IBDP Notes HL/SL Chemistry

Topic 1: Stoichiometric Relationships

## 1.1 Introduction to the particulate nature of matter and chemical change

Matter : Any substance (made up of one or more atoms) with space and mass.

Atom : The smallest particle of a chemical element that can exist.

Compound : Atoms of different elements chemically bond in fixed ratios to form compounds. E.g NaCl (ratio of Na: Cl is 1:1)

Molecule : A group of atoms held together by chemical bonding, representing the smallest fundamental unit of a chemical compound.

Why is water ($H_2O$) a compound and not an element? Water can be chemically broken down into constituent elements.

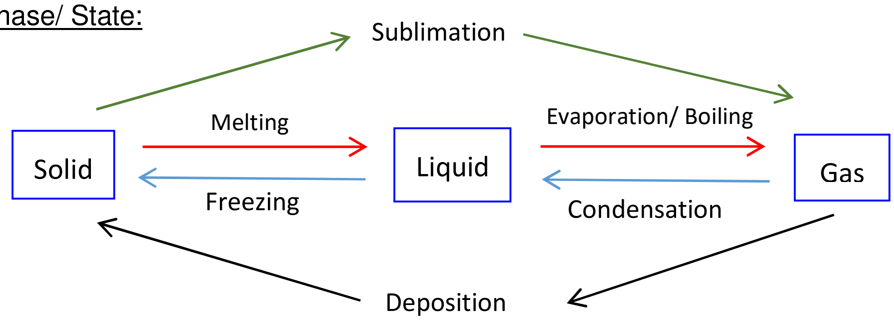
Difference between compounds & elements: Elemental $H_2$ and $O_2$ gas have different chemical & physical properties compared to the compound $H_2O$.

Mixtures contain 2 or more elements or compounds that are not chemically bonded together. Hence, properties of mixtures retain the individual properties of the components.

Homogenous mixture: All components are in the same physical phase and have similar properties. E.g (Salt water, Air)

Heterogenous mixture: Components are in different physical phases and have varying properties. (Oil & Water – subtances remain physically separate)

Changes in phase/ State:




## 1.2 The mole concept

### Mole:

- A unit that describes the amount of substance. Substance could be atoms, molecules, ions or electrons.
- The term "mole" describes a number – **$6.02 \times 10^{23}$**
- One mole of substance = **$6.023 \times 10^{23}$** atoms, molecules, ions or electrons.

1 mole of $^{12}C$ atom: $6.023 \times 10^{23}$ $^{12}C$ atoms
1 mole of $NO_3^-$ ions: $6.023 \times 10^{23}$ of $NO_3^-$ ions

Relative atomic mass ($A_r$) : Ratio of the average mass of the isotope of atoms to $\frac{1}{12}$ mass of an atom of $^{12}C$ atom. No units.

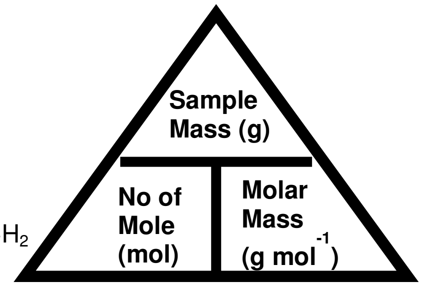
Relative molecular mass ($M_r$) : Ratio of the average mass of the molecule to $\frac{1}{12}$ mass of an atom of $^{12}C$ atom. No units.

Molar Mass (M) : The mass of one mole of a substance. Units: g $mol^{-1}$

E.g. Iron (Fe) has a Molar Mass of 55.85 g $mol^{-1}$. This means that one mole of iron atoms ($6.023 \times 10^{23}$ iron atoms) has a mass of 55.85 g.

Determining the no. of moles of substance (or Molar Mass, Sample mass):

$$no. of\ moles = \frac{sample\ mass\ (g)}{molar\ mass\ (g\ mol^{-1})}$$




Empirical formula: Simplest ratio of the compound - E.g. Empirical formula of $C_2H_4$ is $CH_2$

Molecular formula: Actual number of atoms present in a molecule respectively.

