Neap

Trial Examination 2023

VCE Chemistry Units 1&2

Written Examination

Suggested Solutions

SECTION A – MULTIPLE-CHOICE QUESTIONS

Question	Answer	Question	Answer	Question	Answer
1	D	11	D	21	A
2	C	12	C	22	C
3	B	13	D	23	A
4	B	14	A	24	A
5	D	15	C	25	D
6	C	16	A	26	B
7	A	17	C	27	B
8	C	18	B	28	C
9	C	19	D	29	C
10	B	20	A	30	D

Question 1 D

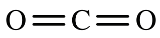
D is correct. Relative isotopic mass does not have a unit of measurement; it measures the mass of an atom of an isotope compared to a carbon-12 atom, which is taken to have a mass of 12 units exactly.

A is incorrect. Relative isotopic mass is not measured in grams.

B and **C** are incorrect. Relative isotopic mass refers to a mass comparison, not a sum of the number of isotopes or the number of atoms.

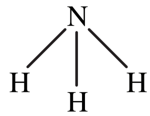
Question 2 C

Option **C** states the correct shapes of all four molecules. The shapes of the molecules are as shown below.

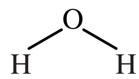


carbon dioxide

hydrogen chloride



ammonia



water

Question 3 B

B is correct. The ionic bonding model explains the generally high melting points of ionic compounds, which are due to the strength of attraction between the positive and negative ions in the lattice. However, the model does not readily explain the great variation in these temperatures for different ionic compounds. Characteristics such as ion size, charge density and type of lattice arrangement can contribute to the variation in temperatures.

A is incorrect. The ionic bonding model explains why ionic compounds tend to shatter. A sideways force causes the layers of ions in a compound's crystal to shift slightly, and so, for example, the negative ions will be close to each other and be repelled. This is observed as the shattering of the crystal.

C is incorrect. The ionic bonding model states that there is very strong attraction between positive and negative ions in ionic compounds, and so a large amount of force is required to disrupt this bonding.

D is incorrect. The ionic bonding model states that ions are held tightly in solid ionic compounds but are free to move in molten compounds. Thus, for example, a solid compound will not conduct electricity as its ions are unable to move, whereas the molten compound will conduct electricity because its ions are free to move.

Question 4 B

B is correct. Diamond has a three-dimensional covalent network lattice in which each carbon atom is bonded to four other carbon atoms. This structure does not have any delocalised electrons, which are required for electrical conduction.

A is incorrect. There is covalent bonding in both diamond and graphite, but graphite also has dispersion forces between its layers.

C is incorrect. In the structure of graphite, each carbon atom is covalently bonded to three other carbon atoms. One electron in the outer shell of each carbon atom is not confined in a covalent bond.

D is incorrect. Graphite has a two-dimensional layer lattice structure, not a three-dimensional network structure.

Question 5 D

D is correct. 'Lead' pencils are composed of powdered graphite and clay that are compressed into the pencil core and surrounded by a wooden casing. When a lead pencil is used, the graphite easily spreads across the page in layers.

A and **B** are incorrect. These options correctly state properties of diamond and graphite, but the properties are not relevant to the applications.

C is incorrect. Powdered graphite, not diamond, is used as a dry lubricant in locks to make them function smoothly. Powdered diamond would damage the metal parts of a lock, rendering the lock useless.

Question 6 C

C is correct. Hydrogen sulfide, H_2S, and water are polar molecules, so the intermolecular bonding between these molecules is dipole–dipole attraction.

A is incorrect. Dispersion forces would be present between H_2S molecules and water molecules, but they are not the strongest bonding between the molecules.

B is incorrect. Covalent bonds are the intramolecular bonds within H_2S molecules and water molecules; they do not form between the molecules.

D is incorrect. As H_2S does not contain a nitrogen, oxygen or fluorine atom bonded to a hydrogen atom, it cannot form hydrogen bonds.

Question 7 A

A is correct. The C–H bonds in methane, CH_4, are slightly polar due to the small difference in electronegativity between the atoms; however, the symmetrical CH_4 molecule cannot be represented as a dipole, so the molecule is non-polar.

B is incorrect. Carbon has a greater electronegativity than hydrogen, and so the carbon atom develops a very small negative charge (δ–).

C is incorrect. Alkane molecules are non-polar.

D is incorrect. CH_4 is a non-polar molecule and so does not interact with polar water molecules to any significant degree, meaning that CH_4 is insoluble in water.

Question 8 C

C is not a principle of green chemistry and is therefore the required response. Production pathways should be designed for maximum energy **efficiency** and minimal environmental damage. That is, the energy use should be minimised, not maximised.

A, **B** and **D** are all green chemistry principles and are therefore not the required response.

Question 9 C

Metal Z had the least amount of reaction in the three tests and so must be the least reactive. Metal X reacted in two of the tests, whereas metal Y reacted in all three tests. Therefore, the order of increasing reactivity is Z, X, Y.

Question 10 B

B is correct. Given that only metal Y reacts with cold water, it can be concluded that if the water were heated, the reaction would occur more vigorously. Furthermore, if steam were used, the reaction would be even more vigorous.

