This book provides full coverage of the IB diploma syllabus in Chemistry and offers support to students preparing for their examinations. The book will help you revise the study material, learn the essential terms and concepts, strengthen your problem-solving skills and improve your approach to IB examinations. The book is packed with worked examples and exam tips that demonstrate best practices and warn against common errors. All topics are illustrated by annotated student answers to questions from past examinations, which explain why marks may be scored or missed.

A separate section is dedicated to data-based and practical questions, which are the most distinctive feature of the syllabus (first assessment in 2016). Numerous examples show how to tackle unfamiliar situations, interpret and analyse experimental data, and suggest improvements to experimental procedures. Practice problems and a complete set of IB-style examination papers provide further opportunities to check your knowledge and skills, boost your confidence and monitor the progress of your studies. Full solutions to all problems and examination papers are given online at www.oxfordsecondary.com/ib-prepared-support.

As any study guide, this book is not intended to replace your course materials, such as textbooks, laboratory manuals, past papers and markschemes, the IB Chemistry syllabus and your own notes. To succeed in the examination, you will need to use a broad range of resources, many of which are available online. The authors hope that this book will navigate you through this critical part of your studies, making your preparation for the exam less stressful and more efficient.

Overview of the book structure

The book is divided into several sections that cover the internal assessment, core SL and additional higher level (AHL) topics, data-based and practical questions, the four options (A–D) and a complete set of practice examination papers.

The largest section of the book, core topics, follows the structure of the IB diploma chemistry syllabus (for first assessment 2016) and covers all understandings and applications and skills assessment statements. Topics 1–11 contain common material for SL and HL students while topics 12–21 are intended for HL students only. The nature of science concepts are also discussed where applicable.

The data-based and practical questions section (chapter 22) provides a detailed analysis of problems and laboratory experiments that often appear in section A of paper 3. Similar to core topics, the discussion is illustrated by worked examples and sample scripts, followed by IB-style practice problems. This section also contains a complete list of laboratory experiments and techniques that may be assessed in papers 2 and 3.

The options section reviews the material assessed in the second part of paper 3. Each of the four options is presented as a series of SL and AHL subtopics.

The internal assessment section outlines the nature of the investigation that you will have to carry out and explains how to select a suitable topic, collect and process experimental data, draw conclusions and present your report in a suitable format to satisfy the marking criteria and achieve the highest grade.

The final section contains IB-style practice examination papers 1, 2 and 3, written exclusively for this book. These papers will give you an opportunity to test yourself before the actual exam and at the same time provide additional practice problems for every topic of core and options material.

The answers and solutions to all practice problems and examination papers are given online at www.oxfordsecondary.com/ib-prepared-support. Blank answer sheets for examination papers are also available at the same address.
A list of commonly used command terms in Chemistry examination questions is given in the table below. Understanding the exact meaning of frequently used command terms is essential for your success in the examination. Therefore, you should explore this table and use it regularly as a reference when answering questions in this book.

<table>
<thead>
<tr>
<th>Command term</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>Annotate</td>
<td>Add brief notes to a diagram or graph</td>
</tr>
<tr>
<td>Calculate</td>
<td>Obtain a numerical answer showing your working</td>
</tr>
<tr>
<td>Comment</td>
<td>Give a judgment based on a given statement or result of a calculation</td>
</tr>
<tr>
<td>Compare</td>
<td>Give an account of the similarities between two or more items</td>
</tr>
<tr>
<td>Compare and contrast</td>
<td>Give an account of similarities and differences between two or more items</td>
</tr>
<tr>
<td>Construct</td>
<td>Present information in a diagrammatic or logical form</td>
</tr>
<tr>
<td>Deduce</td>
<td>Reach a conclusion from the information given</td>
</tr>
<tr>
<td>Describe</td>
<td>Give a detailed account</td>
</tr>
<tr>
<td>Determine</td>
<td>Obtain the only possible answer</td>
</tr>
<tr>
<td>Discuss</td>
<td>Offer a considered and balanced review that includes a range of arguments, factors or hypotheses</td>
</tr>
<tr>
<td>Distinguish</td>
<td>Make clear the differences between two or more items</td>
</tr>
<tr>
<td>Draw</td>
<td>Represent by a labelled, accurate diagram or graph, drawn to scale, with plotted points [if appropriate] joined in a straight line or smooth curve</td>
</tr>
</tbody>
</table>