Molecular formula: n (Empirical Formula) & Molecular Mass: n (Empirical Mass)

Structural formula: Shows the arrangement of atoms and bonds within a molecule. Used in organic chemistry.

You may be asked to find the empirical formula of a compound given its percentage composition by mass.

Example 1: A particular compound is found to be composed of 86% Carbon and 14% Hydrogen by mass, what is its empirical formula?

Step 1 – Write out the percentages in terms of grams (i.e. instead of 86% Carbon out of 100%, we have 86g of Carbon in a sample of 100g of the compound).

Step 2 – Divide by the Molar Mass of the element (determining the number of moles present in 100g of compound.

Carbon: 86 ÷ 12 = 7

Hydrogen: 14 ÷ 1 = 14

Step 3 – Compare the ratio of moles present.

C : H → 7 : 14 = 1 : 2

Therefore, the empirical formula of the compound is $CH_2$.

Example 2a: A compound is found to be composed of 9% Hydrogen, 55% Carbon and 36% Oxygen by mass. Find its empirical formula.

| | Hydrogen | Carbon | Oxygen | |
|---|---|---|---|---|
| % in terms of grams | 9 | 55 | 36 | |
| Number of moles of atoms | 9 ÷ 1 = 9 ×4 | 55 ÷ 12 = 4.58 | 36 ÷ 16 = 2.25 | The no. of moles must be converted into a whole number! |
| | 36 | ≈ 18.32 | 9 | |
| Ratio of moles | ÷9 4 | ≈ 2 | 1 | Now we have the simplest ratio |

So we have an empirical formula of $C_2H_4O$.

Example 2b: Given that the molar mass of the compound is 132, determine its molecular formula.

Empirical formula $C_2H_4O$ has a molar mass of: $(2\times12) + (1\times4) + (1\times16) = 44 g\ mol^{-1}$

$132 \div 44 = 3$

This tells us that the actual molar mass is 3 times larger than the empirical mass. Therefore, the molecular formula is:

$C_6H_{12}O_3$

## Practice Questions from IBDP HL/SL Chemistry (May 2010-2011) for 1.1 & 1.2

a) How many molecules are present in a drop of ethanol, $C_2H_5OH$, of mass $2.3 \times 10^{-3}$ g?

($L = 6.0 \times 10^{23}\ mol^{-1}$)

**b)** The relative molecular mass of a gas is 56 and its empirical formula is $CH_2$. What is the molecular formula of the gas?

**c)** What is the total number of hydrogen atoms in 1.0 mol of benzamide, $C_6H_5CONH_2$?

**d)** Which is both an empirical and a molecular formula?

$C_5H_{12}$ $C_5H_{10}$ $C_4H_8$ $C_4H_{10}$

**e)** What is the mass, in g, of one molecule of ethane, $C_2H_6$?

**f)** Which molecular formula is also an empirical formula?

$PCl_3$ $C_2H_4$ $H_2O_2$ $C_6H_{12}O_6$

**g)** What is the total number of nitrogen atoms in **two** mol of $NH_4NO_3$?

**h)** Which compound has the empirical formula with the largest mass?

$C_2H_6$ $C_2H_4$ $C_2H_2$ $C_3H_6$

## 1.3 Reacting Masses and Volume

- Chemical Equations: Reactants react to produce substances which are known as products in a fixed relationship (i.e mole ratio).
- State symbols: *(s)* solid, *(l)* liquid (or molten), *(g)* gas and *(aq)* aqueous.
- Single arrow (→) : Reaction goes to completion. All the reactants have been consumed to form products completely.
- Reversible arrows (⇆): Reaction that has both reactant and product at the end (i.e equilibrium). The reaction proceeds in both directions, with reactants forming products and products reforming reactants simultaneously.
- Ionic equations: Used to describe the dissociation of ionic compounds in water or precipitation reactions.

