The mole

6

The big picture
- The \( M_r \) of hydrogen, \( H_2 \), is 2. The \( M_r \) of oxygen, \( O_2 \), is 32.
- The mole concept tells us that if we could weigh out exactly 2 g of oxygen and 32 g of hydrogen, they would both contain the same number of molecules.
- This idea helps us to carry out a wide range of calculations, in chemistry.

6.1 Moles and masses

The mole concept
- The mass of an atom of carbon-12 is taken as 12. So the \( A_r \) of carbon is 12.
- A magnesium atom has twice this mass. So the \( A_r \) of magnesium is 24.
- It follows that 24 grams of magnesium contains the same number of atoms as 12 grams of carbon does.

24 g of magnesium is called a mole of magnesium atoms.

A mole of a substance is the amount that contains as many elementary units as the number of atoms in 12 g of carbon-12.

In fact we know how many atoms this is. It is a huge number: \( 6.02 \times 10^{23} \).

So 24 g of magnesium contains \( 6.02 \times 10^{23} \) magnesium atoms. \( 6.02 \times 10^{23} \) is called the Avogadro constant after the scientist who deduced it.

We can apply the same logic to any element, and any compound. Look at these examples:

<table>
<thead>
<tr>
<th>Substance</th>
<th>( A_r ) or ( M_r )</th>
<th>So 1 mole of it is ...</th>
<th>... and contains ...</th>
</tr>
</thead>
<tbody>
<tr>
<td>helium, He</td>
<td>4</td>
<td>4 grams</td>
<td>( 6.02 \times 10^{23} ) helium atoms</td>
</tr>
<tr>
<td>oxygen, O₂</td>
<td>32</td>
<td>32 grams</td>
<td>( 6.02 \times 10^{23} ) oxygen molecules</td>
</tr>
</tbody>
</table>

So a mole of a substance is its \( A_r \) or \( M_r \) given in grams.

Sample calculations

- mass of a given number of moles (g) = mass of 1 mole \( \times \) number of moles

<p>| |</p>
<table>
<thead>
<tr>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1. What is the mass of 6 moles of helium atoms? ( 4 \times 6 = 24 \text{ g} )</td>
</tr>
<tr>
<td>2. What is the mass of 0.5 moles of oxygen molecules? ( 32 \times 0.5 = 16 \text{ g} )</td>
</tr>
</tbody>
</table>

- number of moles in a given mass = \( \frac{\text{mass (g)}}{\text{mass of 1 mole}} \)

<p>| |</p>
<table>
<thead>
<tr>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1. How many moles of helium atoms are in 12 g of helium? ( 12 \div 4 = 3 \text{ moles} )</td>
</tr>
<tr>
<td>2. How many moles of oxygen molecules are in 80 g of oxygen? ( 80 \div 32 = 2.5 \text{ moles} )</td>
</tr>
</tbody>
</table>

Quick check for 6.1
(Answers on page 165)

1. How many molecules are there in 1 mole of molecules?
2. Find the mass of:
   a. 1 mole of chlorine, Cl₂
   b. 0.6 moles of sodium chloride, NaCl
3. Find how many moles of molecules there are in:
   a. 8 g of hydrogen, H₂
   b. 2.55 g of hydrogen sulfide, H₂S
   c. 3.8 g of magnesium chloride, MgCl₂

\( A_r \) values
- H = 1
- Na = 23
- Mg = 24
- S = 32
- Cl = 35.5

12 g of carbon powder contain 1 mole of carbon atoms \((6.02 \times 10^{23} \text{ atoms})\).
6.2 Finding empirical and molecular formulae

- The empirical formula shows the simplest ratio in which the atoms in a compound are combined. For example, the empirical formula of hydrogen peroxide is HO.
- The molecular formula shows the actual number of atoms that combine to form a molecule. The molecular formula of hydrogen peroxide is H₂O₂.

Finding the empirical formula
First, do an experiment to find out what masses of the elements combine to form the compound. Then the empirical formula can be worked out. The steps are:

1. **Find the masses of the elements that combine, in grams.**
2. **Convert the masses to moles of atoms.**
3. **Work out the simplest ratio in which the atoms combine.**
4. **That is the empirical formula.**

**Example** In an experiment, 80 g of carbon combine with 20 g of hydrogen to form a compound. What is its empirical formula?