A is incorrect. Metal X reacted only slowly with hydrochloric acid and it did not react with cold water. There is no evidence provided in the table to suggest that this would change if the water were heated or if steam were used. In addition, this option does not include metal Y.

C and **D** are incorrect. Metal Z did not react with cold water or hydrochloric acid and so it would not be expected to react with steam.

Question 11 D

D is correct. Moving across a period, core charge increases as the number of protons increases, but the shielding effect of the inner shells of electrons remains constant. Thus, outer electrons are attracted more strongly towards the nucleus and so atomic radii decrease across a period. Moving down a group, the number of shells increases and core charge remains constant. (Core charge remains constant because, although the nuclear charge increases, the shielding effect of the inner shells also increases.) Since electronegativity measures electron-attracting power, it decreases down a group as electrons are not attracted as strongly due to their increasing distance from the nucleus.

A, **B** and **C** are incorrect. These options do not show the correct period and group trends.

Question 12 C

C is correct. Element M is a metal with one electron in its outer shell; the electron would be readily lost to produce a singly charged, positive ion. Element Q is a non-metal with six electrons in its outer shell. It would gain two electrons to produce a doubly charged, negative ion. In a reaction between element M and element Q, the positive and negative ions would be attracted to each other by ionic bonds, forming the compound M_2Q.

A is incorrect. M^+ and Q^{2-} form M_2Q, not MQ_2.

B and **D** are incorrect. The compound would be ionic.

Question 13 D

D is correct. This option is the balanced ionic equation that represents the formation of solid barium sulfate, $BaSO_4$, from solutions of barium chloride, $BaCl_2$, and magnesium sulfate, $MgSO_4$. It shows only the ions involved in the reaction.

A is incorrect. This option is not an ionic equation and incorrectly shows magnesium chloride, $MgCl_2$, as a precipitate, when it is actually soluble in aqueous solution.

B is incorrect. This option is the balanced full equation, not the ionic equation.

C is incorrect. This option incorrectly shows the reactant ions as solids.

Question 14 A

A is correct. To prepare a sample of pure $BaSO_4$, soluble ions trapped on the precipitate (substance X) must be removed by washing it with distilled water.

B is incorrect. $MgCl_2$ is in aqueous solution and so would not be trapped by the filter paper.

C is incorrect. $BaSO_4$ is the precipitate, which is trapped by the filter paper, and would not be present in the liquid (liquid Y) that passes through the paper.

D is incorrect. Liquid Y would contain water, magnesium ions and chloride ions. If one of the reactants were in excess, liquid Y may also contain either barium ions or sulfate ions – but not both – because at least one of these ions would be completely used up in forming the precipitate.

Question 15 C

C is correct. Each molecule of the compound contains 12 atoms. The molar mass is 90.0 g mol^{-1}. Therefore, 1 mol of molecules contains $12 \times 6.02 \times 10^{23}$ or 7.22×10^{24} atoms and has a mass of 90.0 g.

A is incorrect. The compound is a covalent molecular species, so there are no ions present.

B is incorrect. The molar mass is 90.0 g mol^{-1}. Relative molecular mass does not have a unit of measurement.

D is incorrect. In any sample of the compound, the number of oxygen atoms equals the number of carbon atoms but, as the relative atomic masses of oxygen and carbon differ, the masses of oxygen and carbon in the compound are not equal.

Question 16 A

A is correct. Intermolecular hydrogen bonding is relatively strong and so requires a considerable amount of energy to be disrupted; it must be disrupted to raise the temperature of water by heating.

B is incorrect. There are no ion-dipole attractions between water molecules.

C is incorrect. Specific heat capacity is relevant to intermolecular bonding not covalent bonding, which is the bonding within water molecules.

D is incorrect. All molecules have intermolecular dispersion forces, but this weak bonding is not responsible for a high specific heat capacity value.

Question 17 C

$[OH^-] = 2 \times 0.100 = 0.200 \text{ M} = 10^{-0.698}$ M

At 25°C, $K_w = 1.00 \times 10^{-14} = [H_3O^+][OH^-] = [H_3O^+] \times 10^{-0.698}$

Thus, $[H_3O^+] = 10^{-13.3}$ M and so $pH = -\log_{10}[H_3O^+] = -\log_{10}10^{-13.3} = 13.3$.

Question 18 B

B is correct. A strong monoprotic acid, such as hydrochloric acid (HCl), at 1.0 M would have a pH of 0. Methanoic acid, HCOOH, is an organic acid. All organic acids are weak acids as they only partially ionise and donate the proton from the COOH group to water. Hence, HCOOH is monoprotic and has a pH greater than 0. Sulfuric acid, H_2SO_4, is a strong acid that can donate two protons per molecule, making it diprotic. The pH is therefore less than 0.

A and **C** are incorrect. HCOOH is a monoprotic, weak acid.

D is incorrect. H_2SO_4 is a strong acid.