For example:

$NaCl$ (s) → $Na^+$ (aq) + $Cl^-$ (aq)    Dissociation of ionic compounds

$Ag^+$ (aq) + $Cl^-$ (aq) → $AgCl$ (s)    Precipitation of ionic compounds

Determining the limiting reagent / reactant:

Reactants that are added in excess to what is required.

E.g.

2 Ca (s) + $O_2$ (g) → 2CaO (s)

<u>What does this chemical equation mean?</u>

**Moles: 2 moles** of Ca reacts with **1 mole** of $O_2$ to produce 2 moles of CaO.

**Mass: 80.2 g** of Calcium reacts with **32.0 g** of Oxygen to produce **112.0 g** of Calcium Oxide.

<u>What kind of questions will appear in exams?</u>

Any **random mass** of the reactants will appear.

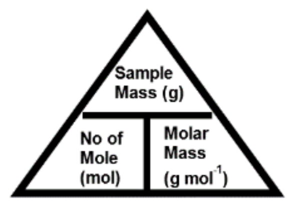
**<u>Type 1: One of the reactant is in excess.</u>**

3.0 g of Calcium reacted in *excess* oxygen to produce Calcium Oxide. Find the mass of Calcium Oxide produced?

2 Ca (s) + $O_2$ (g) → 2CaO (s)

Answer:

No of moles or Amount of Calcium: $\frac{Sample\ Mass}{Molar\ Mass} = \frac{3.0}{40.1} = 0.0748\ mol$




Mole ratio of Ca: CaO → 1:1    Thus, amount of CaO: 0.0748 mol

Mass of Calcium oxide: No of moles x Molar Mass = 0.0748 x 56.1 = **4.196 g**

**<u>Type 2: Determining the limiting agent.</u>**

3.0 g of Calcium reacted with 4.0 g of oxygen to produce Calcium Oxide. Find the mass of Calcium Oxide produced.

<u>Amount of Calcium: 0.0748 mol</u>    <u>Amount of Oxygen : 4.0 / 32 .0 = 0.125 mol</u>

**Recall***    **Moles: 2 mol** of Ca reacts with **1 mol** of $O_2$

From Calcium's point of view: 0.0748 mol of Ca requires only 0.0374 mol of $O_2$.

➔ Oxygen is in excess as we have 0.125 mol.

From Oxygen's point of view: 0.125 mol of $O_2$ requires 0.250 mol of Ca.

➔ Ca is limited as we only have 0.0748 mol.

Only Calcium will be completely used up to form Calcium Oxide. Mole ratio of Ca:CaO is 2:2 = 1:1

Therefore, Amount of Ca = Amount of CaO = 0.0748

Mass of Calcium oxide: No of moles x Molar Mass = 0.0748 x 56.1 = **4.196 g**

**Balancing Equations**

You will frequently be asked to balance equations in exams. Just ensure that each element has the same number of moles present in the reactant and product side of the equation.

Example 1: Na (s) + $H_2O$ (l) → NaOH (aq) + $H_2$ (g)

In this example we have 2 moles of hydrogen atoms on the reactant side, but 3 on the product side. Therefore, the equation is not balanced. In order to balance it, we add numbers in front of each reactant/ product.

$$Na + 2H_2O \rightarrow 2NaOH + H_2$$

Now there are 4 moles hydrogen and 2 moles of Oxygen on both sides. However, the sodium is not balanced.

$$2Na\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq) + H_2\ (g)$$

Practice Balancing Equations

a) $MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + CO_2 + H_2O$

b) $C_4H_8 + O_2 \rightarrow CO_2 + H_2O$

c) $Pb(NO_3)_2 + NaCl \rightarrow PbCl_2 + NaNO_3$

Moles and Concentration

$$Concentration\ (mol\ dm^{-3}) = \frac{No. of\ Moles\ (mol)}{Volume\ (dm^{-3})}$$

This equation is used to calculate the number of moles of a solute (dissolved chemical) in a particular sample of solution.

