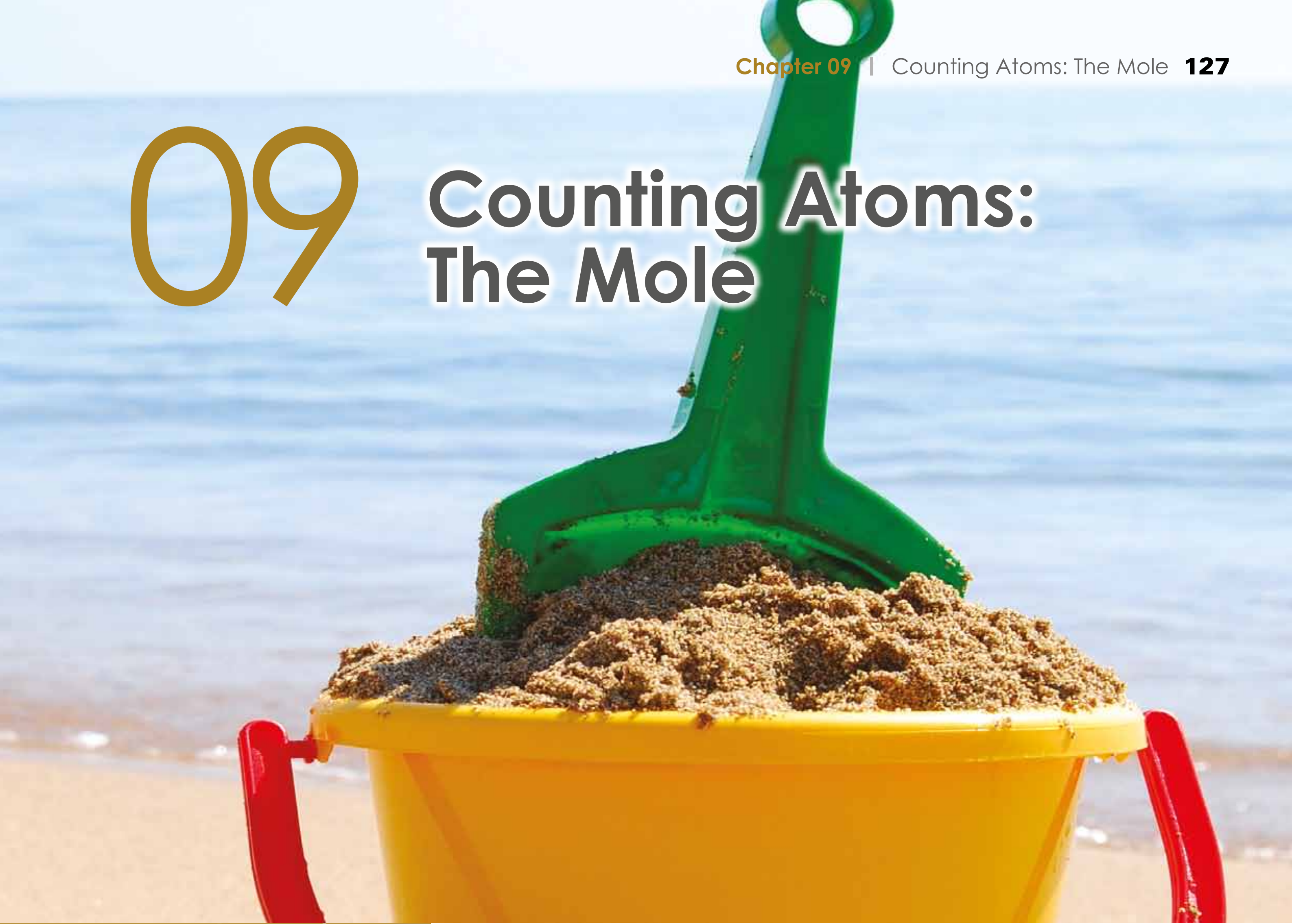


09 Counting Atoms: The Mole

Introduction

In this chapter, students learn how to count particles in chemistry using the mole and to link it with the concepts of molar mass and molar volumes of gases. Students also find out how to derive the empirical formula of a compound from experimental data and how to do basic calculations related to empirical formula.

Chapter Opener (page 129)

1. To open the chapter, the following questions could be discussed. Precise answers are not needed at this stage.

What units do we often use to count shoes and eggs? Suggest a reason why these units are used.

Answer: Shoes — pair. Eggs — dozen. These units make it easier when counting large numbers of shoes or eggs.

It is very difficult to count the numbers of particles in substances. Suggest why.

Answer: The particles in substances are very small so there will be a very large number of particles to count.

What does the term concentration indicate?

Answer: It indicates the amount of a substance/solute dissolved to give a solution. The greater the amount of solute, the greater the concentration of the solution.

2. Carry out an 'Inquiry Preview.'

Learning Outcomes

After completing this chapter, the students should be able to:

- measure number of atoms, molecules or ions in moles
- use molar mass to calculate number of moles and the mass of a substance
- calculate empirical and molecular formulae from relevant data
- use Avogadro's Law and molar volume in calculations involving gases
- calculate the concentration of solutions

Teaching notes for

9.1 How Do We Count Particles in Chemistry? (page 130)

Stimulation

The mole is a difficult concept for many students. Perhaps the best way to think of it is a special number or unit like a pair, a dozen, a ream (of paper), etc. Bring into the classroom a pair of socks, an egg carton, a ream of paper and discuss these as convenient collective units for counting different objects. Lead into the use of the mole as the collective unit chemists use for counting numbers of particles. Introduce the value of the mole, which students must memorise. Draw this value for the mole in its full form on the chalkboard to give students an idea of just how large it is.

This discussion could be linked with the following questions:

1. Give examples of other units used to count everyday objects [*elaborating*]
2. A drop of water contains about 2000 trillion molecules. Write out '2000 trillion' in full to appreciate how large the number is. Is this number larger or smaller than the number in one mole [*visualising, comparing*]
3. Suggest a reason why it is more convenient to use moles in Chemistry to count particles. [*explaining*]
4. We could use moles to count sheets of paper or numbers of people. However, we do not. Why? [*explaining*]

Answers:

1. For example: The unit for football players is a team. The number of players in one team is 11. The unit for paper clips is a gross and the number of paper clips in one gross is 144 (this unit is less commonly used now).
2. 2 000 trillion = $2\ 000 \times 10^{12}$ = 2 000 000 000 000 000.

   This number is much less than that in 1 mole (as the mass of 1 drop of water is less than 18 g).
3. A smaller number (of moles) is used instead of a very large number (of particles). It is easier for us to comprehend small numbers (as these are within our everyday experiences) than huge numbers (which are the amounts of substances chemists use). Also, smaller units are easier to handle in calculations.
4. The numbers of moles of papers or people would be too small for us to deal with. For example, there are about 6 000 000 000 people in the world or 1×10^{-14} moles of people!

Teaching pointers

1. Strictly speaking, the mole is an 'amount' of a substance and not a number. For example, it is more accurate to say there are 6×10^{23} magnesium atoms per mole of magnesium atoms. However, this should not be mentioned to students who, at this level, can regard the mole as just a number.
2. The use of 'mol' as the symbol for 'mole' may be new to students. At 'O' Level, it does not matter if students use 'mole', 'mol' or 'moles'.
3. Examples to illustrate just how large a mole is:
   (a) Write out '2000 trillion' in full on the board to help students appreciate how large the number is. Underneath it, write, in full, the value of one mole and get the class to calculate the number of moles of water molecules in the drop.
   (b) Show the class an iron (steel) nail. The nail contains many millions of atoms. Ask the class this question: Why are there so many atoms in one nail? Get students to estimate the number of atoms in the nail. Then tell them the answer and see who has made the most accurate guess. [An iron nail has about 10^{22} atoms (depending on size).] Students usually underestimate this number.
4. Then ask the class why it is more convenient for chemists to use moles to count particles. **Answer:** A smaller number of moles is used instead of a very large number of particles. Humans can comprehend small numbers (as these are within their everyday experience) but not huge numbers (which occur in the amounts of substances chemists use). Also, the smaller unit is easier to use in calculations.
5. We could use moles to count anything including numbers of people. The world's population is about 7 billion (7×10^9). Ask the class why moles are not used. Answer: The numbers of moles would be too small for us to deal with; 7 billion people in the world is only about 1.17×10^{-14} moles of people!
6. The Extension question ('A mole of coins,' on page 146 of the Textbook) is a useful way to make students aware of just how large 6×10^{23} is.