Question 19 D

D is correct. At the same concentration, each acid has the same number of molecules. As H_2SO_4 has two ionisable hydrogen atoms per molecule, it requires twice the amount of sodium hydrogen carbonate powder, $NaHCO_3$, for neutralisation than HCOOH, which can only donate one hydrogen atom per molecule.

A and **B** are incorrect. The pH is not relevant. The pH of a solution relates to the extent of ionisation of an acid in water, not to the number of hydrogen ions available for reaction with a base. For example, a 1.0 M HCl solution, being monoprotic, would require the same mass of $NaHCO_3$ for neutralisation as HCOOH, even though it has a lower pH than HCOOH.

C is incorrect. At the same concentration, each acid has the same number of molecules. However, H_2SO_4 requires more $NaHCO_3$ powder for neutralisation because of the number of hydrogen ions per molecule that it donates to react with the base. It is diprotic, while methanoic acid is monoprotic.

Question 20 A

A is correct. Co^{3+} accepts an electron from Fe^{2+} in this reaction. Oxidising agents cause oxidation by acting as electron acceptors, allowing other species to lose their electrons. When accepting electrons, the oxidising agent is reduced.

B is incorrect. Co^{3+} accepts an electron; electron acceptors are oxidising agents, not reducing agents.

C is incorrect. The reaction is a redox reaction because electrons are transferred. The total number of electrons accepted equals the total number donated, and so there is no overall change in charge.

D is incorrect. Redox reactions involve electron transfer, not proton transfer. Additionally, the Fe^{2+} ion and the Fe^{3+} ion have the same number of protons as they are both iron ions.

Question 21 A

At 10°C, the pH is 7.3 and so $[H_3O^+] = 10^{-7.3}$ M.

$K_w = [H_3O^+][OH^-]$ and in pure water $[H_3O^+] = [OH^-]$. So $K_w = [H_3O^+]^2 = (10^{-7.3})^2 = 10^{-14.6}$.

Question 22 C

At 45°C, the pH is 6.7 and so $[H_3O^+] = 10^{-6.7}$ M.

In pure water $[H_3O^+] = [OH^-]$, and so $[OH^-] = 10^{-6.7} = 2.0 \times 10^{-7}$ M.

Question 23 A

A is correct. The graph shows that increasing temperature results in the lowering of pH. As it is pure water, $[H_3O^+] = [OH^-]$ at all temperatures and so the water remains neutral.

B and **D** are incorrect. As it is pure water, $[H_3O^+] = [OH^-]$ at all temperatures and so the water remains neutral.

C is incorrect. The pH decreases with increasing temperature and so $[H_3O^+]$ increases. In pure water, $[H_3O^+] = [OH^-]$ and so $[OH^-]$ also increases with increasing temperature.

Question 24 A

24.8 L is the molar volume at standard laboratory conditions (SLC).

Comparing 1 mol of each gas:

$$\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$$

$$V_2 = \frac{p_1V_1T_2}{T_1p_2} = \frac{100 \times 24.8 \times 303}{298 \times 105} = 24.0 \text{ L}$$

Question 25 D

	Accuracy	Precision
A.	This is the definition of reproducibility.	This is the definition of repeatability.
B.	This is the definition of precision.	This is the definition of accuracy
C.	This is the definition of repeatability.	This is the definition of reproducibility.
D.	Accuracy refers to how close measured values are to the true value.	Precision refers to how close measured values are to the mean value.

Question 26 B

B is correct. An electronic pH meter is calibrated regularly and gives results that are consistently reliable. No human judgement is necessary and so the values are consistent with the true value. Repeated measurements of the pH of a solution would be expected to be consistently close to the mean value.

A is incorrect. A natural indicator cannot be used to determine the pH of a solution but rather is used to determine if a solution is acidic or basic.

C and **D** are incorrect. Universal indicator can be used to determine the approximate pH of a solution but requires human judgement to match colours using a standard colour chart. As the pH values are approximate, it is likely that the measurement would not show high levels of accuracy or precision.

Question 27 B

Steps 1, 3, 4 and 6 involve making solutions of a known concentration of a compound, measuring the conductivity of these solutions, constructing a calibration curve, reading the conductivity of a water sample under investigation and using the calibration curve to determine the corresponding salt concentration.

Step 2 would be used in a gravimetric analysis of water salinity and step 5 would be used in a spectroscopic analysis. Therefore, neither of these steps are relevant to analysis by electrical conductivity.

Question 28 C

C is correct.

10.0 g in 2500 L = $\frac{10.0}{2500}$ g per L = 4.00×10^{-3} g L^{-1} = 4.00×10^{-4} g per 100 mL = 4.00×10^{-4} % (m/v)

A is incorrect. 10.0 g in 2500 L = 4.00×10^{-3} g L^{-1}

B is incorrect. 10.0 g in 2500 L = 4.00×10^{-3} g L^{-1} = 4.00 g per 1000 L = 4.00 ppm

D is incorrect. 10.0 g in 2500 L = 4.00×10^{-3} g L^{-1} = $\frac{4.00}{71.0} \times 10^{-3}$ mol L^{-1} = 5.63×10^{-5} mol L^{-1}

Question 29 C

$n(BaSO_4) = \frac{m}{M} = \frac{0.153}{233.4} = 6.5553 \times 10^{-4}$ mol

$n(MgSO_4) = n(SO_4^{2-}) = n(BaSO_4)$

$m(MgSO_4)$ in 20.00 mL = $n \times M = 6.5553 \times 120.4 = 0.0789$ g

total mass of $MgSO_4$ in tank = $0.0789 \times \frac{50\,000}{20.0} = 197$ g

Question 30 D

D is correct. If the precipitate was not completely dry, the recorded mass would have been higher than the calculated average mass of precipitate due to the presence of water.