Example: Find the number of moles of NaCl present in 0.300 $dm^{-3}$ of 2.0 $moldm^{-3}$ aqueous NaCl solution.

Moles = Conc. × Volume
= 2.0 × 0.300
= 0.60 mol

Calculating Moles of Gas

At *room temperature* (25°C / 298K) and pressure, 1 mole of gas has a volume of 24$dm^3$.

$$Moles\ of\ Gas = \frac{Volume\ (dm^3)}{24}$$

1 $dm^3$ = 1000$cm^3$

However, at *standard temperature* (0°C / 273K) and pressure, 1 mole of gas has a volume of 22.7$dm^3$.

$$Moles\ of\ Gas = \frac{Volume\ (dm^3)}{22.7}$$

These two equations are only to be used to calculate the moles of GAS molecules.
Example: Find the number of moles of hydrogen molecules in a container of volume 300 $cm^3$ at:

(a) Room temperature
(b) Standard temperature.

Recall: $300cm^3 = 0.300dm^3$
(a) 0.300 ÷ 24 = 0.0125mol
(b) 0.300 ÷ 22.7 ≈ 0.0132mol

<u>Ideal Gases</u>

When working with gases, we often approximate their behaviour to an "ideal gas" to simplify any necessary calculations. There are 2 important assumptions to note about an ideal gas:

1. The gas molecules do not interact with one another.
2. The volume of each gas particle is zero.

Real gases deviate from this ideal behaviour at:

1. High pressure -
   The gas particles are pushed so closely together that they occupy a larger volume collectively, to the point where the volume of each particle is not negligible.

2. Low temperature –
   Each particle has very little kinetic energy, thus when they move close to other particles, they are more likely to experience some interaction due to intermolecular forces.

The ideal gas equation:

$$PV = nRT$$

| Symbol | Represents… | Unit |
|---|---|---|
| P | Pressure | Pascals (Pa) |
| V | Volume | $m^3$ ( = $1000dm^3$) |
| n | no. of moles | mol |
| R | Gas constant = 8.31 | $J\ K^{-1}\ mol^{-1}$ |
| T | Temperature | K |

This equation allows us to calculate the molar mass of a gas, provided its mass is given.
Example: A sample of a gas is kept in a cylinder of volume 1.00 $dm^3$, at 100 kPa and a temperature of 18.0°C. It is found to have a mass of 1.65g. Calculate the molar mass of the gas.

Convert all given data into appropriate units:

- 1 $dm^3$ = 0.00100 $m^3$
- 100 kPa = 100 000 Pa
- 18.0°C = 291 K

$$PV = nRT$$

Rearrange to make $n$ the subject: $n = \frac{PV}{RT}$

$$n = \frac{100000 \times 0.001}{8.31 \times 291} \approx 0.0414$$

Recall: $Molar\ Mass = \frac{Mass}{Moles}$

$$Molar\ Mass = \frac{1.65}{0.414} \approx 39.9gmol^{-1}$$

Alternatively, we can rearrange this equation to allow us to predict the response of a set amount of gas to a change in conditions.

$$\frac{PV}{T} = nR$$

For a fixed sample of gas (no. of moles does not change), the variables on the right side are both constants. This tells us that the final value of $\frac{PV}{T}$ will not change even if one of the variables involved does.

Example: A fixed sample of gas has the initial conditions of 80 kPa and a temperature of 278 K. The pressure is then increased to 110 kPa, with the volume remaining constant. Calculate the final temperature of the gas.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

1 = Initial

2 = Final

$$T_2 = \frac{T_1P_2V_2}{P_1V_1} = \frac{278 \times 110000}{80000} \approx 382K$$

## Common IB Stoichiometric Relationships Questions

**1.** A toxic gas, A, consists of 53.8 % nitrogen and 46.2 % carbon by mass. At 273 K and $1.01\times10^5$ Pa, 1.048 g of A occupies 462 $cm^3$. Determine the empirical formula of A. Calculate the molar mass of the compound and determine its molecular structure.