Continued on page VI
<table>
<thead>
<tr>
<th>Command term</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>Estimate</td>
<td>Obtain an approximate value</td>
</tr>
<tr>
<td>Explain</td>
<td>Give a detailed account including reasons or causes</td>
</tr>
<tr>
<td>Formulate</td>
<td>Express precisely and systematically a concept or argument</td>
</tr>
<tr>
<td>Identify</td>
<td>Provide an answer from a number of possibilities</td>
</tr>
<tr>
<td>Justify</td>
<td>Give valid reasons or evidence to support an answer or conclusion</td>
</tr>
<tr>
<td>Label</td>
<td>Add labels to a diagram</td>
</tr>
<tr>
<td>List</td>
<td>Give a sequence of brief answers with no explanation</td>
</tr>
<tr>
<td>Outline</td>
<td>Give a brief account or summary</td>
</tr>
<tr>
<td>Predict</td>
<td>Give an expected result</td>
</tr>
<tr>
<td>Sketch</td>
<td>Represent by means of a diagram or graph [labelled as appropriate], giving a general idea of the required shape or relationship</td>
</tr>
<tr>
<td>State</td>
<td>Give a specific name, value or other brief answer without explanation</td>
</tr>
<tr>
<td>Suggest</td>
<td>Propose a solution, hypothesis or other possible answer</td>
</tr>
</tbody>
</table>

A complete list of command terms is available in the subject guide.

### Preparation and exam strategies

In addition to the above suggestions, there are some simple rules you should follow during your preparation study and the exam itself.

1. **Get ready for study.** Have enough sleep, eat well, drink plenty of water and reduce your stress by positive thinking and physical exercise. A good night’s sleep is particularly important before the exam day, as it can improve your score.

2. **Organize your study environment.** Find a comfortable place with adequate lighting, temperature and ventilation. Avoid distractions. Keep your papers and computer files organized. Bookmark useful online and offline material.

3. **Plan your studies.** Make a list of your tasks and arrange them by importance. Break up large tasks into smaller, easily manageable parts. Create an agenda for your studying time and make sure that you can complete each task before the deadline.

4. **Use this book as your first point of reference.** Work your way through the topics systematically and identify the gaps in your understanding and skills. Spend extra time on the topics where improvement is required. Check your textbook and online resources for more information.

5. **Read actively.** Focus on understanding rather than memorizing. Recite key points and definitions using your own words. Try to solve every worked example and practice problem before looking at the answer. Make notes for future reference.

6. **Get ready for the exams.** Practice answering exam-style questions under a time constraint. Learn how to use the Chemistry data booklet quickly and efficiently. Solve as many problems from past papers as you can. Take a trial exam using the papers at the end of this book.

7. **Optimize your exam approach.** Read all questions carefully, paying extra attention to command terms. Keep your answers as short and clear as possible. Double-check all numerical values and units. Label axes in graphs and annotate diagrams. Use exam tips from this book.

8. **Do not panic.** Take a positive attitude and concentrate on things you can improve. Set realistic goals and work systematically to achieve these goals. Be prepared to reflect on your performance and learn from your errors in order to improve your future results.
Key features of the book

Each chapter typically covers one core or option topic, and starts with “You should know” and “You should be able to” checklists. These outline the understandings and applications and skills sections of the IB diploma Chemistry syllabus. Some assessment statements have been reworded or combined together to make them more accessible and simplify the navigation. These changes do not affect the coverage of key syllabus material, which is always explained within the chapter. Chapters contain the features outlined on this page.

Theoretical concepts and key definitions are discussed at a level sufficient for answering typical examination questions. Many concepts are illustrated by diagrams, tables or worked examples. Most definitions are given in a grey side box like this one, and explained in the text.

Assessment tip

This feature highlights the essential terms and statements that have appeared in past mark schemes, warns against common errors and shows how to optimize your approach to particular questions.

Nature of science relates a chemistry concept to the overarching principles of the scientific approach.

Approaches to learning gives advice on the development of communication, social, self-management, research or thinking skills.

Sample student answers show typical student responses to IB-style questions (most of which are taken from past examination papers). In each response, the correct points are often highlighted in green while incorrect or incomplete answers are highlighted in red. Positive or negative feedback on student’s response is given in the green and red pull-out boxes. An example is given below.