Teaching pointers

9.2 How are Moles and Number of Particles Related? (page 131)

1. Emphasise that the mole is just a *number* of atoms, molecules or ions.
2. The analogy of a mole of coins in the Extension exercise on page 146 could be used to show the idea of just how large the value of a mole is.
3. It is extremely useful to visualise the quantity of one mole of an element or a compound. This exercise gives real and concrete meaning to the term *mole*, without which the mole is just an abstract concept.

   To carry out this exercise, you may show labelled bottles containing 1 mole of various substances (such as those in Figure 9.2 on page 131 of the Textbook). Suitable examples are:

   – 12 g of carbon (charcoal)
   – 24 g of magnesium ribbon
   – 32 g of sulfur powder
   – 56 g of iron powder/filings
   – 18 g of water
   – 32 g of oxygen (fill a 24 dm^3 container with air and labelled it as '1 mole of oxygen molecules')
   – 58.5 g of sodium chloride

   Note: At this moment, do *not* mention the *masses* of elements or compounds. You may mention them in Section 9.3 when molar mass is discussed.

4. In doing numerical calculations, it is important that students not only obtain the correct answers, but also understand the working. Get students to explain (qualitatively and not just repeating the numerical steps) the reason behind each step of the working. Research has shown that this kind of exercise will lead to a greater understanding and problem-solving ability.

Skills Practice (page 132)

1. (a) Number of moles of particles in 2.4×10^{23} atoms of H = $2.4 \times 10^{23} \div 6.02 \times 10^{23}$
= 0.4 mol

(b) Number of moles of particles in 4.2×10^{24} atoms of H = $4.2 \times 10^{24} \div 6.02 \times 10^{23}$
= 7.0 mol

(c) Number of moles of particles in 3×10^{22} atoms of H = $3 \times 10^{22} \div 6.02 \times 10^{23}$
= 0.05 mol

2. (a) Number of particles in 3 moles of C atoms = $3 \times 6.02 \times 10^{23}$
= 1.8×10^{24} atoms

(b) Number of particles in 0.75 moles of NH_3 molecules = $0.75 \times 6.02 \times 10^{23}$
= 4.5×10^{23} molecules

(c) Number of particles in 25 moles of NaC*l* = $25 \times 6.02 \times 10^{23}$
= 1.5×10^{25} units

3. (a) 1 mol of C atoms

(b) 2.5 mol of C atoms

(c) 0.13 mol of C atoms

Teaching pointers

9.3 What is Molar Mass? (page 132)

1. This section shows that chemists do not actually count numbers of moles. Rather, they calculate the number using knowledge of masses and molar masses of substances.
2. Molar mass, like the mole, is a physical quantity. As it is a mass, its base unit is the kilogram, though the gram is commonly used.
3. Show the class the same labelled bottles containing 1 mole of various substances that were used in Section 9.2. This time link the mole of the substances to their molar masses. (S – 32 g, Cu – 63.5 g, Hg – 201 g, H_2O – 18 g, $C_{12}H_{22}O_{11}$ – 342 g.)
4. In Activity 9.1 of the Theory Workbook, students get some practice in calculating the numbers of moles of substances from their masses and relative molecular masses.
5. In Exercise 9.1, students get practice in calculating numbers of moles of substances from mass and relative molecular mass.
6. At this stage, you may like to introduce the idea of using spreadsheets for chemical calculations. A worksheet (Additional Exercise 1) is provided at the end of this chapter. You may photocopy and distribute the worksheet to the class. Refer also to the notes under 'Calculations with a Spreadsheet' on the next page.

(page 132)

Mystery Clue

Molar masses – Fe_2O_3 = 160 g, Fe_3O_4 = 232 g.

Chemistry Inquiry (page 133)

How Many Moles of Particles do the Substances Contain?

To teachers, the steps used to solve these problems might seem to be simple. However, research with students into the solving of such problems has shown this not to be the case. For this reason, the steps in the problem solving strategies have been stated explicitly.

One learning strategy is for the teacher to write the question on the board, then get the class to close their textbooks and in working in groups, to try to work out the steps needed to solve the problem.

Some students also have the misconception that if the masses of different substances are the same, then the number of moles would also be the same. Discussion question 2 should help them overcome this misconception.

Group Discussion

1. It is inconvenient to count particles because they are so small and there are so many of them in a substance. It is much easier to find the mass of a substance. Then, by using relative atomic masses, the number of (moles of) atoms or molecules in the substance can be easily determined.
2. The numbers of moles of particles in 50 g of each substance are: Copper 0.00781 mol of Cu *atoms* ($\frac{0.5}{64}$); Water 0.0278 mol of H_2O *molecules* ($\frac{0.5}{18}$); Oxygen 0.0156 mol of O_2 *molecules* ($\frac{0.5}{32}$).
3. From mass and molar mass, the number of moles of particles in a substance is calculated.

Skills Practice (page 134)

1. **(a)** 40 g **(b)** 32 g **(c)** 23 g **(d)** 32 g **(e)** 18 g **(f)** 64 g **(g)** 100 g

2. **(a)** Number of moles of atoms in 32 g of oxygen atoms = 32 ÷ 16 = 2 mol
(b) Number of moles of atoms in 8 g of calcium = 8 ÷ 40 = 0.2 mol
(c) Number of moles of atoms in 69 g of sodium = 69 ÷ 23 = 3 mol

3. **(a)** Mass of 3 mol of helium atoms: 3 × 4 = 12 g
(b) Mass of 0.5 mol of nitrogen atoms: 0.5 × 14 = 7 g
(c) Mass of 4 mol of phosphorus atoms: 4 × 31 = 124 g

4. **(a)** Number of moles of SO_2 in 32 g = 32 ÷ (32 + 16 + 16)
= 0.5 mol

(b) Number of moles of H_2SO_4 in 4.9 g = 4.9 ÷ (1 + 1+ 32 + 4 × 16)
= 0.05 mol

(c) Number of moles of $CaCO_3$ in 20 g = 20 ÷ (40 + 12 + 3 × 16)
= 0.2 mol

5. **(a)** Mass of 0.5 mol of H_2O = 0.5 × (1 + 1 + 16)
= 9 g

(b) Mass of 2 mol of Na*Cl* = 2 × (23 + 35.5)
= 117 g

(c) Mass of 4 mol of O_2 = 4 × (16 + 16)
= 128 g

6. **(a)** Number of moles of C atoms in 12 g = 12 ÷ 12 = 1 mol
1 mole of Mg contains the same number of atoms as 1 mole of C.
Mass of 1 mole of Mg = 1 × 24 = 24 g

(b) Number of moles of N_2 molecules in 56 g = 56 ÷ 28 = 2 mol
2 moles of O_2 contains the same number of atoms as 2 moles of N_2.
Mass of 2 moles of O_2 = 2 × 32 = 64 g

(c) Number of moles of H_2O molecules in 36 g = 36 ÷ 18 = 2 mol
2 moles of Cl_2 contains the same number of atoms as 2 moles of H_2O.
Mass of 2 moles of Cl_2 = 2 × 71 = 142 g

(page 134)

Mystery Clue

(a) Mass of Fe_2O_3 = 32 000 g or 32 kg.

(b) Mass of Fe = 22 400 g or 22.4 kg.

Calculations with a spreadsheet

This activity provides an introduction to the use of spreadsheets for chemical calculations. Students who are familiar with spreadsheets in Excel will have no difficulty with this exercise. For students who are less familiar with this program, teacher guidance will be necessary and may include formatting details in the spreadsheet such as changing the margin size or centralising text in the columns. To check the answers obtained using the spreadsheet, get students to compare the answers obtained with and without the spreadsheet.