**(Total 3 marks)**

**2.** empirical formula = CN;

*Working must be shown to get point.*

$M_r$ = 51.9 (g $mol^{-1}$);
$:N{\equiv}C{-}C{\equiv}N:$; 3

**[3]**

**3.** The data below is from an experiment used to determine the percentage of iron present in a sample of iron ore. This sample was dissolved in acid and all of the iron was converted to $Fe^{2+}$. The resulting solution was titrated with a standard solution of potassium manganate(VII), $KMnO_4$. This procedure was carried out three times. In acidic solution, $MnO_4^-$ reacts with $Fe^{2+}$ ions to form $Mn^{2+}$ and $Fe^{3+}$ and the end point is indicated by a slight pink colour.

| **titre** | **1** | **2** | **3** |
|---|---|---|---|
| **initial burette reading / $cm^3$** | 1.00 | 23.60 | 10.00 |
| **Final burette reading / $cm^3$** | 24.60 | 46.10 | 32.50 |

| | |
|---|---|
| **Mass of iron ore / g** | $3.682 \times 10^{-1}$ |
| **concentration of $KMnO_4$ solution / mol $dm^{-3}$** | $2.152 \times 10^{-2}$ |

(a) Deduce the balanced redox equation for this reaction in **acidic** solution.

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**(2)**

(b) Identify the reducing agent in the reaction.

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**(1)**

(c) Calculate the amount, in moles, of $MnO_4^-$ used in the titration.

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**(2)**

(d) Calculate the amount, in moles, of Fe present in the $3.682 \times 10^{-1}$ g sample of iron ore.

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**(2)**

(e) Determine the percentage by mass of Fe present in the $3.682 \times 10^{-1}$ g sample of iron ore.

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**(2)**
**(Total 9 marks)**

**4.** (a) $MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$
*Award **[2]** if correctly balanced.*
*Award **[1]** for correctly placing $H^+$ and $H_2O$.*
*Award **[1 max]** for correct balanced equation but with electrons shown.*
*Ignore state symbols.* 2

(b) $Fe^{2+}$/ iron(II);
*Do not accept iron.* 1

(c) $n = 2.152 \times 10^{-2} \times 2.250 \times 10^{-2}$;
$4.842 \times 10^{-4}$ (mol);
*Award **[1]** for correct volume*
*Award **[1]** for correct calculation.* 2

(d) 1 mol of $MnO_4^-$ reacts with 5 mol of $Fe^{2+}$;
$5 \times 4.842 \times 10^{-4} = 2.421 \times 10^{-3}$ (mol);
(same number of moles of Fe in the iron ore)
*Allow ECF from part (a) and (c) provided some mention of mole ratio is stated.* 2

(e) $2.421 \times 10^{-3} \times 55.85 = 0.1352$(g);
$\frac{0.1352}{0.3682} \times 100 = 36.72$ %;
*Allow ECF from part (d).* 2

**[9]**

**5.** The percentage of iron(II) ions, $Fe^{2+}$, in a vitamin tablet can be estimated by dissolving the tablet in dilute sulfuric acid and titrating with standard potassium manganate(VII) solution, $KMnO_4$(aq). During the process iron(II) is oxidized to iron(III) and the manganate(VII) ion is reduced to the manganese(II) ion, $Mn^{2+}$(aq). It was found that one tablet with a mass of 1.43 g required 11.6 $cm^3$ of $2.00 \times 10^{-2}$ mol $dm^{-3}$ $KMnO_4$(aq) to reach the end-point.

(a) (i) State the half-equation for the oxidation of the iron(II) ions.

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**(1)**

(ii) State the half-equation for the reduction of the $MnO_4^-$ ions in acidic solution.

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**(1)**

(iii) Deduce the overall redox equation for the reaction.

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**(1)**

(b) (i) Calculate the amount, in moles, of $MnO_4^-$ ions present in 11.6 $cm^3$ of $2.00 \times 10^{-2}$ mol $dm^{-3}$ $KMnO_4$(aq).