Practice problems are given at the end of each chapter. These are IB-style questions that provide you with an opportunity to test themselves and improve your problem-solving skills. Some questions introduce factual or theoretical material from the syllabus that can be studied independently.
When substances are mixed together physically, they can be combined in any proportion. Mixtures can be homogeneous (with uniform properties throughout, for example, air) or heterogeneous (in which the composition varies and components may be in different phases, like a mixture of gravel and water). Mixtures can usually be separated by physical processes such as filtration or distillation. However, when substances react to give a chemical compound, their proportions are fixed in a stoichiometric ratio and they can only be separated again by a chemical reaction.

Stoichiometric calculations are central to chemistry. For a general stoichiometric equation of the form:

\[ aA + bB \rightarrow xX + yY \]

in which \( a \) moles of A reacts with \( b \) moles of B, \( a, b, x \) and \( y \) are the stoichiometric coefficients. These stoichiometric coefficients show the ratios in which chemical species react with one another. An equation with correct stoichiometric coefficients is said to be balanced, with the same number of each type of atom on each side.

To formulate and balance stoichiometric equations quickly, it is useful to memorize the formulas and charges of common ions (table 1.1.1).

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula and charge</th>
<th>Name</th>
<th>Formula and charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>ammonium</td>
<td>( \text{NH}_4^+ )</td>
<td>nitrite</td>
<td>( \text{NO}_2^- )</td>
</tr>
<tr>
<td>carbonate</td>
<td>( \text{CO}_3^{2-} )</td>
<td>nitrate</td>
<td>( \text{NO}_3^- )</td>
</tr>
<tr>
<td>hydrogencarbonate</td>
<td>( \text{HCO}_3^- )</td>
<td>sulfite</td>
<td>( \text{SO}_3^{2-} )</td>
</tr>
<tr>
<td>ethanedioate (oxalate)</td>
<td>( \text{C}_2\text{O}_4^{2-} )</td>
<td>sulfate</td>
<td>( \text{SO}_4^{2-} )</td>
</tr>
<tr>
<td>phosphate</td>
<td>( \text{PO}_4^{3-} )</td>
<td>thiosulfate</td>
<td>( \text{S}_2\text{O}_3^{2-} )</td>
</tr>
</tbody>
</table>

Table 1.1.1 The names, formulas and charges of common polyatomic ions

Chemical equations often include state symbols: solid (s), liquid (l), gas (g) and aqueous solution (aq), which means dissolved in water.
You should know:
✔ masses of atoms are measured relative to $^{12}\text{C}$ and expressed as relative atomic mass ($A_r$) and relative formula/molecular mass ($M_r$), which have no units;
✔ the mole is a measure of the amount of substance, $n$, and refers to a very large, fixed number of entities ($6.02 \times 10^{23}$);
✔ molar mass (mass of one mole of a substance), $M$, has the derived SI unit g mol$^{-1}$;
✔ an empirical formula is the simplest ratio of the atoms of each element in a compound;
✔ a molecular formula is the actual number of atoms of each element in a molecule.

You should be able to:
✔ calculate the molar masses of atoms, ions, molecules and formula units;
✔ solve numerical problems involving the relationships between $n$, $m$ and $M$;
✔ calculate empirical and molecular formulas and percentage composition by mass from given data.

In order to determine stoichiometric ratios from observations, chemists need a way to calculate the amount of substance—the number of atoms, molecules or ions in a known mass of that substance.

The masses of atoms of most elements have been measured with a high degree of accuracy. For example, an atom of carbon has a mass of $1.993 \times 10^{-26}$ kg. However, it is more convenient to express masses of atoms and molecules as ratios relative to the mass of the $^{12}\text{C}$ atom, which is defined as 12.00 on the relative scale. These ratios are known as relative atomic mass ($A_r$) and relative molecular mass ($M_r$), respectively, and have no units.

The SI (Système International d’Unités) is the metric system of measurement. It has seven base units, one of which is the mole, the SI unit for amount of substance, symbol $n$. One mole contains $6.02 \times 10^{23}$ elementary entities, just as one dozen represents a collection of 12 objects. This number is the fixed numerical value of the Avogadro constant, $N_A$.

The mole applies to elementary entities (atoms, molecules, ions, electrons, other particles, or specified groups of such particles).

**Example 1.1.1.**
Formulate a balanced equation, including state symbols, for the reaction of potassium hydroxide, KOH, with phosphoric acid, $\text{H}_3\text{PO}_4$, in aqueous solution.

