SO_4) reacts with ammonia (NH_3) to produce the fertiliser ammonium sulphate ($(NH_4)_2SO_4$). What is the theoretical yield of ammonium sulphate that can be obtained from $2,0$ kg of sulphuric acid? It is found that $2,2$ kg of fertiliser is formed. Calculate the % yield.

WORKED EXAMPLE 17: INDUSTRIAL REACTION TO PRODUCE FERTILISER (continued)

SOLUTION

Step 1: Write the balanced equation

$$H_2SO_4(aq) + 2NH_3(g) \rightarrow (NH_4)_2SO_4(aq)$$

Step 2: Calculate the number of moles of the given substance

$$\begin{aligned} n_{H_2SO_4} &= \frac{m}{M} \\ &= \frac{2000 \text{ g}}{98,12 \text{ g·mol}^{-1}} \\ &= 20,38 \text{ mol} \end{aligned}$$

Step 3: Find the mole ratio

From the balanced equation, the mole ratio of H_2SO_4 in the reactants to $(NH_4)_2SO_4$ in the product is $1:1$. Therefore, $20,38$ mol of H_2SO_4 forms $20,38$ mol of $(NH_4)_2SO_4$.

Step 4: Write the answer

The maximum mass of ammonium sulphate that can be produced is calculated as follows:

$$\begin{aligned} m &= nM \\ &= (20,38 \text{ mol})(114,04 \text{ g·mol}^{-1}) \\ &= 2\ 324,14 \text{ g} \end{aligned}$$

The maximum amount of ammonium sulphate that can be produced is $2,324$ kg.

Step 5: Calculate the % yield

$$\begin{aligned} \text{\% yield} &= \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \\ &= \frac{2,2\ kg}{2,324\ kg} \times 100 \\ &= 94,66\% \end{aligned}$$

WORKED EXAMPLE 18: CALCULATING THE MASS OF REACTANTS AND PRODUCTS

QUESTION

Barium chloride and sulphuric acid react according to the following equation to produce barium sulphate and hydrochloric acid.

$$BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$$

If you have 2 g of $BaCl_2$:

WORKED EXAMPLE 18: CALCULATING THE MASS OF REACTANTS AND PRODUCTS (continued)

1. What quantity (in g) of H_2SO_4 will you need for the reaction so that all the barium chloride is used up?
2. What mass of HCl is produced during the reaction?

SOLUTION

Step 1: Find the number of moles of barium chloride

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{2\text{ g}}{208,2\text{ g}\cdot\text{mol}^{-1}} \\ &= 0,0096\text{ mol} \end{aligned}$$

Step 2: Find the number of moles of sulphuric acid

According to the balanced equation, 1 mole of $BaCl_2$ will react with 1 mole of H_2SO_4. Therefore, if $0,0096$ mol of $BaCl_2$ react, then there must be the same number of moles of H_2SO_4 that react because their mole ratio is $1:1$.

Step 3: Find the mass of sulphuric acid

$$\begin{aligned} m &= nM \\ &= (0,0096\text{ mol})(98,12\text{ g}\cdot\text{mol}^{-1}) \\ &= 0,94\text{ g} \end{aligned}$$

(answer to 1)

Step 4: Find the moles of hydrochloric acid

According to the balanced equation, 2 moles of HCl are produced for every 1 mole of the two reactants. Therefore the number of moles of HCl produced is $2(0,0096)$, which equals $0,0192$ mol.

Step 5: Find the mass of hydrochloric acid

$$\begin{aligned} m &= nM \\ &= (0,0192\text{ mol})(36,46\text{ g}\cdot\text{mol}^{-1}) \\ &= 0,7\text{ g} \end{aligned}$$

(answer to 2)

5 CHAPTER SUMMARY

- **The mole (n)** (abbreviation mol) is the SI (Standard International) unit for amount of substance.
- The number of particles in a mole is called **Avogadro's number** and its value is $6,022 \times 10^{23}$. These particles could be atoms, molecules or other particle units, depending on the substance.
- The **molar mass (M)** is the mass of one mole of a substance and is measured in grams per mole or g·mol^{-1}. The numerical value of an element's molar mass is the same as its relative atomic mass. For a covalent compound, the molar mass has the same numerical value as the molecular mass of that compound. For an ionic substance, the molar mass has the same numerical value as the formula mass of the substance.
- The relationship between moles (n), mass in grams (m) and molar mass (M) is defined by the following equation:

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

- In a balanced chemical equation, the number in front of the chemical symbols describes the **mole ratio** of the reactants and products.
- The **empirical formula** of a compound is an expression of the relative number of each type of atom in the compound.
- The **molecular formula** of a compound describes the actual number of atoms of each element in a molecule of the compound.
- The formula of a substance can be used to calculate the **percentage by mass** that each element contributes to the compound.
- The **percentage composition** of a substance can be used to deduce its chemical formula.
- We can use the products of a reaction to determine the formula of one of the reactants.
- We can find the number of moles of waters of crystallisation.
- One mole of gas occupies a volume of $22,4$ dm^3 at S.T.P.. The **concentration** of a solution can be calculated using the following equation,

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

  where C is the concentration (in mol·dm^{-3}), n is the number of moles of solute dissolved in the solution and V is the volume of the solution (in dm^3). The concentration is a measure of the amount of solute that is dissolved in a given volume of liquid.
- The concentration of a solution is measured in mol·dm^{-3}.
- **Stoichiometry** is the calculation of the quantities of reactants and products in chemical reactions. It is also the numerical relationship between reactants and products.

- The theoretical yield of a reaction is the maximum amount of product that we expect to get out of a reaction.

6 EXERCISES

6.1 Exercise 1

1. How many atoms are there in:
   1.1 1 mole of a substance
   1.2 2 moles of calcium
   1.3 5 moles of phosphorus
   1.4 $24,3$ g of magnesium
   1.5 $24,0$ g of carbon
2. Complete the following table:

Element	Relative atomic mass (u)	Sample mass (g)	Number of moles in the sample
Hydrogen	1,01	1,01	
Magnesium	24,3	24,3	
Carbon	12,0	24,0	
Chlorine	35,45	70,9	
Nitrogen	14,0	42,0	

6.2 Exercise 2

1. Give the molar mass of each of the following elements:
   1.1 Hydrogen gas
   1.2 Nitrogen gas
   1.3 Bromine gas
2. Calculate the number of moles in each of the following samples:
   2.1 $21,6$ g of boron(B)
   2.2 $54,9$ g of manganese(Mn)
   2.3 $100,3$ g of Mercury(Hg)
   2.4 50 g of Barium(Ba)
   2.5 40 g of Lead (Pb)

6.3 Exercise 3

1. Calculate the number of moles in each of the following samples:

   1.1 $5,6$ g of calcium $0,02$ g of manganese

   1.2 40 g of aluminium

2. A lead sinker has a mass of 5 g. Answer the following questions.

   2.1 Calculate the number of moles of lead the sinker contain

   2.2 How many lead atoms are in the sinker?

3. Calculate the mass of each of the following samples:

   3.1 $2,5$ mol magnesium

   3.2 12 mol lithium

   3.3 $4,5 \times 10^{25}$ atoms of silicon

6.4 Exercise 4

1. Calculate the molar mass of the following chemical compounds:

   1.1 KOH

   1.2 $FeCl_3$

   1.3 $Mg(OH)_2$

2. How many moles are present in:

   2.1 10 g of Na_2SO_4

   2.2 34 g of $Ca(OH)_2$

   2.3 $2,45 \times 10^{23}$ molecules of CH_4

3. For a sample of $0,2$ moles of magnesium bromide $MgBr_2$, calculate:

   3.1 The number of moles of Mg^{2+} ions

   3.2 The number of moles of Br^- ions

4. You have a sample containing 3 mol of calcium chloride. Answer the following questions.

   4.1 What is the chemical formula of calcium chloride?

   4.2 How many calcium atoms are in the sample?

5. Calculate the mass of:

   5.1 3 moles of NH_4OH

   5.2 $4,2$ moles of Ca$(NO_3)_2$

6.5 Exercise 5

1. Calcium chloride is produced as the product of a chemical reaction.

    1.1 What is the formula of calcium chloride?

    1.2 What is the percentage mass of each of the elements in a molecule of calcium chloride?

    1.3 If the sample contains 5 g of calcium chloride, what is the mass of calcium in the sample?

    1.4 How many moles of calcium chloride are in the sample?

2. 13 g of zinc combines with $6,4$ g of sulphur.

    2.1 What is the empirical formula of zinc sulphide?

    2.2 What mass of zinc sulphide will be produced?

    2.3 What is the percentage mass of each of the elements in zinc sulphide?

    2.4 The molar mass of zinc sulphide is found to be $97,44$ g·mol^{-1}.Determine the molecular formula of zinc sulphide

3. A calcium mineral consisted of $29,4$% calcium, $23,5$% sulphur and $41,7$% oxygen by mass. Calculate the empirical formula of the mineral.

4. A chlorinated hydrocarbon compound was analysed and found to consist of $24,24$% carbon, $4,04$% hydrogen and $71,72$% chlorine. From another experiment the molecular mass was found to be 99 g·mol^{-1}. Deduce the empirical and molecular formula.