<table>
<thead>
<tr>
<th>Elements that combine</th>
<th>carbon</th>
<th>hydrogen</th>
</tr>
</thead>
<tbody>
<tr>
<td>Masses that combine</td>
<td>80 g</td>
<td>20 g</td>
</tr>
<tr>
<td>Relative atomic masses (Aᵣ)</td>
<td>12</td>
<td>1</td>
</tr>
<tr>
<td>Moles of atoms that combine</td>
<td>(\frac{80}{12} = 6.67)</td>
<td>(\frac{20}{1} = 20)</td>
</tr>
<tr>
<td>Ratio in which atoms combine</td>
<td>6.67 : 20, or 1:3 in its simplest form</td>
<td></td>
</tr>
</tbody>
</table>

So the empirical formula of the compound is CH₃.

Finding the molecular formula from the empirical formula
The actual formula mass of a compound, \(M_r\), can be found using a machine called a mass spectrometer. From the formula mass and empirical formula, you can work out the molecular formula. The steps are:

1. **Calculate the mass of 1 mole of the compound, using the empirical formula.** This is the empirical mass.
2. **Divide \(M_r\) by the empirical mass, to find how many times bigger \(M_r\) is. Your answer is a number. Let’s call it \(n\).**
3. **Multiply the number of each atom in the empirical formula by \(n\). This gives the molecular formula.**

**Example** The formula mass of a compound was found to be 30. Its empirical formula was found to be CH₃. What is the molecular formula for the compound?

\[
\frac{M_r}{\text{empirical mass}} = \frac{30}{15} = 2
\]

So multiply each atom in the empirical formula CH₃ by 2. The molecular formula is C₂H₆. (Note: you do not put the 2 in front of the formula.)

**Quick check for 6.2**
(Answers on page 165)
1. 1.2 g of carbon combined with 0.1 g of hydrogen, to form a compound.
   a. Find the empirical formula for the compound.
   b. Its \(M_r\) value was found to be 78. Find its molecular formula.
2. In one oxide of nitrogen, the ratio of oxygen atoms to nitrogen atoms is 2.1. Its \(M_r\) is 46. Give the empirical and molecular formulae for this compound.

\[Aᵣ \text{ values:} \]
\[H = 1, \quad C = 12, \quad N = 14, \quad O = 16\]
6.3 Finding masses from chemical equations

**Moles and equations**

The mole concept lets us work out masses that react. For example:

<table>
<thead>
<tr>
<th>Ratio of particles in the equation</th>
<th>(2 \text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{H}_2\text{O} (l))</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 molecules</td>
<td>1 molecule</td>
</tr>
<tr>
<td>2 molecules</td>
<td>2 molecules</td>
</tr>
</tbody>
</table>

Scaling up to moles gives
the mole ratio

<table>
<thead>
<tr>
<th>2 moles of molecules</th>
<th>1 mole of molecules</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mole of molecules</td>
<td>2 moles of molecules</td>
</tr>
</tbody>
</table>

or

<table>
<thead>
<tr>
<th>1 mole of molecules</th>
<th>0.5 mole of molecules</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.5 mole of molecules</td>
<td>1 mole of molecules</td>
</tr>
</tbody>
</table>

Or change moles to masses,
using \(A_r\) and \(M_r\)

<table>
<thead>
<tr>
<th>4 g</th>
<th>32 g</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 g</td>
<td>16 g</td>
</tr>
<tr>
<td>------------------------------------</td>
<td>-------------------------------------------------</td>
</tr>
<tr>
<td>36 g</td>
<td>36 g</td>
</tr>
<tr>
<td></td>
<td>36 g</td>
</tr>
</tbody>
</table>

Calculations from equations

The equation tells you the mole ratio of the substances in the reaction.

<table>
<thead>
<tr>
<th>Equation</th>
<th>(\text{CH}_4 (g) + 2 \text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2 \text{H}_2\text{O} (l))</th>
</tr>
</thead>
</table>

If you know the actual mass of one substance, you can change this to moles.

Then use the mole ratio to find the masses of the other substances.

Example 1 In the complete combustion of methane (\(\text{CH}_4\)), what mass of oxygen combines with 64 g of methane, and how much carbon dioxide is produced?