A is incorrect. Excess reagent, the barium chloride ($BaCl_2$), should have been used in all the precipitation reactions to ensure that all the sulfate ions were precipitated.

B is incorrect. If some of the precipitate was lost from the filter paper, then the final mass of the precipitate would have been less, not greater, than the average mass.

C is incorrect. If too much distilled water was used in washing, some of the precipitate could have been dissolved and lost through the filter paper. This would have resulted in a smaller final mass.

SECTION B

Question 1 (8 marks)

a. **i.** Magnesium metal, Mg, consists of a regular array of Mg^{2+} ions (cations) arranged in a lattice structure. 1 mark

Delocalised electrons released from the outer shells of the Mg atoms move throughout the lattice. 1 mark

ii. *Any one of:*

- Malleability: Applying force to the lattice causes the Mg^{2+} ions to move, which changes the shape of the metal, but does not cause the ions to break away from each other.
  The delocalised electrons are mutually attracted to the Mg^{2+} ions in all directions, and so the ions do not break free but are retained in the lattice structure.
- Ductility: Pulling on the lattice causes the Mg^{2+} ions to move to form a wire without the lattice breaking.
  The delocalised electrons are mutually attracted to the Mg^{2+} ions in all directions, and so the ions do not break free but are retained in the lattice structure.
- Heat conductivity: When the metal is heated, the delocalised electrons gain energy and move faster.
  The electrons bump into other components of the lattice, transferring the heat throughout the metal.

2 marks

b. A chlorine molecule, Cl_2, has two chlorine atoms covalently bonded to each other. 1 mark

As there is no electronegativity difference between the chlorine atoms, electrons are equally shared and so no permanent dipole forms, resulting in a non-polar molecule. 1 mark

c.

Sample	Will conduct electricity	Will not conduct electricity
solid $MgCl_2$		**Ions in the lattice are held so that they do not move apart; therefore, electricity will not be conducted.**
an aqueous solution of $MgCl_2$	**As the ionic lattice is dissolved, ions are able to move and so electricity will be conducted.**	

2 marks

1 mark for each correct explanation.

Question 2 (7 marks)

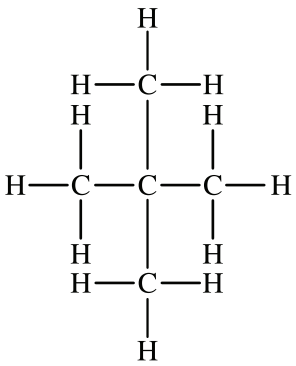
a. pentane 1 mark

b. **i.** methylbutane 1 mark

ii. As the molecules of the unbranched isomer can pack together more closely than the molecules in the branched isomer, the dispersion forces between the unbranched molecules are more intense. 1 mark

This means that a higher temperature is required to disrupt the dispersion forces to reach boiling point for the unbranched isomer. 1 mark

iii.



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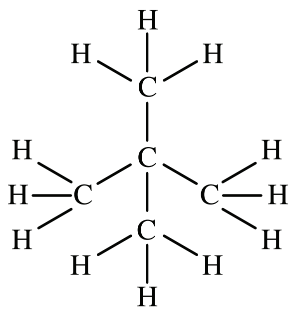


1 mark

c. All the outer shell electrons in the compound's molecules are involved in covalent bonding. 1 mark

As there are no free electrons or other charged particles such as ions, no charge can move in a sample of the compound; therefore, electricity is not conducted. 1 mark

Question 3 (8 marks)

a. In 100 g, the mass ratio of Cu : Fe : S = 34.6 : 30.4 : 35.0.

$$\text{mole ratio} = \frac{34.6}{63.5} : \frac{30.4}{55.8} : \frac{35.0}{32.1} = 0.544 : 0.544 : 1.09$$ 1 mark

Dividing by the lowest value gives 1 : 1 : 2.

Empirical formula: $CuFeS_2$ 1 mark

b. **i.** $^{63}_{29}Cu$ 1 mark

ii. If the abundances were 50 : 50, then the relative atomic mass of copper would be $\frac{62.93 + 64.93}{2} = 63.93$. 1 mark

As the relative atomic mass of copper is 63.54, there must be a higher percentage of the lighter isotope, and so isotope 1 has the greater abundance. 1 mark

c. *Any one of:*

- Copper is in the d-block of the periodic table.
- Copper is a transition metal.
- Copper is in the fourth period of the periodic table.
- Copper is in group 11 of the periodic table.