Here is a suggestion on how you can easily check the work obtained by students using the spreadsheet:

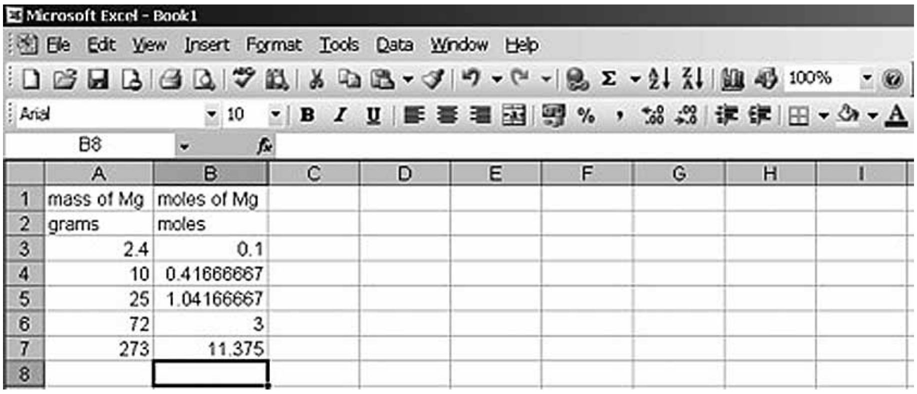
(a) Prepare an overhead transparency with a printout of the answers in the spreadsheet.

(b) Get students to print out their spreadsheets and hand them in.

(c) Compare the answers on the overhead transparency with those obtained by the students; the data should match up.

Answers:
(a) 0.1 mol
(b) 0.42 mol
(c) 1.04 mol
(d) 3 mol
(e) 11.38 mol

Calculations with a spreadsheet are useful if a large number of repetitive calculations need to be carried out.



	A	B
1	mass of Mg	moles of Mg
2	grams	moles
3	2.4	0.1
4	10	0.41666667
5	25	1.04166667
6	72	3
7	273	11.375
8		

Teaching pointers

9.4 What are the Different Kinds of Chemical Formulae? (page 135)

1. Ethane C_2H_6 is a good example to use for teaching the three kinds of chemical formulae. You may use *ball-and-stick* models and diagrams of ethene.

2. Another example (though without models) for molecular and empirical formula is glucose $C_{12}H_{22}O_{11}$.

3. Discuss how to calculate the empirical formula of a compound. The key step is in changing the masses of the elements into moles. (The ratio of elements in the formula is a mole ratio and not a mass ratio.) Some students may have difficulty appreciating this concept. If students do not understand this point, ensure that they can do the calculations.

4. Point out that in conventional experiments on chemical formulae, it is just the empirical formula that is obtained. To obtain the chemical formulae, relative molecular masses must be determined in other experiments (that are not studied in the 'O' Level course).

5. In Experiment 9.1 of the Practical Workbook, the empirical formula for magnesium oxide is determined. However, the results may not be always accurate. Ensure that students appreciate how errors may arise. They are mainly due to:

   (a) some of the magnesium ribbon not burning, and
   (b) some of the magnesium oxide escaping from the crucible as smoke.

6. The work on *structural formulae*, introduced in this section, is developed later in Section 7 on the study of organic compounds.

Skills Practice (page 138)

1. (a)

	Nitrogen (N)	Oxygen (O)
Step 1: Write down the mass of each element.	3.5 g	8.0 g
Step 2: Write down the molar mass of each element.	14 g	16 g
Step 3: Divide each mass by its molar mass to obtain the number of moles.	3.5 ÷ 14 = 0.25	8.0 ÷ 16 = 0.5
Step 4: Divide each number of moles by the smallest number.	0.25 ÷ 0.25 = 1	0.5 ÷ 0.25 = 2

Therefore, the empirical formula is NO_2.

(b)

	Iron (Fe)	Oxygen (O)
Step 1: Write down the mass of each element.	2.8 g	1.2 g
Step 2: Write down the molar mass of each element.	56 g	16 g
Step 3: Divide each mass by its molar mass to obtain the number of moles.	2.8 ÷ 56 = 0.05	1.2 ÷ 16 = 0.075
Step 4: Divide each number of moles by the smallest number.	0.05 ÷ 0.075 = 0.667	0.075 ÷ 0.075 = 1
Ratio	2	3

Therefore, the empirical formula is Fe_2O_3.

(c)

	Carbon (C)	Oxygen (O)	Hydrogen (H)
Step 1: Write down the mass of each element.	2.4 g	6.4 g	0.2 g
Step 2: Write down the molar mass of each element.	12 g	16 g	1 g
Step 3: Divide each mass by its molar mass to obtain the number of moles.	2.4 ÷ 12 = 0.2	6.4 ÷ 16 = 0.4	0.2 ÷ 1 = 0.2 g
Step 4: Divide each number of moles by the smallest number.	0.2 ÷ 0.2 = 1	0.4 ÷ 0.2 = 2	0.2 ÷ 0.2 = 1

Therefore, the empirical formula is CO_2H.

(d)

	Sodium (Na)	Sulphur (S)	Oxygen (O)
Step 1: Write down the mass of each element.	9.2 g	12.8 g	9.6 g
Step 2: Write down the molar mass of each element.	23 g	32 g	16 g
Step 3: Divide each mass by its molar mass to obtain the number of moles.	9.2 ÷ 23 = 0.4	12.8 ÷ 32 = 0.4	9.6 ÷ 16 = 0.6
Step 4: Divide each number of moles by the smallest number.	0.4 ÷ 0.4 = 1	0.4 ÷ 0.4 = 1	0.6 ÷ 0.4 = 1.5
Ratio	2	2	3

Therefore, the empirical formula is $Na_2S_2O_3$.

2. (a)

	Carbon (C)	Hydrogen (H)
Step 1: Write down the mass of each element in 100 g of the compound (i.e. the percentage)	75%	25%
Step 2: Write down the molar mass of each element.	12 g	1 g
Step 3: Divide each mass (i.e. the percentage) by its molar mass to obtain the number of moles.	75 ÷ 12 = 6.25	25 ÷ 1 = 25
Step 4: Divide each number of moles by the smallest number.	6.25 ÷ 6.25 = 1	25 ÷ 6.25 = 4

Therefore, the empirical formula is CH_4.

(b)

	Silicon (Si)	Oxygen (O)
Step 1: Write down the mass of each element in 100 g of the compound (i.e. the percentage)	46.7%	53.3%
Step 2: Write down the molar mass of each element.	28 g	16 g
Step 3: Divide each mass (i.e. the percentage) by its molar mass to obtain the number of moles.	46.7 ÷ 28 = 1.67	53.3 ÷ 16 = 3.33
Step 4: Divide each number of moles by the smallest number.	1.67 ÷ 1.67 = 1	3.33 ÷ 1.67 = 2

Therefore, the empirical formula is SiO_2.

(c)

	Sodium (Na)	Carbon (C)	Oxygen (O)
Step 1: Write down the mass of each element in 100 g of the compound (i.e. the percentage)	43.4%	11.3%	45.3%
Step 2: Write down the molar mass of each element.	23 g	12 g	16 g
Step 3: Divide each mass by its molar mass to obtain the number of moles.	43.4 ÷ 23 = 1.89	11.3 ÷ 12 = 0.94	45.3 ÷ 16 = 2.83
Step 4: Divide each number of moles by the smallest number.	1.89 ÷ 0.94 = 2	0.94 ÷ 0.94 = 1	2.83 ÷ 0.94 = 3

Therefore, the empirical formula is Na_2CO_3.

(d)

	Carbon (C)	Hydrogen (H)	Oxygen (O)
Step 1: Write down the mass of each element in 100 g of the compound (i.e. the percentage)	48.6%	8.1%	43.2%
Step 2: Write down the molar mass of each element.	12 g	1 g	16 g
Step 3: Divide each mass by its molar mass to obtain the number of moles.	48.6 ÷ 12 = 4.05	8.1 ÷ 1 = 8.1	43.2 ÷ 16 = 2.7
Step 4: Divide each number of moles by the smallest number.	4.05 ÷ 2.7 = 1.5	8.1 ÷ 2.7 = 3	2.7 ÷ 2.7 = 1
Ratio	3	6	2

Therefore, the empirical formula is $C_3H_6CO_2$.

3. (a)

	Chlorine (C*l*)	Oxygen (O)
Step 1: Write down the mass of each element.	17.4 – 3.2 = 14.2 g	3.2 g
Step 2: Write down the molar mass of each element.	35.5 g	16 g
Step 3: Divide each mass by its molar mass to obtain the number of moles.	14.2 ÷ 35.5 = 0.4	3.2 ÷ 16 = 0.2
Step 4: Divide each number of moles by the smallest number.	0.4 ÷ 0.2 = 2	0.2 ÷ 0.2 = 1

Therefore, the empirical formula is Cl_2O.

