OXFORD IB PREPARED

CHEMISTRY

IB DIPLOMA PROGRAMME

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Contents

1 STOICHIOMETRIC RELATIONSHIPS

TOPIC 1.1 INTRODUCTION TO THE PARTICULATE NATURE OF MATTER AND CHEMICAL CHANGE

You should know:

- ✔ atoms of different elements combine in fixed ratios to form compounds, which differ in properties from their constituent elements;
- ✔ a mixture is a combination of two or more substances that retain their individual properties;
- ✔ mixtures can be homogeneous or heterogeneous.

You should be able to:

- ✔ deduce chemical equations from given reactants and products;
- ✔ apply state symbols in equations;
- ✔ explain observed changes in physical properties and temperature during a change of state.

- **Chemical stoichiometry** is the relationship between the amounts of the reactants and products in a chemical reaction.
- **Stoichiometric coefficients** describe the ratios in which amounts of species react with one another.

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).

Name	Formula and charge	Name	Formula and charge
ammonium	NH_4^+	nitrite	NO_2^-
carbonate	CO_3^{2-}	nitrate	NO_3^-
hydrogencarbonate	HCO_3^-	sulfite	SO_3^{2-}
ethanedioate (oxalate)	$C_2O_4^{2-}$	sulfate	SO_4^{2-}
phosphate	PO_4^{3-}	thiosulfate	$S_2O_3^{2-}$

▲ **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.

Assessment tip

In some questions, state symbols are required and you will be penalized if these are not included. Remember that the state symbol for water in the liquid phase is (l), not (aq): $H_2O(l)$.

Symbols and names of chemical elements can be found in section 5 of the data booklet.

Example 1.1.1.

Formulate a balanced equation, including state symbols, for the reaction of potassium hydroxide, KOH, with phosphoric acid, H_3PO_4, in aqueous solution.

Solution

First, write the formulas of the reactants and products.

$KOH + H_3PO_4 \rightarrow K_3PO_4 + H_2O$

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

$3KOH + H_3PO_4 \rightarrow K_3PO_4 + 3H_2O$

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

$3KOH(aq) + H_3PO_4(aq) \rightarrow K_3PO_4(aq) + 3H_2O(l)$

Assessment tip

Remember, the chemical formula of a substance should never be changed when balancing chemical equations, only its coefficient.

TOPIC 1.2 THE MOLE CONCEPT

You should know:

- ✔ masses of atoms are measured relative to ^{12}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 (6.02×10^{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 1.993×10^{-26} kg. However, it is more convenient to express masses of atoms and molecules as ratios relative to the mass of the ^{12}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 6.02×10^{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).

- **Relative atomic mass** (A_r) is the ratio of the average mass of an atom of a chemical element in a given sample to one-twelfth of the mass of a carbon-12 atom. Since the value is relative, it has no units. The terms **relative molecular mass** and **relative formula mass** (both M_r) are used for molecules and ionic species, respectively.
- The **amount of substance**, n, is the number of atoms, molecules or ions, expressed in moles, in a given quantity of the substance.
- The **mole** (abbreviated to mol) is the SI unit for amount of substance.
- The **Avogadro constant**, N_A, 6.02×10^{23} mol^{-1}, is the number of particles in 1 mol. Without units, it is called the **Avogadro number**.

The carbon-12 atom (^{12}C) 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.

Assessment tip

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

The amount of substance, n, is calculated from the mass and the molar mass as follows:

$$n\,(\text{mol}) = \frac{m\,(\text{g})}{M\,(\text{g mol}^{-1})}$$

Example 1.2.1.

An extra-strength aspirin tablet contains 500 mg of acetylsalicylic acid, $C_9H_8O_4$. 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 ≡ 6.02×10^{23} molecules.

So 2.78×10^{-3} mol ≡ $(6.02 \times 10^{23})(2.78 \times 10^{-3}) \approx 1.67 \times 10^{21}$ molecules of acetylsalicylic acid.

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.

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.

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 PO_4^{3-} and the magnesium ion is 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 $Mg_3(PO_4)_2$.