<table>
<thead>
<tr>
<th>Equation</th>
<th>(\text{CH}_4 (g) + 2 \text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2 \text{H}_2\text{O} (l))</th>
</tr>
</thead>
</table>

Mole ratio

\(1\)  \(2\)  \(1\)  \(2\)

Moles of known substance \(M_r\) of \(\text{CH}_4 = 16\) so \(64\) g of \(\text{CH}_4 = (64 ÷ 16)\) moles = 4 moles

Using the mole ratio

4 moles of \(\text{CH}_4\) so 8 moles of \(\text{O}_2\) and 4 moles of \(\text{CO}_2\)

Change moles to masses

\(M_r\) of \(\text{O}_2 = 32; \ (8 \times 32) = 256\) g \(M_r\) of \(\text{CO}_2 = 44; \ (4 \times 44) = 176\) g

So ...

256 g of oxygen combines with 64 g of methane and produces 176 g of carbon dioxide

Example 2 Aluminium burns in oxygen. What mass of oxygen combines with 100 g of aluminium, and how much aluminium oxide is produced?

<table>
<thead>
<tr>
<th>Equation</th>
<th>(4\text{Al} (s) + 3\text{O}_2 (g) \rightarrow 2\text{Al}_2\text{O}_3 (s))</th>
</tr>
</thead>
</table>

Mole ratio

\(4\)  \(3\)  \(2\)

but aluminium is the known substance, so make it \(1\) in the ratios:

\(1\)  \(3 ÷ 4 = 0.75\)  \(2 ÷ 4 = 0.5\)

Moles of known substance

\(100 ÷ 27 = 3.704\) moles of \(\text{Al}\)

Using the mole ratio

\(3.704\) of \(\text{Al}\) so \((3.704 \times 0.75) = 2.778\) of \(\text{O}_2\)

and \((3.704 \times 0.5) = 1.852\) of \(\text{Al}_2\text{O}_3\)

Change moles to masses

\(\text{O}_2; \ (2.778 \times 32) = 88.9\) g\(\text{Al}_2\text{O}_3; \ (1.852 \times 102) = 188.9\) g

So ...

88.9 g of oxygen combines with 100 g of aluminium and produces 188.9 g of aluminium oxide

**Ar values**

- \(H = 1\)
- \(C = 12\)
- \(O = 16\)
- \(O = 16\)
- \(Al = 27\)
Concentration: the amount of a solute (in grams or moles) in 1 dm$^3$ of solution.

Units used: g/dm$^3$ or mol/dm$^3$

Remember: 1 dm$^3$ = 1 litre = 1000 cm$^3$

A solution containing 1 mole of a solute in 1 dm$^3$ of solution is often called a molar solution or a 1 M solution.

A 2 M solution contains 2 moles in 1 dm$^3$ of solution, and so on.

## Quick check for 6.3

(Answers on page 165)

1. How much oxygen is required for the complete combustion of 128 g of methane?
2. What mass of hydrogen will combine with 56 g of nitrogen in this reaction?
   \[ \text{N}_2 (g) + 3\text{H}_2 (g) \rightarrow 2\text{NH}_3 (g) \]
3. What mass of ammonia could be produced from 56 g of nitrogen?
4. Magnesium carbonate breaks down on heating, like this:
   \[ \text{MgCO}_3 (s) \rightarrow \text{MgO} (s) + \text{CO}_2 (g) \]
   When 10 g of magnesium carbonate is heated, what mass of each product forms?