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(ii) Calculate the amount, in moles, of $Fe^{2+}$ ions present in the vitamin tablet.

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(iii) Determine the percentage by mass of $Fe^{2+}$ ions present in the vitamin tablet.

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**(Total 7 marks)**

**6.** (a) (i) $Fe^{2+} \rightarrow Fe^{3+} + e^-$; 1

(ii) $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$; 1

(iii) $MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+}\ 4H_2O$;
*Accept e instead of $e^-$.* 1

(b) (i) amount of $MnO_4^- = \dfrac{11.6}{1000} \times 0.0200 = 2.32 \times 10^{-4}$ mol; 1

(ii) amount of $Fe^{2+} = 5 \times 2.32 \times 10^{-4} = 1.16 \times 10^{-3}$ mol; 1

(iii) mass of $Fe^{2+} = 55.85 \times 1.16 \times 10^{-3} = 6.48 \times 10^{-2}$ g;

percentage of $Fe^{2+}$ in tablet $= \frac{6.48 \times 10^{-2}}{1.43} \times 100 = 4.53$ %; 2

**[7]**

**7.** The percentage by mass of calcium carbonate in eggshell was determined by adding excess hydrochloric acid to ensure that all the calcium carbonate had reacted. The excess acid left was then titrated with aqueous sodium hydroxide.

(a) A student added 27.20 $cm^3$ of 0.200 mol $dm^{-3}$ HCl to 0.188 g of eggshell. Calculate the amount, in mol, of HCl added.

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**(1)**

(b) The excess acid requires 23.80 $cm^3$ of 0.100 mol $dm^{-3}$ NaOH for neutralization. Calculate the amount, in mol, of acid that is in excess.

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**(1)**

(c) Determine the amount, in mol, of HCl that reacted with the calcium carbonate in the eggshell.

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**(1)**

(d) State the equation for the reaction of HCl with the calcium carbonate in the eggshell.

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**(2)**

(e) Determine the amount, in mol, of calcium carbonate in the sample of the eggshell.

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**(2)**

(f) Calculate the mass **and** the percentage by mass of calcium carbonate in the eggshell sample.

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**(3)**

(g) Deduce **one** assumption made in arriving at the percentage of calcium carbonate in the eggshell sample.

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**(1)**
**(Total 11 marks)**

**8.** (a) (i) *Copper:*
0 to +2 / increases by 2 / +2 / 2+;
*Allow zero/nought for 0.*

*Nitrogen:*
+5 to +4 / decreases by 1 / –1 / 1–;
*Penalize missing + sign or incorrect notation such as 2+, $2^+$ or II, once only.* 2

(ii) nitric acid/$HNO_3$ / $NO_3^-$/nitrate;
*Allow nitrogen from nitric acid/nitrate but not just nitrogen.* 1

(b) (i) $0.100 \times 0.0285$;
$2.85 \times 10^{-3}$ (mol);
*Award **[2]** for correct final answer.* 2

(ii) $2.85 \times 10^{-3}$ (mol); 1

(iii) $(63.55 \times 2.85 \times 10^{-3}) = 0.181$ g;
*Allow 63.5.* 1

(iv) $\left(\frac{0.181}{0.456} \times 100 =\right)$ 39.7 % 1

(v) $\left(\frac{44.2 - 39.7}{44.2} \times 100 =\right)$ 10/10.2 %;
*Allow 11.3 % i.e. percentage obtained in (iv) is used to divide instead of 44.2 %.* 1

(c) *Brass has:*
delocalized electrons / sea of mobile electrons / sea of electrons free to move;
*No mark for just "mobile electrons".* 1

**[10]**

**9.** Brass is a copper containing alloy with many uses. An analysis is carried out to determine the percentage of copper present in three identical samples of brass. The reactions involved in this analysis are shown below.