**Solution**
First, write the formulas of the reactants and products.

$$\text{KOH} + \text{H}_3\text{PO}_4 \rightarrow \text{K}_3\text{PO}_4 + \text{H}_2\text{O}$$

Then balance the equation so that the numbers of atoms on both sides are equal. Do this by adjusting the coefficients on each side.

$$3\text{KOH} + \text{H}_3\text{PO}_4 \rightarrow \text{K}_3\text{PO}_4 + 3\text{H}_2\text{O}$$

Finally, add the state symbols. Aqueous solutions are involved, so (aq) is used for all species except water.

$$3\text{KOH(aq)} + \text{H}_3\text{PO}_4(aq) \rightarrow \text{K}_3\text{PO}_4(aq) + 3\text{H}_2\text{O(l)}$$

**Assessment tip**
Remember, the chemical formula of a substance should never be changed when balancing chemical equations, only its coefficient.
The carbon-12 atom (¹²C) is an isotope, a concept discussed in topic 2.1.

**Assessment tip**

Prefixes (e.g., M, k, m, μ, p) are frequently used to form decimal multiples and submultiples of SI units. Do not forget to apply conversion factors when using these prefixes. You should also ensure that your final answer is expressed in the units indicated in the question.

**Example 1.2.1.**

An extra-strength aspirin tablet contains 500 mg of acetylsalicylic acid, C₉H₈O₄. Calculate the number of molecules of acetylsalicylic acid in the tablet.

**Solution**

Calculate the molar mass, \(M\), of acetylsalicylic acid (using relative atomic masses from the periodic table in section 6 of the data booklet):

\[
M = (9 \times 12.01) + (8 \times 1.01) + (4 \times 16.00) = 180.17 \text{ g mol}^{-1}
\]

Convert \(m\) (acetylsalicylic acid) from mg to g (1 mg = \(10^{-3}\) g):

\[
500 \text{ mg} = 500 \times 10^{-3} \text{ g} = 0.500 \text{ g}
\]

Calculate the amount \(n\) of acetylsalicylic acid:

\[
n = \frac{0.500 \text{ g}}{180.17 \text{ g mol}^{-1}} \approx 2.78 \times 10^{-3} \text{ mol}
\]

Finally, use the relationship: 1 mol \(\equiv 6.02 \times 10^{23}\) molecules.

\[
2.78 \times 10^{-3} \text{ mol} \equiv (6.02 \times 10^{23})(2.78 \times 10^{-3}) \approx 1.67 \times 10^{21} \text{ molecules of acetylsalicylic acid.}
\]

**Assessment tip**

It is best practice to write relative atomic masses correct to two decimal places, as in the data booklet. For example, \(A_m\) for hydrogen is written as 1.01, not 1. Use of integer values can lead to inaccuracies in multi-step solutions to examination questions.

**Maths skills**

A numerical value should reflect the precision of its measurement. For multiplication or division, the result is expressed based on the measurement with the smallest number of significant figures (sf). For addition or subtraction, the result is expressed based on the measurement with the smallest number of decimal places.

If the number you are rounding to a certain number of significant figures or decimal places is followed by 5, 6, 7, 8 or 9, round the number up. If it is followed by 0, 1, 2, 3 or 4, round the number down.

**Example 1.2.2.**

Determine the percentage of magnesium present in magnesium phosphate, correct to three significant figures.

**Solution**

First, work out the formula for magnesium phosphate:

The phosphate ion is \(\text{PO}_4^{3-}\) and the magnesium ion is \(\text{Mg}^{2+}\) (magnesium belongs to group 2 of the periodic table and loses its two valence electrons when ionized). By balancing the charges, magnesium phosphate will have the chemical formula \(\text{Mg}_3(\text{PO}_4)_2\).

Then calculate the molar mass, \(M\), for \(\text{Mg}_3(\text{PO}_4)_2\):

\[
M = (3 \times 24.31) + (2 \times 30.97) + (8 \times 16.00) = 262.87 \text{ g mol}^{-1}
\]

Finally calculate the percentage of magnesium in \(\text{Mg}_3(\text{PO}_4)_2\):

\[
\%\text{Mg} = \frac{3 \times 24.31}{262.87} \times 100 \approx 27.7\% \text{ to 3 sf.}
\]

This question links topics 1.2, The mole concept, and 4.1, Ionic bonding and structure. Such linkage is common in IB Chemistry examination papers, especially for stoichiometry.