### 6.4 Calculations about solutions

<table>
<thead>
<tr>
<th>To find</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>concentration in mol/dm$^3$</td>
<td>concentration in mol/dm$^3$ = amount of solute (mol) / volume of solution (dm$^3$)</td>
</tr>
</tbody>
</table>
|                        | 1. What is the concentration of a solution containing 2 moles of a compound in 0.5 dm$^3$?  
|                        | 2. What is the concentration of a solution containing 0.1 moles of a compound in 40 cm$^3$?  
|                        | 0.1 + 0.04 = 2.5 mol/dm$^3$                                               |
| concentration in g/dm$^3$ | concentration in g/dm$^3$ = concentration in mol/dm$^3$ × $M_r$          |
|                        | 1. What is the concentration in g/dm$^3$ of a 4 mol/dm$^3$ solution of sodium hydroxide? ($M_r$ = 40)  
|                        | 4 mol/dm$^3$ × 40 = 160 g/dm$^3$                                          |
|                        | 2. What is the concentration in g/dm$^3$ of a 2.5 mol/dm$^3$ solution of sulfuric acid? ($M_r$ = 98)  
|                        | 2.5 mol/dm$^3$ × 98 = 245 g/dm$^3$                                        |
| volume of solution     | volume of solution (dm$^3$) = amount of solute (mol) / concentration (mol/dm$^3$) |
|                        | 1. What volume of a 2 mol/dm$^3$ solution contains 0.6 moles?  
|                        | 0.6 ÷ 2 = 0.3 dm$^3$ or 300 cm$^3$                                         |
|                        | 2. What volume of a 0.5 mol/dm$^3$ solution contains 2 moles?  
|                        | 20.5 = 4 dm$^3$ or 4000 cm$^3$                                            |
| moles of solute        | amount of solute (mol) = concentration (mol/dm$^3$) × volume of solution (dm$^3$) |
|                        | 1. How many moles of solute are there in 2 dm$^3$ of a 2 mol/dm$^3$ solution?  
|                        | 2 × 2 = 4 moles                                                            |
|                        | 2. How many moles of solute are there in 50 cm$^3$ of a 0.25 mol/dm$^3$ solution?  
|                        | 0.25 × 0.05 = 0.0125 moles                                                 |

### Use the calculation triangle!

\[
\text{amount (mol)} = \text{conc. (mol/dm$^3$)} \times \text{vol. (dm$^3$)}
\]
### Calculating the volume of a solution in a reaction

You use the same logic as in section 6.3.

#### Example 1
What volume of 0.5 mol/dm³ hydrochloric acid reacts with 0.12 g of magnesium?

<table>
<thead>
<tr>
<th>Equation</th>
<th>Mg (s) + 2HCl (aq) → MgCl₂ (aq) + H₂ (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mole ratio</td>
<td>1 2</td>
</tr>
<tr>
<td>Moles of known substance</td>
<td>Mg: 0.12 ÷ 24 = 0.005 moles</td>
</tr>
<tr>
<td>Using the mole ratio</td>
<td>0.005 moles of Mg so (0.005 x 2) = 0.01 moles of HCl</td>
</tr>
<tr>
<td>Volume</td>
<td>volume = mol ÷ concentration (see the calculation triangle) so volume of HCl solution = (0.01 ÷ 0.5) = 0.02 dm³ or 20 cm³</td>
</tr>
<tr>
<td>So …</td>
<td>20 cm³ of 0.5 mol/dm³ hydrochloric acid reacts with 0.12 g of magnesium</td>
</tr>
</tbody>
</table>

#### Example 2
What volume of 2.0 mol/dm³ sodium hydroxide neutralises 25 cm³ of 0.5 mol/dm³ sulfuric acid?

<table>
<thead>
<tr>
<th>Equation</th>
<th>2NaOH (aq) + H₂SO₄ (aq) → Na₂SO₄ (aq) + 2H₂O (l)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mole ratio</td>
<td>1 2</td>
</tr>
<tr>
<td>Moles of known substance</td>
<td>H₂SO₄: 0.5 x 0.025 = 0.0125 moles</td>
</tr>
<tr>
<td>Using the mole ratio</td>
<td>0.0125 moles of H₂SO₄ so (2 x 0.0125) = 0.0025 moles of NaOH</td>
</tr>
<tr>
<td>Volume</td>
<td>volume = mol ÷ concentration (see the calculation triangle) so volume of NaOH solution = (0.0025 ÷ 2) = 0.00125 dm³ or 12.5 cm³</td>
</tr>
<tr>
<td>So …</td>
<td>12.5 cm³ of 2 mol/dm³ sodium hydroxide neutralises 25 cm³ of 0.5 mol/dm³ of sulfuric acid.</td>
</tr>
</tbody>
</table>

Quick check for 6.4

1. What is the concentration of a solution containing:
   a. 0.25 moles in 2 dm³ of solution?
   b. 0.6 moles in 100 cm³ of solution?