(b)

	Chlorine (C*l*)	Oxygen (O)
Step 1: Write down the mass of each element.	66.8 – 16.0 = 50.8 g	16.0 g
Step 2: Write down the molar mass of each element.	127 g	16 g
Step 3: Divide each mass by its molar mass to obtain the number of moles.	50.8 ÷ 127 = 0.4	16 ÷ 16 = 1
Step 4: Divide each number of moles by the smallest number.	0.4 ÷ 0.4 = 1	1 ÷ 0.4 = 2.5
Ratio	2	5

Therefore, the empirical formula is I_2O_5.

4. Let the molecular formula be $(COH_3)_n$.
The relative molecular mass $= (12 + 16 + 3)n$
$= 31n$

$31n = 62$
$n = 62 \div 31$
$= 2$

Thus, the molecular formula is $C_2O_2H_6$.

5. Let the molecular formula be $(H_2CO)_n$.
The relative molecular mass $= (2 + 12 + 16)n$
$= 30n$

$30n = 90$
$n = 90 \div 30$
$= 3$

Thus, the molecular formula is $H_6C_3O_3$.

6. Let the molecular formula be $(H_2CO_2)_n$.
The relative molecular mass $= (2 + 12 + 2 \times 16)n$
$= 46n$

$46n = 46$
$n = 46 \div 46$
$= 1$

Thus, the molecular formula is H_2CO_2.

Notes for Teachers

Notes on chemical formulae

- Chemical formulae show the composition of elements in compounds in terms of mole ratios and *not* mass ratios.
- Structural formulae are used for covalent elements and compounds but not ionic compounds.
- For compounds with giant structures, such as sodium chloride Na*Cl* and silicon dioxide SiO_2, it is meaningless to talk about molecules and molecular formulae. In these cases, *empirical* formula is used.

Teaching pointers

9.5 What is the Molar Volume of Gases? (page 138)

1. Revise the idea of a law (see page xi of the Textbook). Until a few years ago, Avogadro's Law was known as Avogadro's *hypothesis* as there was no direct experimental evidence for it.
2. Point out to students again that it was Avogadro why first suggested the idea of the molecule as a particle distinct from an atom.
3. Molecules of different gases have different sizes. As there are large spaces between the molecules, the size does not affect the overall volume of the gas. Hence, different gases with the same number of molecules have the same volume.
4. Molar volume, like molar mass, is not a relative measure and therefore has a unit. Get students to accept the value for the molar volume of a gas as there is no evidence provided in this course for it. To enable students to have a concrete idea of the volume of a mole of gas as 24 dm^3, show them a (cardboard) cube of side 29 cm.
5. Ensure students appreciate that molar volume and molar mass are different. Compare the two formulae below for calculating the numbers of moles using the molar volume and molar mass respectively.

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

$$\text{Number of moles} = \frac{\text{Mass (in g)}}{\text{Molar mass (in g)}}$$

Skills Practice (page 141)

1. (a) Yes. According to Avogadro's law.
   (b) No, as the mass of a He atom/molecule is less than the mass of a CO_2 molecule.
2. (a) Mass of 12 dm^3 of $H_2 = \frac{12}{24} \times 2$
   $= 1$ g
   (b) Mass of 4.8 dm^3 of $CO_2 = \frac{4.8}{24} \times 44$
   $= 8.8$ g
   (c) Mass of 6 dm^3 of $NH_3 = \frac{6}{24} \times 17$
   $= 4.25$ g
3. (a) Volume of 14 g of $N_2 = \frac{14}{28} \times 24$
   $= 12$ dm^3
   (b) Volume of 2 g of $H_2 = \frac{2}{2} \times 24$
   $= 24$ dm^3
   (c) Volume of 3.2 g of $SO_3 = \frac{3.2}{80} \times 24$
   $= 0.96$ dm^3
4. (a) Number of moles of CO in 6 dm^3 $= 6 \div 24$
   $= 0.25$ mol
   (b) 0.25*x* molecules
   (c) Mass of sample $= 0.25 \times (12 + 16)$
   $= 7$ g

Notes for Teachers

Avogadro's Number

In recognition of the idea expressed in Avogadro's Law, the number of fundamental particles in a mole of substance (6×10^{23}) is called the **Avogadro Number** (or **Avogadro's Number**). Although it was not Avogadro who came up with this value, his hypothesis did lead to the eventual determination of this number.

Teaching pointers

9.6 Why Does the Concentration of Solutions Matter in Chemistry? (page 141)

1. You may use orange juice to demonstrate the idea of concentration. It is advisable to use a powdered orange drink that readily dissolves in water to give clear solutions.
2. Introduce the idea of concentration using coloured solutions of potassium dichromate(VI) or powdered orange drink. Refer to 'Notes for Teachers' below on how to do this. The worksheet (Additional Exercise 2) given at the end of this chapter may be useful for this task. It can be photocopied and distributed to the class.
3. When discussing Figure 9.10, also get the class to compare the concentrations of Solution A and Solution C. Answer: The volume of Solution B is the same as the volume of Solution B but has twice the amount of solute (2 spoonfuls). Thus the concentration of Solution B is twice the concentration of Solution A.
4. Following the discussion of concentration with coloured solutions, the concentration of *colourless* solutions can be discussed. Show the class some bottles of acids and alkalis and explain the need for the concentrations of these solutions to be shown on the bottles.
5. As some students may be confused over the terms *concentration* and *density*, discuss the differences between the two.
6. The spreadsheet calculations in Additional Exercise 9.3 at the end of this chapter can help to further develop the IT skills of students. You should provide some guidance on the formatting details in the spreadsheet such as changing the margin size or centralising of text in the columns.

Chemistry Inquiry (page 142)

What is the Concentration of the Solution?

Again, write the question on the board, get the class to close their textbooks and in groups to try to work out the steps needed to solve the problem.

Group Discussion

1. Concentration involves amount of solute. Amount can be measured as mass of solute or moles moles of solute (number of solute particles) to give concentration as mass concentration and molar concentration respectively.
2. Mass concentration.
3. The SI unit for amount is the mole (mol) and for volume is the cubic metre (m^3). Thus, the SI unit for concentration is mole per cubic metre, symbol mol/m^3.

Skills Practice (page 144)

1. (a) Concentration in mol/dm^3 $= \frac{0.4}{2}$
$= 0.2$ mol/dm^3
Concentration in g/dm^3 $= 0.2 \times 98 = 19.6$ g/dm^3
(b) Concentration in mol/dm^3 $= \frac{0.5}{250} \times 1000$
$= 2.0$ mol/dm^3
Concentration in g/dm^3 $= 2.0 \times 36.5$
$= 73$ g/dm^3
(c) Concentration in mol/dm^3 $= \frac{0.01}{50} \times 1000$
$= 0.2$ mol/dm^3
Concentration in g/dm^3 $= 0.2 \times (23 + 16 + 1)$
$= 8$ g/dm^3
2. (a) Number of grams of H_2SO_4 in 1 dm^3 $= \frac{9.8}{250} \times 1000$
$= 39.2$ g
(b) Concentration in mol/dm^3 $= \frac{39.2}{98}$
$= 0.4$ mol/dm^3
3. (a) Mass of sulfuric acid in 1 dm^3 $= 1.83 \times 1000 \times 0.94$
$= 1720$ g
(b) (i) Concentration in g/dm^3 $= 1720$ g/dm^3
(ii) Concentration in mol/dm^3 $= \frac{1720}{98}$
$= 17.6$ mol/dm^3
4. (a) Number of moles of $HCl = 4 \times 0.5$
$= 2$ mol
(b) Number of moles of NaOH $= \frac{0.4}{1000} \times 25$
$= 0.01$ mol
5. Number of moles of KOH in 500 cm^3 $= \frac{0.4}{1000} \times 500$
$= 0.2$ g
Mass of KOH in 500 cm^3 $= 0.2 \times 56$
$= 11.2$ g

Notes for Teachers

Teaching the concept of concentration

Research with Chemistry students has found that they have a good *intuitive* idea of concentration when applied to everyday coloured solutions (e.g. concentrated and dilute orange juice). However, they have great difficulty manipulating the two variables of amount of solute and volume (especially the latter) to account for the differences in concentration of solutions.