Then calculate the molar mass, M, for $Mg_3(PO_4)_2$:

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

Finally calculate the percentage of magnesium in $Mg_3(PO_4)_2$:

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

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.

Example 1.2.3.

Salbutamol, a drug used to treat asthma, contains carbon, hydrogen, nitrogen and oxygen, and has molar mass $M = 239.35$ 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:

Element	%	*n* / mol	Divide by smallest value of *n*
C	65.2	$65.2/12.01 \approx 5.43$	$5.43/0.42 \approx 13$
H	8.9	$8.9/1.01 \approx 8.8$	$8.8/0.42 \approx 21$
N	5.9	$5.9/14.01 \approx 0.42$	$0.42/0.42 \approx 1$
O	20.0	$20.0/16.00 \approx 1.25$	$1.25/0.42 \approx 3$

Empirical formula $= C_{13}H_{21}NO_3$

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

Since M(molecular formula) is also 239.35 g mol^{-1}, the empirical formula for salbutamol is the same as its molecular formula, $C_{13}H_{21}NO_3$.

- **Empirical formula** is the simplest ratio of the atoms of each element in a compound.
- **Molecular formula** is the actual number of atoms of each element in a molecule. For example, for benzene the molecular formula is C_6H_6, but the empirical formula is CH.

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.

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:

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

- A **limiting reactant** is the reactant that is used up completely, and that limits the amount of product formed in a chemical reaction.
- An **excess reactant** is present in a reaction mixture in a quantity greater than needed to react with another reactant.
- **Percentage yield** = (experimental yield/theoretical yield) × 100%.

Example 1.3.1.

5.25 kg of hydrogen, H_2, reacts with 28.2 kg of nitrogen, N_2, to form 15.5 kg of ammonia, NH_3.

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) $3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$

b) Step 1: Work out the amount, in mol, of each reactant, $n(H_2)$ and $n(N_2)$.

In the equation $n = \frac{m}{M}$, m is expressed in g. Hence, you need to convert kg to g.

$$n(H_2) = \frac{5.25 \times 10^3}{2 \times 1.01} \approx 2.60 \times 10^3 \text{ mol}$$

$$n(N_2) = \frac{28.2 \times 10^3}{2 \times 14.01} \approx 1.01 \times 10^3 \text{ mol}$$

Step 2: Consider the stoichiometric ratio between $N_2(g)$ and $H_2(g)$.

$$1 \text{ mol } N_2(g) \equiv 3 \text{ mol } H_2(g)$$

$$1.01 \times 10^3 \text{ mol } N_2(g) \equiv 3.03 \times 10^3 \text{ mol } H_2(g)$$

Step 3: $n(H_2)$ reacting with N_2 is 2.60×10^3 mol

$n(H_2)$ needed for complete reaction = 3.03×10^3 mol

Since $n(H_2)$ used < $n(H_2)$ needed, hydrogen is the limiting reactant.

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

$$3 \text{ mol } H_2(g) \equiv 2 \text{ mol } NH_3(g)$$

$$2.60 \times 10^3 \text{ mol } H_2(g) \equiv \frac{2}{3}(2.60 \times 10^3) \text{ mol } NH_3(g) \approx 1.73 \times 10^3 \text{ mol } NH_3(g)$$