Step 1: $Cu(s) + 2HNO_3(aq) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$

Step 2: $4I^-(aq) + 2Cu^{2+}(aq) \rightarrow 2CuI(s) + I_2(aq)$

Step 3: $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

(a) (i) Deduce the change in the oxidation numbers of copper and nitrogen in step 1.

Copper:

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Nitrogen:

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**(2)**

(ii) Identify the oxidizing agent in step 1.

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**(1)**

(b) A student carried out this experiment three times, with three identical small brass nails, and obtained the following results.

Mass of brass = 0.456 g ± 0.001 g

| **Titre** | **1** | **2** | **3** |
|---|---|---|---|
| **Initial volume of 0.100 mol dm$^{-3}$ $S_2O_3^{2-}$ (± 0.05 cm$^3$)** | 0.00 | 0.00 | 0.00 |
| **Final volume of 0.100 mol dm$^{-3}$ $S_2O_3^{2-}$ (± 0.05 cm$^3$)** | 28.50 | 28.60 | 28.40 |
| **Volume added of 0.100 mol dm$^{-3}$ $S_2O_3^{2-}$ (± 0.10 cm$^3$)** | 28.50 | 28.60 | 28.40 |
| **Average volume added of 0.100 mol dm$^{-3}$ $S_2O_3^{2-}$(± 0.10 cm$^3$)** | 28.50 | | |

(i) Calculate the average amount, in mol, of $S_2O_3^{2-}$ added in step 3.

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**(2)**

(ii) Calculate the amount, in mol, of copper present in the brass.

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**(1)**

(iii) Calculate the mass of copper in the brass.

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**(1)**

(iv) Calculate the percentage by mass of copper in the brass.

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**(1)**

(v) The manufacturers claim that the sample of brass contains 44.2 % copper by mass. Determine the percentage error in the result.

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**(1)**

(c) With reference to its metallic structure, describe how brass conducts electricity.

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**(1)**
**(Total 10 marks)**

**10.** (a) $n$(HCl)(= 0.200 mol dm$^{-3}$ × 0.02720 dm$^3$) = 0.00544 / 5.44 × 10$^{-3}$ (mol); 1

(b) $n$(HCl) excess (= 0.100 mol dm$^{-3}$ × 0.02380 dm$^3$) = 0.00238 / 2.38 × 10$^{-3}$ (mol);

*Penalize not dividing by 1000 once only in (a) and (b).* 1

(c) $n$(HCl) reacted (= 0.00544 – 0.00238) = 0.00306 / 3.06 × 10$^{-3}$ (mol); 1

(d) $2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ /
$2H^+(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$;
*Award **[1]** for correct reactants and products.*
*Award **[1]** if this equation correctly balanced.*
*Award **[1 max]** for the following equations:*
*$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2CO_3(aq)$*
*$2H^+(aq) + CaCO_3(s) \rightarrow Ca^{2+}(aq) + H_2CO_3(aq)$*
*Ignore state symbols.* 2

(e) $n(CaCO_3) = (\frac{1}{2}n(HCl)) = \frac{1}{2} \times 0.00306$;
$= 0.00153 / 1.53 \times 10^{-3}$ (mol);
*Award **[2]** for correct final answer.* 2

(f) $M_r(CaCO_3)$ (= 40.08 + 12.01 + 3 × 16.00) = 100.09 /100.1 /
$M$ = 100.09 /100.1 (g $mol^{-1}$);
*Accept 100.*

$m(CaCO_3)(= nM) = 0.00153$ (mol) × 100.09 (g $mol^{-1}$) = 0.153 (g);

$\%CaCO_3 \left(= \frac{0.153}{0.188} \times 100\right) = 81.4\ \% / 81.5\ \%$

*Accept answers in the range 79.8 to 81.5 %.*
*Award **[3]** for correct final answer.* 3

(g) only $CaCO_3$ reacts with acid / impurities are inert/non-basic / impurities do not react with the acid / nothing else in the eggshell reacts with acid / no other carbonates;
*Do not accept "all calcium carbonate reacts with acid".* 1

**[11]**