2. What volume of a 0.5 mol/dm³ solution contains:
   a. 2 moles?
   b. 0.01 moles?

3. Find the number of moles of solute in:
   a. 3 dm³ of a 0.05 mol/dm³ solution
   b. 40 cm³ of a 2.5 mol/dm³ solution

4. How many grams of potassium hydroxide (M_r = 56) are there in:
   a. 2 dm³ of a molar solution?
   b. 20 cm³ of a 0.5 mol/dm³ solution?

5. What volume of 2 mol/dm³ nitric acid will react with 10 g of calcium carbonate (M_r = 100)? The equation for the reaction is:
   \[ \text{CaCO}_3 (s) + 2\text{HNO}_3 (aq) \rightarrow \text{Ca(NO}_3)_2 (aq) + \text{H}_2\text{O (l)} + \text{CO}_2 (g) \]

6. What volume of 0.5 mol/dm³ sulfuric acid will neutralise 15 cm³ of sodium hydroxide solution, of concentration 1 mol/dm³?
### 6.5 Calculating volumes of gases

**Key idea:** 1 mole of a gas has a volume of 24 dm³ at room temperature and pressure (rtp). So 24 dm³ is called the *molar gas volume*.

<table>
<thead>
<tr>
<th>To find</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>volume</td>
<td>volume of gas at rtp (dm³) = no. of moles × 24 dm³</td>
</tr>
<tr>
<td></td>
<td>1. What volume does 3 moles of a gas occupy, at rtp? 3 × 24 = 72 dm³ or 72 000 cm³</td>
</tr>
<tr>
<td></td>
<td>2. What volume does 14 g of nitrogen (N₂) occupy, at rtp? (A_r: N = 14) Number of moles of nitrogen gas = 14 ÷ 28 = 0.5 Volume = 0.5 × 24 = 12 dm³ or 12 000 cm³</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>moles</th>
<th>no. of moles = volume of gas (dm³) / 24 dm³</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>1. How many moles are there in 48 dm³ of a gas, at rtp? 48 ÷ 24 = 2 moles</td>
</tr>
<tr>
<td></td>
<td>2. How many moles are there in 120 cm³ of a gas, at rtp? 120 ÷ 24000 = 0.005 moles</td>
</tr>
</tbody>
</table>

#### Calculating the volume of a gas in a reaction

**Example 1** What volume of hydrogen forms, at rtp, when 0.12 g of magnesium reacts with hydrochloric acid?

**Equation**

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

**Mole ratio**

\[1 \quad 1\]

**Moles of known substance**

\[A_r \text{ of Mg} = 24 \text{ so } (0.12 \div 24) = 0.005 \text{ moles of Mg}\]

**Using the mole ratio**

0.005 moles of Mg so 0.005 moles of H₂

**Volume at rtp**

\[\text{volume} = \text{no. of mol} \times 24 \text{ dm}^3\]

so volume of H₂ = (0.005 × 24) = 0.12 dm³ or 120 cm³

**So …** 120 cm³ of hydrogen is released.

**Example 2** What volumes of nitrogen and hydrogen give 50 cm³ of ammonia gas, at rtp?

1 mole of every gas has the same volume at rtp (24 dm³).

So in reactions that involve only gases, we can say: **volume ratio = mole ratio**.

**Equation**

\[\text{N}_2 (g) + 3\text{H}_2 (g) \rightarrow 2\text{NH}_3 (g)\]

**Volume ratio = mole ratio**

\[1 \quad 3 \quad 2\]

**Using the volume ratio**

25 cm³ of N₂ 75 cm³ of H₂ 50 cm³ of NH₃

**So …** 25 cm³ of nitrogen and 75 cm³ of hydrogen give 50 cm³ of ammonia.

**Quick check for 6.5** *(Answers on page 165)*

1. Calculate the volume at rtp of:
   a. 6 g of oxygen, O₂
   b. 6.4 g of sulfur dioxide, SO₂ \((A_r: O = 16, S = 32)\)

2. What volume of carbon dioxide at rtp is released when 5.3 g of sodium carbonate \((M_r = 106)\) reacts with hydrochloric acid? The equation is:
   \[\text{Na}_2\text{CO}_3 (s) + 2\text{HCl} (aq) \rightarrow 2\text{NaCl} (aq) + \text{H}_2\text{O} (l) + \text{CO}_2 (g)\]