1 mark

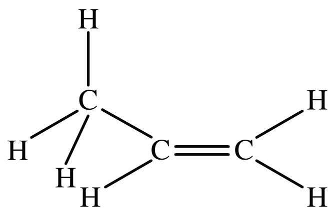
d. A linear economy uses resources to make items and then discards the items into landfill when they are no longer useful. 1 mark

Recycling copper items that are no longer useful keeps the copper metal in circulation, so new resources are not required; this is typical of a circular economy. 1 mark

Question 4 (6 marks)

a. a type of polymer that forms when the C=C double bonds in monomer molecules are broken and the resulting units are joined by covalent bonds in a long carbon chain 1 mark

b.



1 mark

c. A thermoplastic polymer has only weak bonding, such as dispersion forces, between the polymer chains. 1 mark

A thermosetting polymer has covalent bonds as well as dispersion forces between the polymer chains, forming a cross-linked structure. 1 mark

d. *Any one of:*

- A metal rod can be heated for a short time and then pushed against the polymer sample to see if the sample softens.
- A thermoplastic polymer will soften when in contact with the rod and harden if the rod is removed, whereas a thermosetting polymer will not soften.
- In a fume hood, a small sample of a polymer can be held above a Bunsen burner flame.
- A thermoplastic polymer will soften and melt, whereas a thermosetting polymer will char.

2 marks

1 mark for outlining the method.

1 mark for stating the expected result.

Question 5 (9 marks)

a. $n\left(\text{adipic acid, } C_6H_{10}O_4\right) = \frac{m}{M} = \frac{23.6}{146.1} = 0.1615 \text{ mol}$ 1 mark

1 mol of $C_6H_{10}O_4$ contains 10 mol of H atoms, and so n(H atoms) = 10 × 0.1615 = 1.615 mol. 1 mark

number of H atoms = $n \times N_A$ = 1.615 × 6.02 × 10^{23} = 9.72 × 10^{23} 1 mark

b. The strongest bonding between the molecules is hydrogen bonding, which is disrupted at the reasonably high temperature of 152°C and so melting occurs. 1 mark

This temperature is not sufficient to disrupt the covalent bonds that hold the atoms together in the molecule, and so the molecules are not broken down. 1 mark

c. The two COOH groups at both ends of the molecule are highly polar and there is only a short, non-polar hydrocarbon chain between them. 1 mark

The COOH groups can form hydrogen bonds and dipole–dipole attraction with water molecules, and so the molecule has high solubility, especially as temperature increases. 1 mark

d. *For example:*

Process A uses crude oil, which is a finite resource with low reserves, meaning that its supply is severely limited and not renewable. Process B uses a renewable plant source. 1 mark

Process A uses conditions that are heavily energy-dependent, uses hazardous chemicals and produces pollutants in the form of waste materials. Process B uses mild conditions, mild catalysts and produces little waste. 1 mark

Question 6 (6 marks)

a. One component, W, in the mixture lines up with the standard in lane 3 and so one component may possibly be identified. *(It is not certain that component W is standard 3. There may be other substances with the same R_f value.)* 1 mark

b. $R_f = \frac{\text{distance travelled by spot}}{\text{distance travelled by solvent front}} = \frac{5.0}{7.0} = 0.714 = 0.71$ 1 mark

c. The non-polar solvent would have dissolved any non-polar substances and so carried these closer to the solvent front. 1 mark

Component V is closest to the solvent front and so has the lowest overall polarity. 1 mark

d. Each component in the mixture has a different structure and so interacted with the mobile and stationary phases differently. 1 mark

As the components moved up the chromatography paper, the process of absorption–desorption occurred depending on the strength of these interactions, and so the components were able to be separated. 1 mark

Question 7 (9 marks)

a. $n(H_2) = \frac{m}{M} = \frac{100\,000}{2.0} = 50\,000 \text{ mol}$ 1 mark

$n(HCl) = 2 \times n(H_2) = 2 \times 50\,000 = 100\,000 \text{ mol}$ 1 mark

$pV = nRT$

$V = \frac{nRT}{p} = \frac{100\,000 \times 8.31 \times 293}{101.3} = 2.4 \times 10^6 \text{ L}$ 1 mark

b. Hydrogen chloride gas ionises completely when it reacts with water to produce hydrogen ions, which give the solution its acidic properties. 1 mark

No hydrogen ions are produced when hydrogen chloride gas dissolves in methylbenzene so this solution does not exhibit properties that are typical of an acid. 1 mark

c. **i.** $2HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$

2 marks

1 mark for the correct reactants and products.
1 mark for correct balancing and state symbols.

ii. $2NO_3^-(aq) + 10H^+(aq) + 8e^- \rightarrow N_2O(g) + 5H_2O(l)$

2 marks

1 mark for the correct reactants and products.
1 mark for correct balancing and state symbols.