**Assessment tip**

Note that the final mark given for the correct numerical answer would be lost if the answer were not given to the correct number of sf.
Salbutamol, a drug used to treat asthma, contains carbon, hydrogen, nitrogen and oxygen, and has molar mass \( M = 239.35 \text{ g mol}^{-1} \). In a laboratory analysis, the drug was found to contain 65.2\% C, 8.9\% H and 5.9\% N by mass. Deduce the molecular formula of salbutamol.

**Solution**

The mass percent of oxygen in salbutamol can be worked out from \( 100 - (65.2 + 8.9 + 5.9) = 20.0\% \).

Now we can determine the empirical formula of salbutamol:

<table>
<thead>
<tr>
<th>Element</th>
<th>%</th>
<th>( n )/mol</th>
<th>Divide by smallest value of ( n )</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>65.2</td>
<td>65.2/12.01 ≈ 5.43</td>
<td>5.43/0.42 ≈ 13</td>
</tr>
<tr>
<td>H</td>
<td>8.9</td>
<td>8.9/1.01 ≈ 8.8</td>
<td>8.8/0.42 ≈ 21</td>
</tr>
<tr>
<td>N</td>
<td>5.9</td>
<td>5.9/14.01 ≈ 0.42</td>
<td>0.42/0.42 ≈ 1</td>
</tr>
<tr>
<td>O</td>
<td>20.0</td>
<td>20.0/16.00 ≈ 1.25</td>
<td>1.25/0.42 ≈ 3</td>
</tr>
</tbody>
</table>

Empirical formula = \( \text{C}_{13}\text{H}_{21}\text{NO}_3 \)

\( M(\text{empirical formula}) = (13 \times 12.01) + (21 \times 1.01) + (14.01) + (3 \times 16.00) = 239.35 \text{ g mol}^{-1} \)

Since \( M(\text{molecular formula}) \) is also 239.35 g mol\(^{-1}\), the empirical formula for salbutamol is the same as its molecular formula, \( \text{C}_{13}\text{H}_{21}\text{NO}_3 \).

**TOPIC 1.3 REACTING MASSES AND VOLUMES**

**You should know:**
- the amount of limiting reactant controls the amount of product formed in a chemical reaction;
- the experimental yield is usually lower than the theoretical yield;
- Avogadro’s law states that equal volumes of gases measured at the same temperature and pressure contain equal numbers of molecules;
- the molar volume of an ideal gas is a constant at a specified temperature and pressure;
- the molar concentration of a solute, \( c \), is the amount of solute, \( n \), in a given volume, \( V \), of the solution;
- a standard solution is one with a known concentration of solute.

**You should be able to:**
- solve numerical problems involving reacting quantities, limiting reactants, and theoretical, experimental and percentage yields;
- calculate reacting volumes of gases by applying Avogadro’s law;
- solve problems and analyse graphs involving \( T, p \) and \( V \) for a fixed mass of an ideal gas;
- solve numerical problems using the ideal gas equation, \( pV = nRT \);
- explain why real gases deviate from ideal behaviour at high pressure and low temperature;
- solve problems involving dilution, mixing of solutions and titration.

Mole ratios in chemical equations can be used to calculate reacting ratios by mass, concentration and volume.

When two substances react with each other, the one that is used up completely is called the *limiting reactant*. The reactant that is not entirely consumed is said to be present in *excess*. The expected amount of product from the reaction, the theoretical yield, is calculated from the amount of the limiting reactant, but is rarely obtained in practice because of side reactions and losses on separation and purification. The percentage yield can be calculated as follows:

\[
\text{percentage yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100\%
\]

**Assessment tip**

If the subscripts representing the number of atoms in the calculated empirical formula are not integer values, multiply all the subscripts by a factor to generate integer values for the number of atoms. For example, if a subscript is 0.25, multiply all of the subscripts by a factor of 4.
Example 1.3.1.

5.25 kg of hydrogen, H₂, reacts with 28.2 kg of nitrogen, N₂, to form 15.5 kg of ammonia, NH₃.

a) Formulate a balanced chemical equation for this reaction, including state symbols.

b) Deduce the limiting reactant.

c) Calculate the theoretical yield of ammonia, in kg, correct to three significant figures.

d) Determine the percentage yield of ammonia, correct to one decimal place.

Solution

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

b) Step 1: Work out the amount, in mol, of each reactant, \(n(\text{H}_2)\) and \(n(\text{N}_2)\).