5. Magnesium sulphate has the formula $MgSO_4.nH_2O$. A sample containing $5,0$ g of magnesium sulphate was heated until all the water had evaporated. The final mass was found to be $2,6$ g. How many water molecules were in the original sample?

6.6 Exercise 6

1. $5,95$ g of potassium bromide was dissolved in 400 dm^3 of water. Calculate its concentration.

2. 100 g of sodium chloride NaCl is dissolved in 450 cm^3 of water.

    2.1 How many moles of NaCl are present in solution?

    2.2 What is the volume of water dm^3 ?

    2.3 Calculate the concentration of the solution.

3. What is the molarity of the solution formed by dissolving 80 g of sodium hydroxide NaOH in 500 cm^3 of water?

4. What mass (g) of hydrogen chloride HCl is needed to make up 1000 cm^3 of a solution of concentration 1 mol·dm^{-3}?

5. How many moles of H_2SO_4 are there in 250 cm^3 of a $0,8$ $mol{\cdot}dm^{-3}$ sulphuric acid solution? What mass of acid is in this solution?

6.7 Exercise 7

1. Diborane,B_2H_6 , was once considered for use as a rocket fuel. The combustion reaction for diborane is:

$$B_2H_6 + 3O_2 \rightarrow 2HBO_2 + 2H_2O$$

If we react $2,37$ g of diborane, how many grams of water would we expect to produce?

2. Sodium azide is a commonly used compound in airbags. When triggered, it has the following reaction:

$$2Na_2N_3 \rightarrow 2Na + 3N_2$$

If $23,4$ g of sodium azide is used, how many moles of nitrogen gas would we expect to produce? What volume would this nitrogen gas occupy at STP?

3. Photosynthesis is a chemical reaction that is vital to the existence of life on Earth. During photosynthesis, plants and bacteria convert carbon dioxide gas, liquid water, and light into glucose ($C_6H_{12}O_6$) and oxygen gas.

   3.1 Write down the balanced equation for the photosynthesis reaction.

   3.2 If 3 mol of carbon dioxide are used up in the photosynthesis reaction, what mass of glucose will be produced?

7 ANSWERS TO EXERCISES

7.1 Exercise 1

1.1 $6,022 \times 10^{23}$

1.2 $12,044 \times 10^{23}$

1.3 $30,11 \times 10^{23}$

1.4 $6,022 \times 10^{23}$

1.5 $12,044 \times 10^{23}$

2.1 1 mol

2.2 1 mol

2.3 2 mol

2.4 2 mol

2.5 3 mol

7.2 Exercise 2

1.1 $1,01$ g·mol^{-1}

1.2 $14,01$ g·mol^{-1}

1.3 $79,9$ g·mol^{-1}

2.1 2 mol

2.2 1 mol

2.3 $0,5$ mol

2.4 $0,36$ mol

2.5 $0,19$ mol

7.3 Exercise 3

1.1 $0,14$ mol

1.2 $3,64 \times 10^{-4}$ mol

1.3 $1,48$ mol

2.1 $0,024$ mol

2.2 $1,44 \times 10^{22}$ atoms

3.1 $60,78$ g

3.2 $83,28$ g

3.3 $2099,17$ g

7.4 Exercise 4

1.1 $56,1$ g·mol^{-1}

1.2 $162,2$ g·mol^{-1}

1.3 $58,33$ g·mol^{-1}

2.1 $0,07$ mol

2.2 $0,46$ mol

2.3 $4,07$ mol

3.1 $0,24$ moles

3.2 $0,2$ moles

4.1 $CaCl_2$

4.2 $1,81 \times 10^{24}$ atoms

5.1 $105,18$ g

5.2 $689,22$ g

7.5 Exercise 5

1.1 $CaCl_2$

1.2 Ca% = $36,15$% and Cl% = $63,89$%

1.3 $1,81$ g

1.4 $0,05$ mol

2.1 ZnS

2.2 $19,49$ g

2.3 Zn% = $67,1$% and S% = $32,9$%

2.4 ZnS

3. $CaSO_4$

4. CH_2Cl

5. 6

7.6 Exercise 6

1.1 $1,25 \times 10^{-4}$

2.1 $1,71$ mol

2.2 $0,45$ dm^3

2.3 $3,8$ mol·dm^{-3}

3. 4 mol·dm^{-3}

4. $36,46$ g

5. $19,62$ g

7.7 Exercise 7

1. $3,1$ g

2. n = $0,54$ mol and Volume = $12,1$ dm^3

3.1 $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

3.2 $90,06$ g