3. Hydrogen and chlorine react like this:
   \[\text{H}_2 (g) + \text{Cl}_2 (g) \rightarrow 2\text{HCl} (g)\]

   What volumes of hydrogen and chlorine give 250 dm³ of hydrogen chloride?
% yield

- The **yield** is the amount of product obtained from a chemical reaction.
- We can calculate the amount of product from the equation (the **calculated mass**).
- But the **actual mass** we get is less – for example because some reactant remains unreacted, or some product is lost in the separation process.

\[
\text{The } % \text{ yield } = \frac{\text{actual mass}}{\text{calculated mass}} \times 100\%
\]

**Example** In an experiment, 100 g of aluminium is burnt in oxygen, giving aluminium oxide. 150 g of aluminium oxide is obtained. What is the % yield for the experiment?

Equation

\[
4\text{Al (s)} + 3\text{O}_2 (\text{g}) \rightarrow 2\text{Al}_2\text{O}_3 (\text{s})
\]

Mole ratio

\[
4 \quad 3 \quad 2
\]

Calculated mass

The calculated mass of aluminium oxide, from 100 g of Al, is 188.9 g. (See Example 2 on page 41 for the working for this.)

Actual mass

150 g of aluminium oxide

% yield

\[
(150 \div 188.9) \times 100 = 79.4 \%
\]

So ... the % yield for the experiment was 79.4 %.

% purity

- When we carry out a reaction, we obtain a certain mass of product.
- But it is impure. There may still be some reactant mixed with it, for example.

\[
\text{The } % \text{ purity of the product from the reaction } = \frac{\text{mass of pure product}}{\text{mass of impure product}} \times 100 \%
\]

**Example** Aluminium was burned in oxygen, to give aluminium oxide. The equation is:

\[
4\text{Al (s)} + 3\text{O}_2 (\text{g}) \rightarrow 2\text{Al}_2\text{O}_3 (\text{s})
\]

150 g of product was obtained. But it was found to contain 5 g of impurities. What is the % purity of the aluminium oxide obtained?

<table>
<thead>
<tr>
<th>Mass of impure product obtained</th>
<th>150 g</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of pure product</td>
<td>150 - 5 = 145 g</td>
</tr>
<tr>
<td>% purity</td>
<td>(145 ÷ 150) × 100 = 96.7%</td>
</tr>
<tr>
<td>So ...</td>
<td>the aluminium oxide obtained was 96.7% pure.</td>
</tr>
</tbody>
</table>

Quick check for 6.6

(Answers on page 165)

1. What would be the ideal yield, in an experiment?
2. According to the equation, a reaction should give you 60 g of product. But you obtain only 45 g. What is the % yield?
3. You obtain 7.5 g of compound X, in an experiment. After purifying, its mass is 6 g. What was the % purity of X, in your experiment?
4. Some sea water is evaporated. The salt obtained is 91% sodium chloride. How much sodium chloride would you obtain from 20 tonnes of this salt?
5. Electrolysis of 102 kg of aluminium oxide gave 45 kg of aluminium. What was the % yield? (A; Al = 27, O = 16)
Questions on Section 6

Answers for these questions are on page 165.

**Extended curriculum**

1. Marble reacts with dilute hydrochloric acid according to the equation:

   \[ \text{CaCO}_3 (s) + 2\text{HCl} (aq) \rightarrow \text{CaCl}_2 (aq) + \text{CO}_2 (aq) + \text{H}_2\text{O} (l) \]

   One piece of marble, 0.3 g, was added to 5 cm³ of hydrochloric acid, 1.00 mol/dm³.

   a. Which reagent is in excess? Give a reason for your choice.

   b. Use your answer to a to calculate the maximum volume of carbon dioxide produced, measured at rtp.

   The volume of one mole of any gas is 24 dm³ at room temperature and pressure (rtp).

2. Crystals of sodium sulfate-10-water, Na₂SO₄.10H₂O, are prepared by titration.

   a. 25.0 cm³ of aqueous sodium hydroxide is pipetted into a conical flask.

      A few drops of an indicator are added. Using a burette, dilute sulfuric acid is slowly added until the indicator just changes colour. The volume of acid needed to neutralise the alkali is noted.