In the equation \(n = \frac{m}{M}\), \(m\) is expressed in g.

Hence, you need to convert kg to g.

\[
\begin{align*}
n(\text{H}_2) &= \frac{5.25 \times 10^3}{2 \times 1.01} \approx 2.60 \times 10^3 \text{ mol} \\
n(\text{N}_2) &= \frac{28.2 \times 10^3}{2 \times 14.01} \approx 1.01 \times 10^3 \text{ mol}
\end{align*}
\]

Step 2: Consider the stoichiometric ratio between \(\text{N}_2\) and \(\text{H}_2\).

\[
\begin{align*}
1 \text{ mol N}_2(g) &= 3 \text{ mol H}_2(g) \\
1.01 \times 10^3 \text{ mol N}_2(g) &= 3.03 \times 10^3 \text{ mol H}_2(g)
\end{align*}
\]

Step 3: \(n(\text{H}_2)\) reacting with \(\text{N}_2\) is \(2.60 \times 10^3\) mol \(n(\text{H}_2)\) needed for complete reaction = \(3.03 \times 10^3\) mol

Since \(n(\text{H}_2)\) used < \(n(\text{H}_2)\) needed, hydrogen is the limiting reactant.

c) Determine the amount, in mol, of ammonia expected from the limiting reactant:

\[
\begin{align*}
3 \text{ mol H}_2(g) &= 2 \text{ mol NH}_3(g) \\
2.60 \times 10^3 \text{ mol H}_2(g) &= \frac{2}{3} (2.60 \times 10^3) \text{ mol NH}_3(g) \\
&= 1.73 \times 10^3 \text{ mol NH}_3(g)
\end{align*}
\]

Convert this amount to mass in g, using the expression \(n = \frac{m}{M}\):

\[
m(\text{NH}_3) = n \times M = (1.73 \times 10^3 \text{ mol})(17.04 \text{ g mol}^{-1}) = 2.95 \times 10^4 \text{ g}
\]

Finally, convert the mass into kg and express your answer to 3 sf:

\[
n(\text{NH}_3) = 29.5 \text{ kg}
\]

d) Percentage yield = \(\frac{15.5 \text{ kg}}{29.5 \text{ kg}} \times 100\% \approx 52.5\%\)

The behaviour of ideal gases can be described by three laws. Boyle’s law states that the pressure of a fixed mass of an ideal gas is inversely proportional to its volume at a constant temperature, \(p \propto \frac{1}{V}\).

Charles’s law states that the volume of a fixed mass of an ideal gas is proportional to its absolute temperature (in kelvin) at constant pressure, \(V \propto T\), and finally Gay-Lussac’s law states that \(p \propto T\) for absolute temperature and a constant volume of gas. Together, these gas laws give the expression: \(\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}\)
For reactions in the gas phase, reacting ratios can be calculated using Avogadro’s law: equal volumes of gases measured at the same temperature and pressure contain equal numbers of molecules. This proportionality, combined with the gas laws and a constant $R$, the gas constant, gives the ideal gas equation, or equation of state:

$$pV = nRT$$

It follows that 1 mol of any ideal gas has the same volume at a specified temperature and pressure. Under standard conditions (STP) of $T = 273$ K (0°C) and $p = 100$ kPa, the molar volume of an ideal gas is 22.7 dm$^3$ mol$^{-1}$.

An ideal gas obeys the gas laws exactly, but real gases deviate from ideal gas behaviour because some intermolecular forces of attraction exist between the gaseous particles, slightly altering their speeds and collision behaviour, and because particles in a real gas occupy space. These deviations become noticeable at high pressures and low temperatures:

- At high pressures, the gas is compressed, so the space occupied by gas particles is no longer negligible compared with the volume of the gas, so the volume is larger than that for an ideal gas.
- At low temperatures, gas particles have little kinetic energy to overcome attractive forces between them, so the volume is smaller than that for an ideal gas.

**Example 1.3.2.**

Calculate the volume of hydrogen gas produced, in cm$^3$, at 32°C and 90.5 kPa, when 6.55 g of gallium reacts with an excess of hydrochloric acid.

$$2\text{Ga(s)} + 6\text{HCl(aq)} \rightarrow 2\text{GaCl}_3\text{(aq)} + 3\text{H}_2\text{(g)}$$

**Solution**

Since the question states that hydrochloric acid is in excess, gallium must be the limiting reactant.