To introduce concentration, *coloured* solutions (e.g. potassium dichromate(VI) solution or powdered orange drink) should be used as differences in concentration are readily observable and intuitively understood due to the differences in colour .

Show coloured solutions of differing concentration such as those in Figure 9.10 on page 141 of the Textbook. Discuss how the concentration of a solution depends on:

(a) the amount of solute, and

(b) the volume of the solution.

Research has also shown that students are able to understand the concept of concentration better when these two factors are manipulated *separately* rather than introducing them together as in a formula. For Figure 9.10:

- Comparing Solutions A and B shows how the amount of solute affects the concentration. This idea is more easily understood as the volumes are kept the same.
- Comparing Solutions B and C shows the relationship between volume and concentration, i.e. as the volume increases (for the same amount of solute), the concentration decreases.
- Comparing Solutions A and C involves the manipulation of *both* the amount of solute and the volume of solutions.

When teaching students to do calculations involving concentration, get them to generate and describe the steps of the solutions in a *qualitative* way first in order to understand what they are doing. Then carry out the numerical calculations. Research has shown that this exercise helps students to develop better conceptual understanding and leads to improved problem-solving abilities.

09 Chapter Review

Self-Management

Misconception Analysis (page 144–145)

1. **False** The symbol for the mole is 'mol'. The symbol 'm' is for the metre.

2. **False** The unit for molar mass is the gram. Unlike relative atomic mass or relative molecular mass, molar mass is not relative and therefore must have a unit.

3. **True** The mole is just a number and can be used to count any objects. It is normally used only in Chemistry to count the large numbers of particles involved.

4. **False** In some compounds, such as CuO, the empirical formula is the same as the molecular formula. **Note:** The molecular formula can never be simpler than the empirical formula.

5. **False** According to Avogadro's Law, these gases contain equal numbers of *molecules*.

6. **False** The molar gas volume is 24 dm^3 only at room temperature and pressure (25 °C and 1 atmosphere).

7. **False** One mole of all gases has the same volume (at the same temperature and pressure) but *not* the same mass. For example, 1 mole of H_2 has a mass of 2 g whereas 1 mole of O_2 has a mass of 32 g.

8. **True** Concentration compares amount of solute (with any appropriate unit such as grams or moles) to volume of solution (also with any appropriate unit such as cm^3 or dm^3).

Practice

Structured Questions (pages 145–146)

1. (a) $MgCl_2$
   (b) Relative formula mass of $MgCl_2$ = 24 + 35.5 + 35.5
   = 95
   (c) Number of moles of $MgCl_2$ in 26.7 g of the compound
   = 26.7 ÷ 95
   = 0.28 mol
   (d) 1 mol of $MgCl_2$ has 2 mol of Cl^- ions
   0.5 mol of $MgCl_2$ has 1 mol of Cl^- ions

2. (a) Mass of hydrogen in the compound = 64 – 56 = 8 g
   (b)

	Nitrogen (N)	Hydrogen (H)
Step 1: Write down the mass of each element.	56 g	8 g
Step 2: Write down the molar mass of each element.	14 g	1 g
Step 3: Divide each mass by its molar mass to obtain the number of moles.	56 ÷ 14 = 4	8 ÷ 1 = 8
Step 4: Divide each number of moles by the smallest number.	4 ÷ 4 = 1	8 ÷ 4 = 2

   Therefore, the empirical formula is NH_2.

   (c) Let the molecular formula be $(NH_2)_n$.
   The relative molecular mass = $(14 + 2)n$
   = $16n$
   $16n = 32$
   $n = 32 \div 16$
   $= 2$
   Thus, the molecular formula is N_2H_4.

   (d) The compound is likely to have a low melting and boiling point, be volatile, be insoluble in water and not to conduct electricity (any two properties).

3. (a) Mass of zinc = 153.27 – 150.00 = 3.27 g
   Mass of product = 154.07 – 150.00 = 4.07 g
   Mass of oxygen in product = 4.07 – 3.27 = 0.80 g
   (b)

	Zinc (Zn)	Oxygen (O)
Step 1: Write down the mass of each element.	3.27 g	0.8 g
Step 2: Write down the molar mass of each element.	65 g	16 g
Step 3: Divide each mass by its molar mass to obtain the number of moles.	3.27 ÷ 65 = 0.05	0.8 ÷ 16 = 0.05
Step 4: Divide each number of moles by the smallest number.	0.05 ÷ 0.05 = 1	0.05 ÷ 0.05 = 1

   Therefore, the empirical formula is ZnO.

4. (a) Connect lithium with a bulb and electric cell to see if it conducts electricity. All metals conduct electricity.

(b) (i)

	Lithium (Li)	Sulfur (S)	Oxygen (O)
Step 1: Write down the mass of each element in 100 g of the compound (i.e. the percentage)	8.0%	36.8%	55.2%
Step 2: Write down the molar mass of each element.	7 g	32 g	16 g
Step 3: Divide each mass (i.e. the percentage) by its molar mass to obtain the number of moles.	8.0 ÷ 7 = 1.14	36.8 ÷ 32 = 1.15	55.2 ÷ 16 = 3.45
Step 4: Divide each number of moles by the smallest number.	1.14 ÷ 1.14 = 1	1.15 ÷ 1.14 = 1	3.45 ÷ 1.14 = 3

Therefore, the empirical formula is $LiSO_3$.

(ii) Let the molecular formula be $(LiSO_3)_n$.

The relative molecular mass = $(7 + 32 + 16 + 16 + 16)n$
= $87n$

$87n = 174$
$n = 174 \div 87$
$= 2$

Thus, the molecular formula is $Li_2S_2O_6$.

(c) This compound is likely to have a high melting point and to conduct electricity when molten.

5. (a) Heat the solid again and see if the mass decreases further. The decomposition is complete when there is no further decrease in mass.

(b) (i) Mass of water produced in the decomposition
= 17.22 – 13.62 = 3.6 g

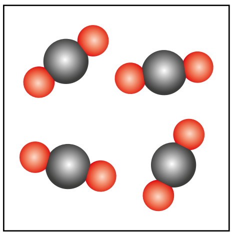
(ii) Moles of water produced = 3.6 ÷ 18 = 0.2 mol

(c) (i) Mass of anhydrous sodium carbonate produced in the decomposition = 13.62 – 11.50 = 2.12 g

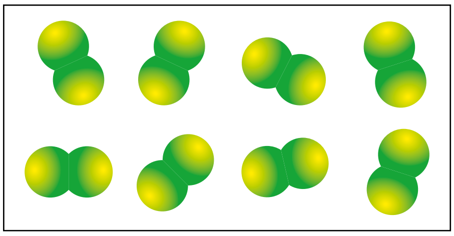
(ii) Number of moles of Na_2CO_3 produced
= 2.12 ÷ (23 + 23 + 12 + 3 × 16)
= 0.02 mol

(d) 0.02 moles of Na_2CO_3 produce 0.2 moles of water.
1 mole of Na_2CO_3 produces 10 moles of water
Therefore, $n = 10$

6. (a)



(b)



7. (a) Relative molecular mass of chlorine = 35.5 + 35.5 = 71

(b) Number of moles of chlorine = 168 ÷ 24 000 = 0.007 mol

(c) Mass of chlorine gas in the sample = 0.007 × 71 = 0.5 g

8. (a) Relative formula mass of $NaCl$ = 23 + 35.5 = 58.5

(b) Moles of $NaCl$ in 500 cm^3 of seawater
= 12.4 ÷ 58.5 = 0.212 mol

(c) Number of $NaCl$ units in 500 cm^3 of seawater
= $0.212 \times 6.02 \times 10^{23}$
= 1.27×10^{23} units

(d) Concentration of the solution in mol/dm^3
$= \frac{0.212}{500} \times 1000$
= 0.424 mol/dm^3

9. (a) Relative molecular mass of ethanol
= 12 + 12 + 6 + 16 = 46

(b) Concentration of the solution in g/dm^3
$= \frac{12.5}{200} \times 1000$
= 62.5 g/dm^3

(c) Concentration of the solution in mol/dm^3
$= \frac{62.5}{46}$
= 1.36 mol/dm^3

Free Response Questions (page 146)

1. The following inferences could be made:
   - Molar mass of chlorine = 71 g. Molar mass of carbon dioxide = 44 g.
   - Number of moles of chlorine = 71 ÷ 71 = 1 mol
   - Number of moles of carbon dioxide = 44 ÷ 44 = 1 mol
   - Both gases contain the same number of molecules. Number of molecules = 6×10^{23} molecules
   - From Avogadro's Law, both gases have the same volume.
   - From the data given, the actual volumes of the gases cannot be determined.