Question 8 (5 marks)

a. **i.** The boiling points of group 16 hydrides increase with increasing relative molecular mass, that is, increasing from H_2S to H_2Se to H_2Te. 1 mark

ii. The major attraction between the hydride molecules is dispersion forces (some dipole–dipole bonding also occurs). 1 mark

Larger molecules have more intense dispersion forces between them, and so the boiling points increase with increasing molecular mass as more energy is needed to disrupt the stronger intermolecular bonding. 1 mark

b. As liquid water cools and approaches its freezing temperature of 0°C, the molecules form a unique arrangement, creating an ice crystal. 1 mark

Each water molecule forms hydrogen bonds with four neighbouring molecules in a very open arrangement, resulting in fewer molecules in a given volume (that is, a smaller mass for a given volume), which causes a sharp decrease in density. 1 mark

Question 9 (9 marks)

a. $Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$

2 marks
1 mark for the correct reactants and products.
1 mark for correct balancing and state symbols.

b. **i.** $n(HCl) = cV = 0.1000 \times 0.02045 = 0.002045$ mol — 1 mark

ii. $n(Na_2CO_3) = \frac{1}{2} \times n(HCl) = \frac{1}{2} \times 0.002045 = 0.001023$ mol — 1 mark

Note: Consequential on answer to ***Question 9b.i****.*

iii. $n(Na_2CO_3)$ in 250.0 mL = $0.001023 \times \frac{250.0}{20.00} = 0.01279$ mol — 1 mark

Note: Consequential on answer to ***Question 9b.ii****.*

iv. $m(Na_2CO_3) = n \times M = 0.01279 \times 106.0 = 1.356$ g — 1 mark

Note: Consequential on answer to ***Question 9b.iii****.*

v. $m(H_2O) = 1.585 - 1.356 = 0.229$ g — 1 mark

Note: Consequential on answer to ***Question 9b.iv****.*

vi. $n(H_2O) = \frac{0.229}{18.0} = 0.01277$ mol — 1 mark

The mole ratio of $Na_2CO_3 : H_2O = 0.01279 : 0.01277 = 1:1$
and so $x = 1$. — 1 mark

Note: Consequential on answers to ***Questions 9b.iii.*** *and* ***9b.v****.*
Note: The formula of hydrated sodium carbonate is $Na_2CO_3 \cdot H_2O$.

Question 10 (8 marks)

a. $AlCl_3(s) + 3Na(s) \rightarrow Al(s) + 3NaCl(s)$

2 marks
1 mark for the correct reactants and products.
1 mark for correct balancing and state symbols.

b. **i.** In a metal displacement reaction, there is always a transfer of electrons from the more reactive metal to the less reactive metal ion; therefore, it is a redox reaction. — 1 mark

ii. $Al(s) \rightarrow Al^{3+}(s) + 3e^-$ — 1 mark

iii. $Cr^{3+}(s) + 3e^- \rightarrow Cr(s)$ — 1 mark

c. $n(\text{Al}) = \frac{m}{M} = \frac{10^6}{27.0} = 3.7037 \times 10^4$ 1 mark

$n(\text{CO}_2) = \frac{3}{4} \times n(\text{Al}) = \frac{3}{4} \times 3.7037 \times 10^4 = 2.7778 \times 10^4$ 1 mark

$m(\text{CO}_2) = n \times M = 2.7778 \times 10^4 \times 44.0 = 1.2 \times 10^6$ g = 1.2 tonnes 1 mark

Question 11 (6 marks)

a.

Gas	Could this gas be gas B?	Explanation
N_2	yes	The nitrogen molecule is non-polar and so is only attracted to water molecules by dispersion forces, resulting in a similar low solubility to that of oxygen gas.
NH_3	no	The ammonia molecule is polar and capable of forming hydrogen bonds with water molecules; it should have a very high solubility in water.
CO_2	no	Carbon dioxide is a non-polar molecule but it reacts with water to form H^+ and CO_3^{2-} ions. Therefore, its solubility would be higher than that of oxygen.

3 marks
1 mark for each correct row.

b. i. The solubility of sodium chloride does not change significantly over the temperature range of 0 to 100°C. Most salts change solubility considerably over this range. 1 mark

ii. At 100°C, 34 g of the solute would have dissolved in 100 g of water (17 g in 50 g of water), so 30 − 17 = 13 g of crystals were not dissolved. 1 mark

At 20°C, 32 g of the solute would have dissolved in 100 g of water (16 g in 50 g of water) and so 17 − 16 = 1 g came out of the solution. Therefore, the total mass of crystals that would have been isolated is 13 + 1 = 14 g. 1 mark

Question 12 (9 marks)

a. At 510 nm, the coloured complex formed by *o*-phenanthroline with iron(II) ions (Fe^{2+}) absorbs at the maximum level, allowing for accurate absorbance readings. Any other chemical species present do not absorb at this wavelength. 1 mark

b. pipette 1 mark

c. An absorbance of 0.35 corresponds to 25×10^{-2} g L^{-1}. 1 mark

The original sample was diluted by a factor of 125 (× 25 × 5) in steps 2 and 3,

so the concentration in the sample in step 1 is $25 \times 10^{-2} \times 125 = 31.25$ g L^{-1} 1 mark

In the 2.0 mL sample in step 1, $c(Fe^{2+}) = \frac{31.25}{55.8} = 0.56$ M 1 mark

d. **i.** The data is classified as primary data because it was generated and used by the experimenter and not obtained from another source. 1 mark