Therefore, to deduce the amount of hydrogen gas produced, first calculate the amount of gallium.

The atomic mass of gallium is 69.72 g mol$^{-1}$.

So, $n = \frac{6.55 \text{ g}}{69.72 \text{ g mol}^{-1}} \approx 0.0939 \text{ mol}$

Then consider the stoichiometric ratio between gallium and hydrogen:

$$2 \text{ mol Ga(s)} = 3 \text{ mol H}_2\text{(g)}, \text{ so } 1 \text{ mol Ga(s)} = \frac{3}{2} \text{ mol H}_2\text{(g)}$$

Hence, 0.0939 mol Ga(s) = $\frac{3}{2} (0.0939)$ mol H$_2$ (g) = 0.141 mol H$_2$ (g)

To calculate the volume $V$ of H$_2$ (g), use the ideal gas equation, $pV = nRT$.

Collect all the required data and ensure that correct units are used:

$n = 0.141 \text{ mol}, R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}, T = 32 + 273 = 305 \text{ K}, $ $p = 90.5 \text{ kPa} = 9.05 \times 10^4 \text{ Pa}$

Rearranging the equation and inserting the data gives:

$$V = \frac{nRT}{p} = \frac{0.141 \text{ mol} \times 8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 305 \text{ K}}{9.05 \times 10^4 \text{ Pa}}$$

$$\approx 3.95 \times 10^{-3} \text{ m}^3$$

Finally convert m$^3$ to cm$^3$: $V = 3.95 \times 10^4 \text{ cm}^3$

**Assessment tip**

Remember to convert °C to K for temperature in calculations.
For reactions taking place in solution, quantities can be calculated from concentrations. The molar concentration of a solute (dissolved substance), \( c \), in mol dm\(^{-3} \), is related to the amount of the solute, \( n \), in mol and the volume, \( V \), of the solution in dm\(^3 \) by the expression:

\[
c = \frac{n}{V}
\]

Other typical units of concentration, \( c \), are g dm\(^{-3} \) and ppm (1 ppm = 1 mg dm\(^{-3} \)).

When the concentration of a solute is not known, it can be found by reacting it with a standard solution and comparing their reacting volumes, taking into account the stoichiometric equation for the reaction. This is the principle of titration.

**Example 1.3.3.**

Sodium hydroxide reacts with sulfuric acid in aqueous solution to form a salt and water.

**a)** Formulate a balanced chemical equation for this reaction, including state symbols.

**b)** Calculate the volume, in dm\(^3 \), of 0.350 mol dm\(^{-3} \) sodium hydroxide solution that will neutralize 25.0 cm\(^3 \) of 0.250 mol dm\(^{-3} \) sulfuric acid solution in a titration.

**Solution**

**a)**

\[
2\text{NaOH(aq)} + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O(l)}
\]

**b)**

\[
V(\text{H}_2\text{SO}_4) = 25.0 \text{ cm}^3 = 0.0250 \text{ dm}^3
\]

\[
n(\text{H}_2\text{SO}_4) = 0.0250 \text{ dm}^3 \times 0.250 \text{ mol dm}^{-3} \approx 0.00625 \text{ mol}
\]

2 mol NaOH = 1 mol H\(_2\)SO\(_4\), so 

\[
n(\text{NaOH}) = 2 \times 0.00625 \text{ mol} = 0.0125 \text{ mol}
\]

\[
V(\text{NaOH}) = \frac{0.0125 \text{ mol}}{0.350 \text{ mol dm}^{-3}} \approx 0.0357 \text{ dm}^3
\]

Titrations involving redox and acid–base reactions are discussed in topics 9.1 and 18.3, respectively. Chemical stoichiometry is also linked to equilibrium calculations in topic 17.1.

**2.478 g of white phosphorus was used to make phosphine according to the equation:**

\[
P_4(s) + 30\text{H}^-(aq) + 3\text{H}_2\text{O}(l) \rightarrow \text{PH}_3(g) + 3\text{H}_3\text{PO}_2(aq)
\]

**a)** Calculate the amount, in mol, of white phosphorus used. [1]

**b)** This phosphorus was reacted with 100 cm\(^3 \) of 5.00 mol dm\(^{-3} \) aqueous sodium hydroxide. Deduce, showing your working, which was the limiting reactant. [1]

**c)** Determine the excess amount, in mol, of the other reactant. [1]

**d)** Determine the volume of phosphine, measured in cm\(^3 \) at standard temperature and pressure, that was produced. [1]
Problem 1
Formulate a balanced equation, including state symbols, for the reaction of nitric acid with calcium hydroxide.