2. Empirical formula of a compound shows the simplest whole number ratio of the atoms present.

   For a description of an experiment that could be carried out to determine the empirical formula of magnesium oxide, refer to the steps listed in the procedure for Experiment 9.1 on pages 61–62 of the Practical Workbook.

   Possible sources of error include:
   - Some of the magnesium had not burnt.
   - Some of the magnesium oxide escaped as smoke.
   - Weighing was done before the crucible and its content were cooled to room temperature.

Extension (page 146)

A Mole of Coins
The mole (6×10^{23}) is such a large number that it is very difficult for us to appreciate its value. One way of visualising the value of a mole is to use the analogy in this activity. The aim of this activity is for students to realise:
- how large the number of atoms in one mole is.
- that it is not sensible to use moles for ordinary objects such as coins.

Answers:
(a) Ten coins have a thickness of 1.4 cm.
So, 1 mole (6×10^{23}) of coins is 8×10^{22} cm or 800 000 000 000 000 000 km high.
(b) Approximate distances are:
Woodlands to Sentosa: 30 km
Singapore to London: 10 000 km
Earth to the Sun: 150 000 000 km
Earth to the nearest star outside the Solar System (Alpha Centauri): 40 000 000 000 000 km
(c) The pile of coins would stretch from Earth to the Sun and back 2 000 million times or from the Earth to Alpha Centauri and back 10 000 times.

Note: As such distances are outside our experience, it is impossible for us to fully grasp the magnitude of the number.
(d) No, it is not sensible to use moles for coins because 1 mole of coins simply does not exist.

A bag of coins from the bank, for example, contain only a tiny fraction of a mole. The mole is too large a unit for counting everyday objects.

As moles are used to count chemical particles, the magnitude of the mole gives us an impression of the number of particles involved in, for example a chemical reaction. Also, as the counting of everyday objects such as coins and apples involves relatively small numbers, the mole is not a sensible unit for counting them.

Additional Teaching Material

Additional Exercise 1: Using a Spreadsheet to Calculate Number of Moles

Objective

- To calculate the number of moles of various masses of a substance

Key Competency

ICS: management of information [*using spreadsheet software*]

Computer spreadsheets are a convenient way of processing data in numerical problems or data obtained from experiments. Spreadsheets save a lot of time and effort compared with manual calculations. A spreadsheet is very useful for carrying out repetitive calculations. The example on this page shows the way to use a spreadsheet to calculate numbers of moles using the formula:

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

Note: There are several different types of spreadsheets. The comments here apply to *Microsoft Excel*.

Use of the spreadsheet

Suppose we want to use a spreadsheet to calculate the number of moles of magnesium atoms in 2.4 g, 10 g, 25 g, 72 g and 273 g of magnesium.

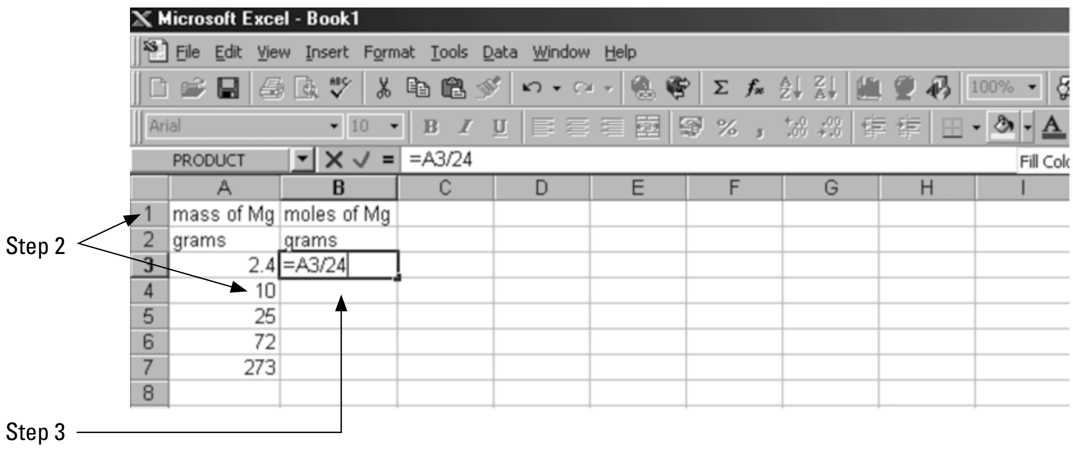
1. Open the spreadsheet. You will see a page similar to the one below (though it will be a *blank* sheet).

2. The masses of magnesium are entered into Column A of the spreadsheet.

3. To calculate the number of moles, we use the formula:

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

This must be changed into a **spreadsheet formula**. In this case, it is '=A3/24' and it is entered into Cell B3, i.e., row 3 of column B of the spreadsheet. (A3 = cell A3; 24 = molar mass of magnesium; / is the symbol for division.) When you leave cell B3, the answer for the number of moles of magnesium (0.1) appears in this cell.

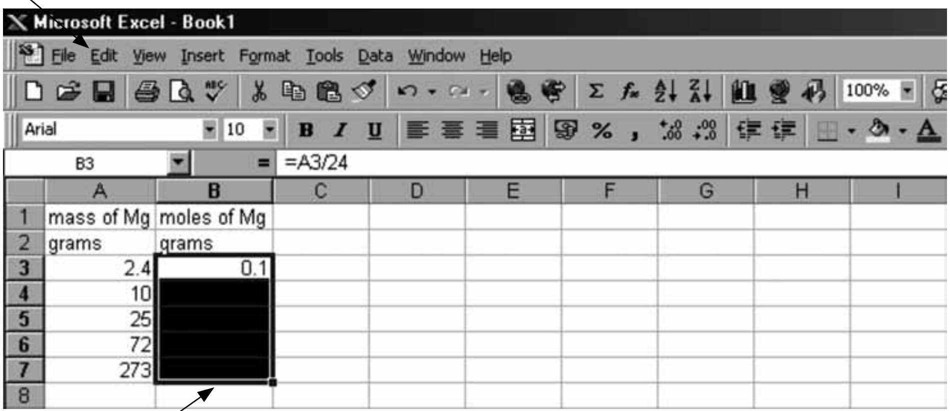


4. The calculations for *all* the moles of magnesium can now be done in *one* step.

(a) Hold down the left mouse key in cell B3 to highlight the cell. Then drag the mouse down

(b) Click 'Edit' and go to 'Fill'. You are then given several choices. Click on 'Down'. The answers for the other calculations wikk appear in the highlighted area.

Step 4(b): Click on 'Edit'. Go to 'Fill' and click 'Down'.



Step 4(a): Select these cells.

Additional Teaching Material

Additional Exercise 2: What Factors Affect the Concentration of a Solution?

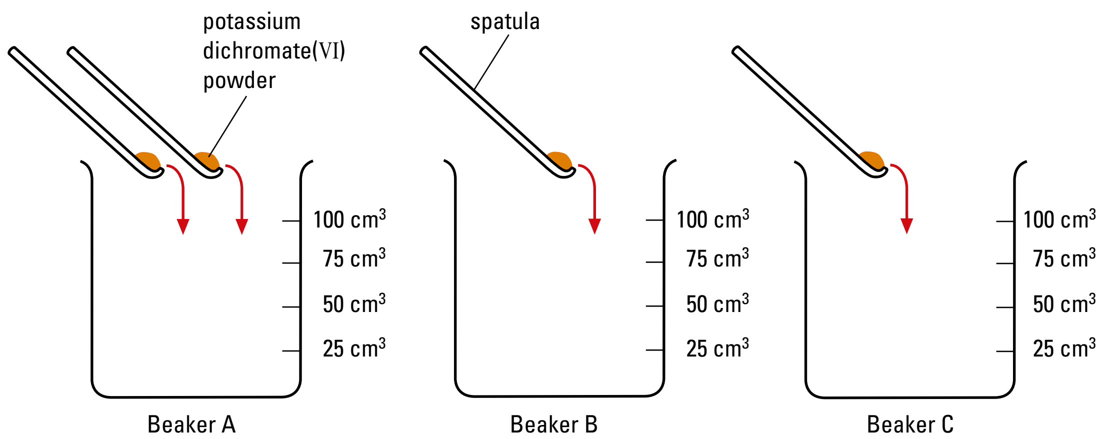
Objective

- To investigate the factors involved in concentration

Key Competency

CIT: sound reasoning [*observing, comparing, manipulating variables*]

Observe as your teacher carries out the experiment below.