The data is quantitative because it was measured using instruments and is not solely descriptive. 1 mark

ii. *For example*:

Using glassware that requires the experimenter to ensure that the lower part of the meniscus is level with the markings on the glassware is open to random error. 1 mark

Note: Accept any other reasonable and accurate explanation of the source of random error.

iii. *For example*:

Repeating the experiment and taking an average of the results could minimise the effect of random error as the average value would be closer to the true value. 1 mark

Note: Accept any other reasonable and accurate explanation of ways of minimising the effect of the random error from ***part d.ii****.*

Neap

Trial Examination 2023

VCE Chemistry Units 1&2

Written Examination

Data Booklet

Instructions

This data booklet is provided for your reference. A question and answer booklet is provided with this data booklet.

Students are NOT permitted to bring mobile phones and/or any other unauthorised electronic devices into the examination room.

1. Periodic table of the elements

Key: atomic number 79; symbol of element **Au**; relative atomic mass 197.0; name of element gold

1 **H** 1.0 hydrogen																	2 **He** 4.0 helium
3 **Li** 6.9 lithium	4 **Be** 9.0 beryllium											5 **B** 10.8 boron	6 **C** 12.0 carbon	7 **N** 14.0 nitrogen	8 **O** 16.0 oxygen	9 **F** 19.0 fluorine	10 **Ne** 20.2 neon
11 **Na** 23.0 sodium	12 **Mg** 24.3 magnesium											13 **Al** 27.0 aluminium	14 **Si** 28.1 silicon	15 **P** 31.0 phosphorus	16 **S** 32.1 sulfur	17 **Cl** 35.5 chlorine	18 **Ar** 39.9 argon
19 **K** 39.1 potassium	20 **Ca** 40.1 calcium	21 **Sc** 45.0 scandium	22 **Ti** 47.9 titanium	23 **V** 50.9 vanadium	24 **Cr** 52.0 chromium	25 **Mn** 54.9 manganese	26 **Fe** 55.8 iron	27 **Co** 58.9 cobalt	28 **Ni** 58.7 nickel	29 **Cu** 63.5 copper	30 **Zn** 65.4 zinc	31 **Ga** 69.7 gallium	32 **Ge** 72.6 germanium	33 **As** 74.9 arsenic	34 **Se** 79.0 selenium	35 **Br** 79.9 bromine	36 **Kr** 83.8 krypton
37 **Rb** 85.5 rubidium	38 **Sr** 87.6 strontium	39 **Y** 88.9 yttrium	40 **Zr** 91.2 zirconium	41 **Nb** 92.9 niobium	42 **Mo** 96.0 molybdenum	43 **Tc** (98) technetium	44 **Ru** 101.1 ruthenium	45 **Rh** 102.9 rhodium	46 **Pd** 106.4 palladium	47 **Ag** 107.9 silver	48 **Cd** 112.4 cadmium	49 **In** 114.8 indium	50 **Sn** 118.7 tin	51 **Sb** 121.8 antimony	52 **Te** 127.6 tellurium	53 **I** 126.9 iodine	54 **Xe** 131.3 xenon
55 **Cs** 132.9 caesium	56 **Ba** 137.3 barium	57–71 lanthanoids	72 **Hf** 178.5 hafnium	73 **Ta** 180.9 tantalum	74 **W** 183.8 tungsten	75 **Re** 186.2 rhenium	76 **Os** 190.2 osmium	77 **Ir** 192.2 iridium	78 **Pt** 195.1 platinum	79 **Au** 197.0 gold	80 **Hg** 200.6 mercury	81 **Tl** 204.4 thallium	82 **Pb** 207.2 lead	83 **Bi** 209.0 bismuth	84 **Po** (210) polonium	85 **At** (210) astatine	86 **Rn** (222) radon
87 **Fr** (223) francium	88 **Ra** (226) radium	89–103 actinoids	104 **Rf** (261) rutherfordium	105 **Db** (262) dubnium	106 **Sg** (266) seaborgium	107 **Bh** (264) bohrium	108 **Hs** (267) hassium	109 **Mt** (268) meitnerium	110 **Ds** (271) darmstadtium	111 **Rg** (272) roentgenium	112 **Cn** (285) copernicium	113 **Nh** (280) nihonium	114 **Fl** (289) flerovium	115 **Mc** (289) moscovium	116 **Lv** (292) livermorium	117 **Ts** (294) tennessine	118 **Og** (294) oganesson

57 **La** 138.9 lanthanum	58 **Ce** 140.1 cerium	59 **Pr** 140.9 praseodymium	60 **Nd** 144.2 neodymium	61 **Pm** (145) promethium	62 **Sm** 150.4 samarium	63 **Eu** 152.0 europium	64 **Gd** 157.3 gadolinium	65 **Tb** 158.9 terbium	66 **Dy** 162.5 dysprosium	67 **Ho** 164.9 holmium	68 **Er** 167.3 erbium	69 **Tm** 168.9 thulium	70 **Yb** 173.1 ytterbium	71 **Lu** 175.0 lutetium
89 **Ac** (227) actinium	90 **Th** 232.0 thorium	91 **Pa** 231.0 protactinium	92 **U** 238.0 uranium	93 **Np** (237) neptunium	94 **Pu** (244) plutonium	95 **Am** (243) americium	96 **Cm** (247) curium	97 **Bk** (247) berkelium	98 **Cf** (251) californium	99 **Es** (252) einsteinium	100 **Fm** (257) fermium	101 **Md** (258) mendelevium	102 **No** (259) nobelium	103 **Lr** (262) lawrencium

The value in the brackets indicates the mass number of the longest-lived isotope.