Problem 2
Calculate the number of ions present in 0.25 mol of calcium nitrate.

Problem 3
Compound X has an empirical formula CH$_2$O and a molar mass of 60.06 g mol$^{-1}$. Deduce the molecular formula of X.

Problem 4
Compound Y is a hydrocarbon and has a molar mass of 86.20 g mol$^{-1}$. Upon combustion, Y produces 1.75 g CO$_2$ and 0.836 g H$_2$O. Deduce the molecular formula for Y.

Problem 5
1.7 g of NaNO$_3$ ($M = 85.00$) is dissolved in water to prepare 0.10 dm$^3$ of solution. What is the concentration of the resulting solution in mol dm$^{-3}$?

A. $2.0 \times 10^{-4}$  B. $1.0 \times 10^{-1}$  C. $2.0 \times 10^{-1}$  D. 5.0

Problem 6
4.00 g of propane, C$_3$H$_8$, undergoes combustion in 68.2 g of oxygen.

a) Formulate a balanced chemical equation for this reaction, including state symbols.

b) Deduce the limiting reactant.

c) Calculate the theoretical yield, in g, of carbon dioxide formed.

Problem 7
Calculate the volume, in dm$^3$, of a balloon filled with 0.350 mol of hydrogen gas, at a temperature of 26.0°C and a pressure of 1.15 $\times$ 10$^2$ kPa.
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Steroids and cholesterol

Most steroids act as chemical messengers (hormones) that regulate metabolism, immune responses and reproductive functions. Anabolic steroids stimulate the growth of muscle tissue and have many medical uses, but are also abused in sports as performance-enhancing drugs. All steroids in the human body are synthesized from cholesterol, shown in section 34 of the data booklet, which is also an important component of cell membranes.

Example B.3.1.

Cholesterol is synthesized in the liver and has various biological functions.

a) Suggest, with a reason, whether cholesterol is soluble in water or not.

b) Describe how cholesterol is transported around the body.

Solution

a) The cholesterol molecule has a large hydrocarbon backbone and only one hydroxyl group. Its overall polarity is low, so it is insoluble in water.

b) Cholesterol is transported from the liver to body tissues by the blood in the form of complexes with low-density lipoproteins (LDL). High-density lipoproteins (HDL) form more stable complexes with cholesterol and transport it back to the liver, where it is metabolized.

Sunflower oil contains stearic, oleic and linoleic fatty acids. The structural formulas of these acids are given in section 34 of the data booklet.

a) Explain which one of these fatty acids has the highest boiling point. [2]

b) 10.0 g of sunflower oil reacts completely with 123 cm³ of 0.500 mol dm⁻³ iodine solution. Calculate the iodine number of sunflower oil to the nearest whole number. [3]

This answer could have achieved 4/5 marks:

a) Stearic acid, as it is saturated and so molecules can pack closer together, giving stronger London dispersion forces between molecules.

b) \[ n(I_2) = 0.123 \times 0.500 = 0.0615 \text{ mol}; \]
\[ m(I_2) = 126.9 \times 0.0615 \approx 7.8 \text{ g} \]
\[ \frac{10.0 \text{ g acid}}{7.8 \text{ g iodine}} \Rightarrow \frac{78 \text{ g iodine}}{100 \text{ g acid}}. \]

Iodine number is 78.

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Steroids and cholesterol

Key structures are shown or described. Responses that correctly explain the difference in the character of the three fatty acids are awarded up to 1 mark. For example: “Stearic acid is saturated and has the highest boiling point.” However, to score 2 marks, they should also explain the difference in boiling point in terms of London forces, but not the detail given in my mark scheme: “Stearic acid is saturated and has the highest boiling point because its hydrocarbon chain packs more tightly with London forces, giving it a higher boiling point than the other fatty acids, which are unsaturated and have lower boiling points.” Any given reason, which does not proceed logically from the above, is awarded up to 1 mark. A response that is not supported by any of the above, or which does not include a reason, is awarded up to 1 mark.

b) [3 marks]

Iodine number: 78

Assessment tips

HDL cholesterol (HDL-C) and LDL cholesterol (LDL-C) are sometimes called “good cholesterol” and “bad cholesterol”, respectively. You should never use such colloquial names in examinations, as they will not be accepted.

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