1. Add 50 cm^3 of water to Beakers A and B. Then dissolve potassium dichromate(VI) powder in the two beakers as follows:

Beaker A: 2 spatula (or spoon) full
Beaker B: 1 spatulas (or spoons) full

(a) Which solution is more concentrated? Explain.

____

____

(b) Shade the diagrams of Solutions A and B to show the differences in concentration.

2. Now add 25 cm^3 of water to Beaker C. Then add 1 spatula (or spoon) full of potassium dichromate(VI) powder.

(a) Which solution is more concentrated — B or C? Explain.

____

____

(b) Shade the diagram of the solution in Beaker C to show the difference in concentration.

3. Look at Beakers A and C.
Compare the concentrations of the two solutions.

____

____

4. The concentrations of acid and alkali solutions are more difficult to compare than that of potassium dichromate(VI) solutions.

(a) What is the reason for this?

____

____

(b) What do chemists do to ensure that they know the correct concentration of an acid or alkali that is in a bottle of reagent?

____

____

Conclusions

Complete the sentences in the box below.

- The concentration of a solution depends on two factors:
  - The amount of ____________________.
  - The ______________________ of the solution.
- Units commonly used in Chemistry for:
  - The amount of solute are the _________________ and the _________________.
  - The volume are ______________________ and ______________________.
  - The concentration are ______________________ and ______________________.

Additional Teaching Material

Additional Exercise 3: What Factors Affect the Concentration of a Solution?

Objective

- To calculate the concentrations of various solutions

Key Competency

ICS: management of information [*using spreadsheet software*]

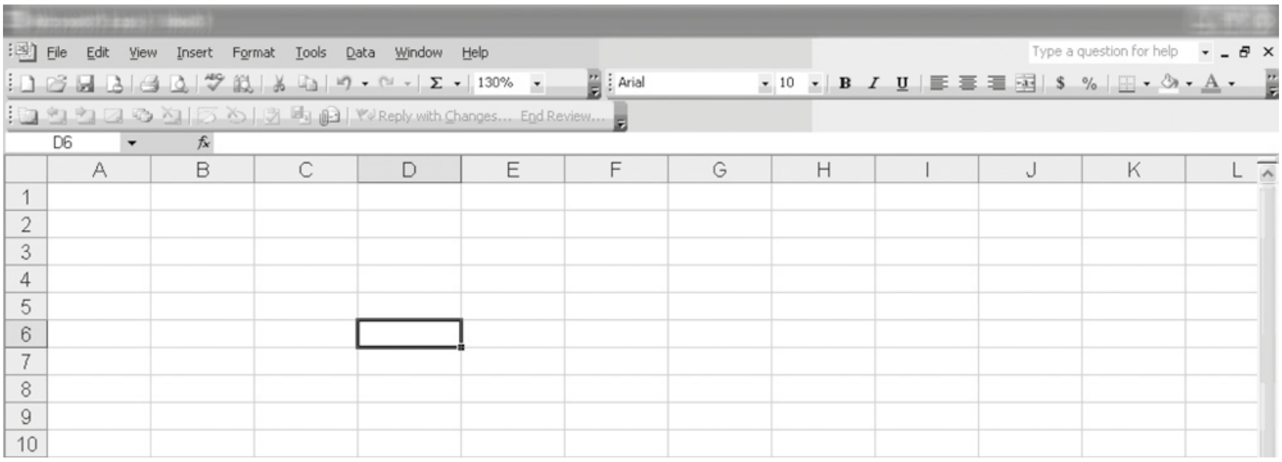
In this exercise, you will use a spreadsheet to calculate the concentrations of some solutions. The instructions enable you to do the work on a computer.

You will use the spreadsheet to calculate the concentrations of the following four sodium hydroxide solutions in (i) g/dm^3 and (ii) mol/dm^3.

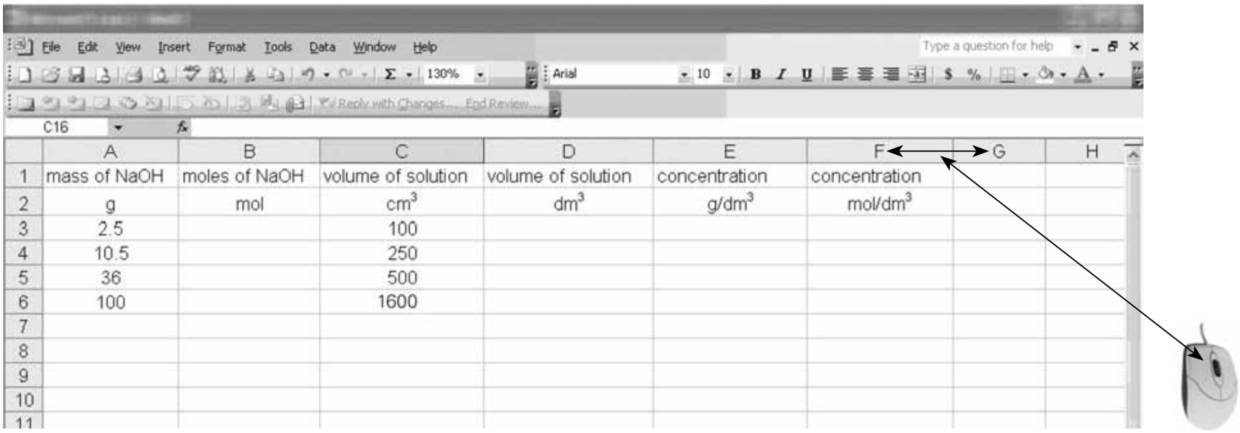
- Solution 1: 2.5 g in 100 cm^3 of solution
- Solution 2: 10.5 g in 250 cm^3 of solution
- Solution 3: 36 g in 500 cm^3 of solution
- Solution 4: 100 g in 1500 cm^3 of solution

The molar mass of sodium hydroxide is 40 g.

1. Open the spreadsheet. You will see a page similar to the following:



2. Enter the data given into the spreadsheet as shown below.



	A	B	C	D	E	F	G	H
1	mass of NaOH	moles of NaOH	volume of solution	volume of solution	concentration	concentration		
2	g	mol	cm^3	dm^3	g/dm^3	mol/dm^3		
3	2.5		100					
4	10.5		250					
5	36		500					
6	100		1600					

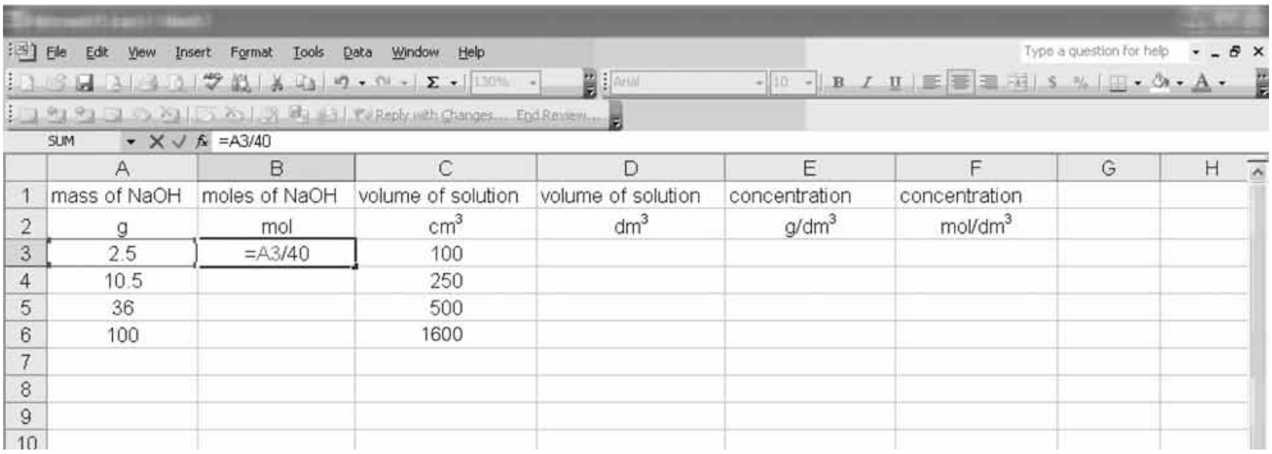
Note: To fit all the words in the columns, you will need to make the columns wider. Do this by placing the cursor on a column dividing line, holding down the left mouse key and dragging the divider to the right.