      Suggest how you would continue the experiment to obtain pure, dry crystals of sodium sulfate-10-water.

   b. Using 25.0 cm³ of aqueous sodium hydroxide, 2.24 mol/dm³, 3.86 g of crystals were obtained.

      \[ 2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O} \]
      \[ \text{Na}_2\text{SO}_4 + 10\text{H}_2\text{O} \rightarrow \text{Na}_2\text{SO}_4\cdot10\text{H}_2\text{O} \]

      Calculate:

      i. The number of moles of NaOH used.
      ii. The maximum number of moles of Na₂SO₄.10H₂O that could be formed.
      iii. The maximum yield of sodium sulfate-10-water.
      iv. The percentage yield.

3. Basic lead(II) carbonate has a formula of the type \(x\text{PbCO}_3\cdot y\text{Pb(OH)}_2\) where \(x\) and \(y\) are whole numbers. Determine \(x\) and \(y\) from the following information.

   \[ \text{PbCO}_3 \rightarrow \text{PbO} + \text{CO}_2 \quad \text{Pb(OH)}_2 \rightarrow \text{PbO} + \text{H}_2\text{O} \]

   When heated, the basic lead(II) carbonate gave 2.112 g of carbon dioxide and 0.432 g of water.

   a. Calculate the number of moles of \(\text{CO}_2\) formed.
   b. Calculate the number of moles of \(\text{H}_2\text{O}\) formed.
   c. The formula of basic lead(II) carbonate is …………………. 

   Mass of one mole of \(\text{CO}_2\) = 44 g
   Mass of one mole of \(\text{H}_2\text{O}\) = 18 g
4 a Define the following:
   i the mole
   ii the Avogadro constant

b Which two of the following contain the same number of molecules?
Show how you arrived at your answer.
2.0 g of methane, CH₄
8.0 g of oxygen, O₂
2.0 g of ozone, O₃
8.0 g of sulfur dioxide, SO₂
c 4.8 g of calcium is added to 3.6 g of water. The following reaction occurs.
   Ca + 2H₂O → Ca(OH)₂ + H₂
   i Calculate the number of moles of Ca and the number of moles of H₂O
   ii Which reagent is in excess? Explain your choice.
   iii Calculate the mass of the reagent named in ii which remained at the end
   of the experiment.

   CIE 0620 June ’13 Paper 3 Q8

5 A sample of rust had the following composition:
51.85 g of iron    22.22 g of oxygen    16.67 g of water.
Calculate the following and then write the formula for this sample of rust.

a number of moles of iron atoms, Fe
b number of moles of oxygen atoms, O
c number of moles of water molecules, H₂O
d simplest mole ratio Fe : O : H₂O
e formula for this sample of rust is …………

   CIE 0620 June ’12 Paper 3 Q8b

6 There are three possible equations for the thermal decomposition of sodium hydroncarbonate.
   2NaHCO₃(s) → Na₂O(s) + 2CO₂(g) + H₂O(g) equation 1
   NaHCO₃(s) → NaOH(s) + CO₂(g) equation 2
   2NaHCO₃(s) → Na₂CO₃(s) + CO₂(g) + H₂O(g) equation 3

The following experiment was carried out to determine which one of the above is the
correct equation.

A known mass of sodium hydroncarbonate was heated for ten minutes.
It was then allowed to cool and was weighed.

Results
Mass of sodium hydroncarbonate = 3.36 g
Mass of the residue = 2.12 g

a Calculate the number of moles of NaHCO₃ used.
b i If residue is Na₂O, calculate the number of moles of Na₂O.
   ii If residue is NaOH, calculate the number of moles of NaOH.
   iii if residue is Na₂CO₃, calculate the number of moles of Na₂CO₃.
c Use the number of moles calculated in a and b to decide which one of the three
equations is correct. Explain your choice.

   CIE 0620 Nov ’11 Paper 3 Q7c

7 Two gases react as shown.
   X₂ + Y₂ → 2XY
For this reaction, what is the missing number below?

   volume of product
   total volume of reactants = …… at rtp

   CIE 0620 June ’13 Paper 3 Q5

---

Extended

The mole

<table>
<thead>
<tr>
<th>Substance</th>
<th>M_r (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaHCO₃</td>
<td>84</td>
</tr>
<tr>
<td>Na₂O</td>
<td>62</td>
</tr>
<tr>
<td>NaOH</td>
<td>40</td>
</tr>
<tr>
<td>Na₂CO₃</td>
<td>106</td>
</tr>
</tbody>
</table>