2. Electrochemical series

Reaction	Standard electrode potential (E^0) in volts at 25°C
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87
$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$	+1.77
$Au^+(aq) + e^- \rightleftharpoons Au(s)$	+1.68
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$	+0.68
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$S(s) + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(g)$	+0.14
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	–0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	–0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	–0.25
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	–0.28
$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$	–0.40
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	–0.44
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	–0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	–0.83
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	–1.18
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	–1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	–2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	–2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	–2.87
$K^+(aq) + e^- \rightleftharpoons K(s)$	–2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	–3.04

3. Reactivity series for some metals

Element	Reactivity
K	most reactive ↑
Na	
Ca	
Mg	
Al	
(C)	
Zn	
Fe	
Sn	
Pb	
(H_2)	
Cu	
Ag	
Au	least reactive

4. Chemical relationships

Name	Formula
number of moles of a substance	$n = \frac{m}{M}$; $n = cV$; $n = \frac{V}{V_m}$
universal gas equation	$pV = nRT$
energy needed to heat a substance	$E = mc\Delta T$

5. Physical constants and standard values

Name	Symbol	Value
Avogadro constant	N_A or L	6.02×10^{23} mol^{-1}
molar gas constant	R	8.31 J mol^{-1} K^{-1}
molar volume of an ideal gas at SLC (25°C and 100 kPa)	V_m	24.8 L mol^{-1}
specific heat capacity of water	c	4.18 kJ kg^{-1} K^{-1} or 4.18 J g^{-1} K^{-1}
density of water at 25°C	d	997 kg m^{-3} or 0.997 g mL^{-1}
ionic product for water	K_W	1.00×10^{-14} mol^2 L^{-2} at 298 K (self-ionisation constant)

6. Unit conversions

Measured value	Conversion
0°C	273 K
100 kPa	750 mm Hg or 0.987 atm
1 litre (L)	1 dm^3 or 1×10^{-3} m^3 or 1×10^3 cm^3 or 1×10^3 mL

7. Metric (including SI) prefixes

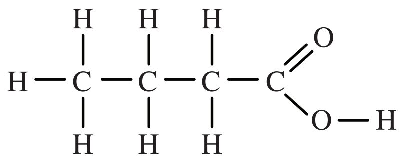
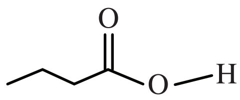
Metric (including SI) prefixes	Scientific notation	Multiplying factor
giga (G)	10^9	1 000 000 000
mega (M)	10^6	1 000 000
kilo (k)	10^3	1000
deci (d)	10^{-1}	0.1
centi (c)	10^{-2}	0.01
milli (m)	10^{-3}	0.001
micro (μ)	10^{-6}	0.000001
nano (n)	10^{-9}	0.000000001
pico (p)	10^{-12}	0.000000000001

8. Acid–base indicators

Name	pH range	Colour change from lower pH to higher pH in range
thymol blue (1st change)	1.2–2.8	red → yellow
methyl orange	3.1–4.4	red → yellow
bromophenol blue	3.0–4.6	yellow → blue
methyl red	4.4–6.2	red → yellow
bromothymol blue	6.0–7.6	yellow → blue
phenol red	6.8–8.4	yellow → red
thymol blue (2nd change)	8.0–9.6	yellow → blue
phenolphthalein	8.3–10.0	colourless → pink

9. Representations of organic molecules

The following table shows different representations of organic molecules, using butanoic acid as an example.

Formula	Representation
molecular formula	$C_4H_8O_2$
structural formula	
semi-structural (condensed) formula	$CH_3CH_2CH_2COOH$ or $CH_3(CH_2)_2COOH$
skeletal structure	

10. A solubility table

High solubility	Low solubility
Compounds containing the following ions are soluble in water. • Na^+, K^+, NH_4^+, NO_3^-, CH_3COO^- • Cl^-, Br^-, I^- (unless combined with Ag^+ or Pb^{2+}) • SO_4^{2-} (however $PbSO_4$ and $BaSO_4$ are not soluble, Ag_2SO_4 and $CaSO_4$ are slightly soluble)	Compounds containing the following ions are generally insoluble, unless combined with Na^+, K^+ or NH_4^+. • CO_3^{2-}, PO_4^{3-}, S^{2-} • OH^- ($Ba(OH)_2$ and $Sr(OH)_2$ are soluble, $Ca(OH)_2$ is slightly soluble)

END OF DATA BOOKLET