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

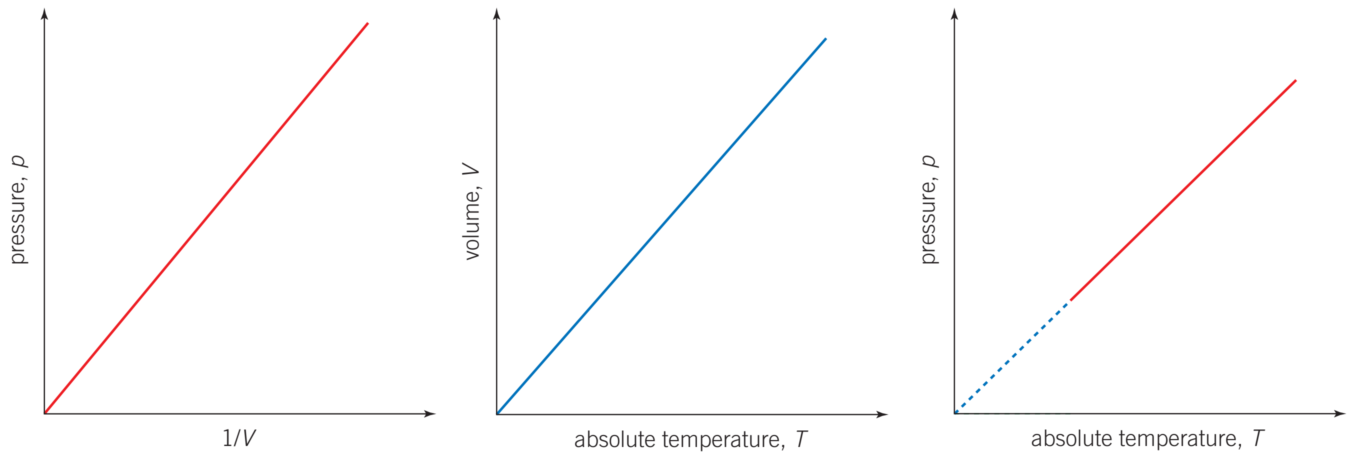
$$m(NH_3) = n \times M = (1.73 \times 10^3 \text{ mol})(17.04 \text{ g mol}^{-1}) \approx 2.95 \times 10^4 \text{ g}$$

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

$n(NH_3) = 29.5$ 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: $\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$



▲ **Figure 1.3.1.** The behaviour of ideal gases: Boyle's law (left), Charles's law (middle) and Gay-Lussac's law (right)

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.

- **Avogadro's law** states that equal volumes of gases measured at the same temperature and pressure contain equal number of molecules
- An **ideal gas** is a gas that obeys the equation of state, $pV = nRT$, also known as the **ideal gas equation**. The particles of an ideal gas have negligible volume and collide elastically.

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.

$2Ga(s) + 6HCl(aq) \rightarrow 2GaCl_3(aq) + 3H_2(g)$

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

Then consider the stoichiometric ratio between gallium and hydrogen:

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

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

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:

$n = 0.141$ mol, $R = 8.31\ \text{J K}^{-1}\ \text{mol}^{-1}$, $T = 32 + 273 = 305$ K,
$p = 90.5\ \text{kPa} = 9.05 \times 10^4$ 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: $V = 3.95 \times 10^3\ cm^3$

The ideal gas equation can be found in section 1 of the data booklet. The value of the gas constant R is listed in section 2, as is the conversion from m^3 to cm^3: $10^{-3}\ m^3 = 10^3\ cm^3$.

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}).

A **standard solution** is one with a known concentration of a solute.

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*.

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.

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) $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

b) $V(H_2SO_4) = 25.0\ cm^3 = 0.0250\ dm^3$

$n(H_2SO_4) = 0.0250\ dm^3 \times 0.250\ mol\ dm^{-3} \approx 0.00625\ mol$

$2\ mol\ NaOH \equiv 1\ mol\ H_2SO_4$, so $n(NaOH) = 2 \times 0.00625\ mol = 0.0125\ mol$

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

Assessment tip

Always ensure that you are using the correct units in numerical questions. In this question, both cm^3 and dm^3 are used, so the calculations involve conversions.

Volume conversion factors are given in section 2 of the data booklet: 1 dm^3 = 10^3 cm^3

SAMPLE STUDENT ANSWER

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

$P_4(s) + 3OH^-(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2^-(aq)$

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]

This answer could have achieved 1/4 marks:

a) $n = (2.478)/(4 \times 31) = 0.02$ mol

b) $n(NaOH) = (100 \times 5.00)/(1000) = 0.500$ mol

1 mol $P_4 \equiv 3$ mol OH^-

0.02 mol P_4 used $\equiv 0.06$ mol OH^-

$n(NaOH)$ used $= 0.500$ mol and $n(NaOH)$ needed $=$ 0.06 mol

c) 0.500 mol

d) 1 mol of a gas at STP occupies 22.7 dm^3. Since 1 mol $P_4 \equiv 1$ mol PH_3, 0.02 mol $P_4 \equiv 0.02$ mol PH_3. This is equivalent to (0.02×22.7) $dm^3 = 0.454$ dm^3.