3. To calculate the concentrations of the sodium hydroxide solutions, you need to use the formulae:
moles = mass (in g) / molar mass (in g)
volume (in dm^3) = volume in cm^3 / 1000
concentration = mass of solute (in g) / volume of solution (in dm^3)
concentration = moles of solute (in mol) / volume of solution (in dm^3)

To do this, enter the following **spreadsheet formulae**:

- Cell B3, i.e., row 3 of column B: =A3/40 (A3 = mass of NaOH in cell A3; 40 = molar mass of NaOH)
- Cell D3, i.e., row 3 of column D: =C3/1000 (C3 = volume of sodium hydroxide in cell C3 in cm^3; divide by 1000 to get volume in dm^3.
- Cell E3, i.e., row 3 of column E: =A3/D3 (A3 = mass of NaOH in cell A3; D3 = volume of sodium hydroxide in cell D3 in dm^3)
- Cell F3, i.e., row 3 of column F: =B3/D3 (B3 = moles of NaOH in cell B3; D3 = volume of sodium hydroxide in cell D3 in dm^3)

The entry in Cell B3 is shown below:



SUM =A3/40

	A	B	C	D	E	F	G	H
1	mass of NaOH	moles of NaOH	volume of solution	volume of solution	concentration	concentration		
2	g	mol	cm^3	dm^3	g/dm^3	mol/dm^3		
3	2.5	=A3/40	100					
4	10.5		250					
5	36		500					
6	100		1600					

When you leave cell B3, notice that the answer for the number of moles of carbon appears in this cell. Now enter the spreadsheet formula for volume of sodium hydroxide into cell D3.

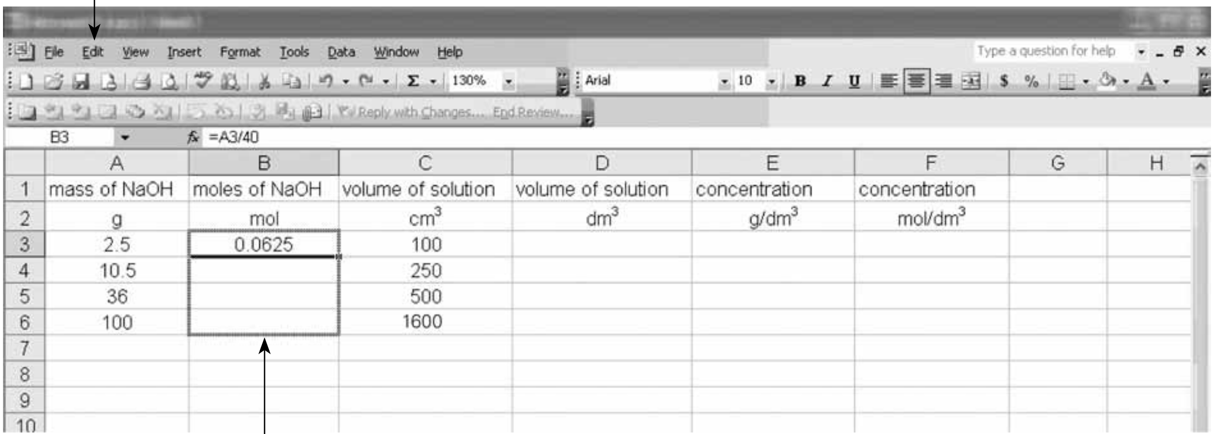
Then do the same for cells E3 and F3.

4. You can now do other calculations in one step.

(a) Hold down the left mouse key in cell B3 to highlight the cell. Then drag the mouse down to highlight the cells shown below:

(b) Click 'Edit' and go to 'Fill'. You are then given several choices. Click on 'Down'. The answers for the other calculations will appear.

(4b) Click on 'Edit'. Go to 'Fill' and click 'Down'.



	A	B	C	D	E	F	G	H
1	mass of NaOH	moles of NaOH	volume of solution	volume of solution	concentration	concentration		
2	g	mol	cm^3	dm^3	g/dm^3	mol/dm^3		
3	2.5	0.0625	100					
4	10.5		250					
5	36		500					
6	100		1600					
7								
8								
9								
10								

(4a) Select these 4 cells.

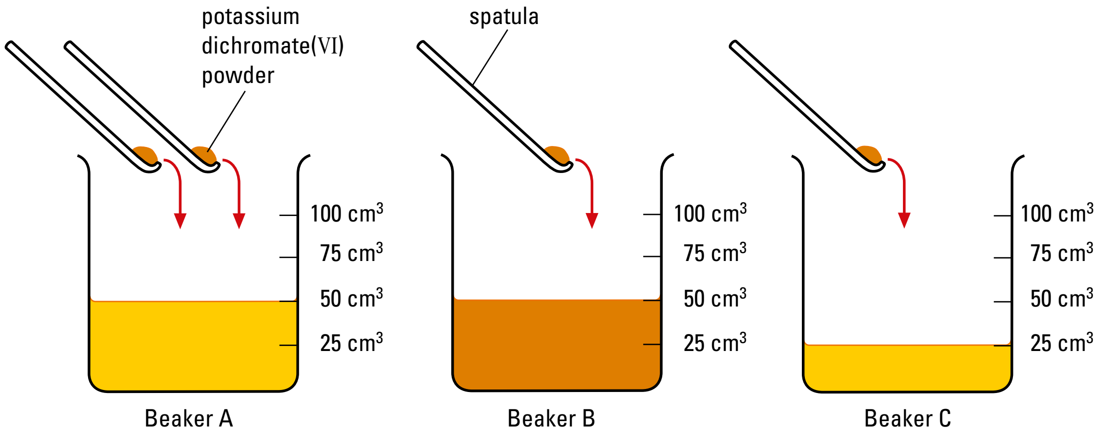
5. Now read from your spreadsheet the concentrations of the solutions:

NaOH solution	Concentration	
• Solution 1: 2.5 g in 100 cm^3 of solution	________ g/dm^3	________ mol/dm^3
• Solution 2: 10.5 g in 250 cm^3 of solution	________ g/dm^3	________ mol/dm^3
• Solution 3: 36 g in 500 cm^3 of solution	________ g/dm^3	________ mol/dm^3
• Solution 4: 100 g in 1500 cm^3 of solution	________ g/dm^3	________ mol/dm^3

(Take the readings correct to 2 decimal places.)

Answers

Additional Exercise 2:



1. **(a)** Solution A is more concentrated as it has more solute than Solution B in the same volume of solution. Solution A is darker in colour.
   (b) See shading above.

2. **(a)** Solution C is more concentrated as it has the same amount of solute as Solution B but only half the volume of solution. Solution C is darker in colour.
   (b) See shading above.

3. Solutions A and C have the same concentration as the amount of solute per unit volume in each solution is the same. They are of the same colour.

Conclusions

Complete the sentences in the box below.

- The concentration of a solution depends on two factors:
  – The amount of ___solute___.
  – The ___volume___ of the solution.
- Units commonly used in Chemistry for:
  – The amount of solute are the ___gram___ and the ___mol___.
  – The volume are ___cm^3___ and ___dm^3___.
  – The concentration are ___g/dm^3___ and ___mol/dm^3___.

Additional Exercise 3:

5. **Concentration**

- ___25___ g/dm^3 ___0.63___ mol/dm^3
- ___42___ g/dm^3 ___1.05___ mol/dm^3
- ___72___ g/dm^3 ___1.80___ mol/dm^3
- ___62.5___ g/dm^3 ___1.56___ mol/dm^3

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