▲ Correct expression for n ($n = m/M$) and A_r of P correctly multiplied by four, as species is P_4, but exact value for A_r of P (30.97) should have been used

▼ The candidate has ignored significant figures: the correct answer is $n = 0.02000$ mol

▼ Calculation is correct, but the question is not answered, as the limiting reactant (P_4) must be explicitly stated

▼ This is the amount of NaOH used, not the amount in excess; the excess could be calculated as follows:
excess $= 0.500 - 0.0600 = 0.440$ mol

▼ Correct method, but the units of cm^3 were required, so the correct answer is 454 cm^3

Practice problems for Topic 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_2O 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_2O.

Deduce the molecular formula for Y.

Problem 5
1.7 g of $NaNO_3$ ($M_r = 85.00$) 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. 2.0×10^{-4} **B.** 1.0×10^{-1} **C.** 2.0×10^{-1} **D.** 5.0

Problem 6
4.00 g of propane, C_3H_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 1.15×10^2 kPa.

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CHEMISTRY

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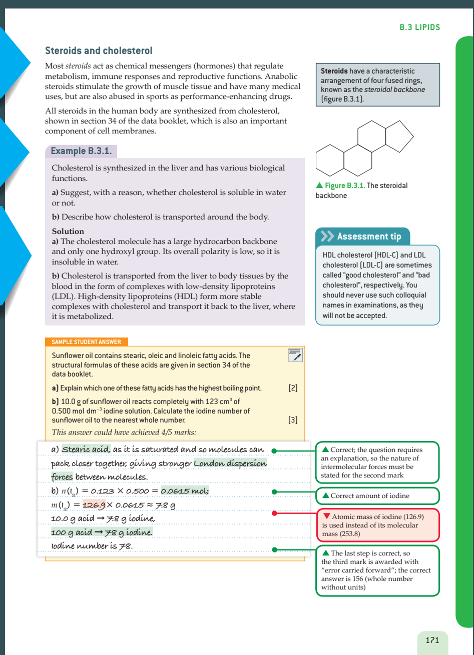

B.3 LIPIDS

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.

Steroids have a characteristic arrangement of four fused rings, known as the *steroidal backbone* (figure B.3.1).

▲ Figure B.3.1. The steroidal backbone

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.

Assessment tip

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.

SAMPLE STUDENT ANSWER

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 cm^3 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) $n(I_2) = 0.123 \times 0.500 = 0.0615$ mol;

$m(I_2) = 126.9 \times 0.0615 \approx 7.8$ g

10.0 g acid → 7.8 g iodine,

100 g acid → 78 g iodine.

Iodine number is 78.

▲ Correct; the question requires an explanation, so the nature of intermolecular forces must be stated for the second mark

▲ Correct amount of iodine

▼ Atomic mass of iodine (126.9) is used instead of its molecular mass (253.8)

▲ The last step is correct, so the third mark is awarded with "error carried forward"; the correct answer is 156 (whole number without units)

171

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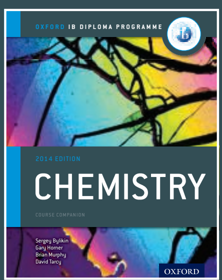

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FOR FIRST ASSESSMENT IN 2016

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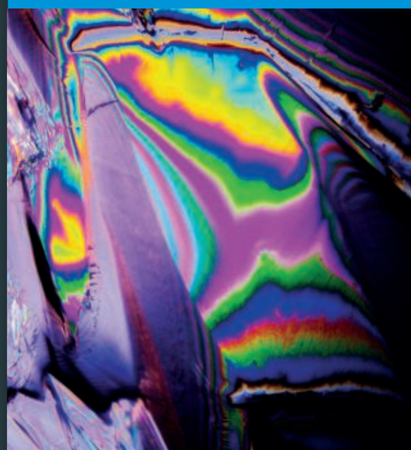

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ISBN 978-0-19-842367-6