

FREE SAMPLE: THE FIRST QUESTION

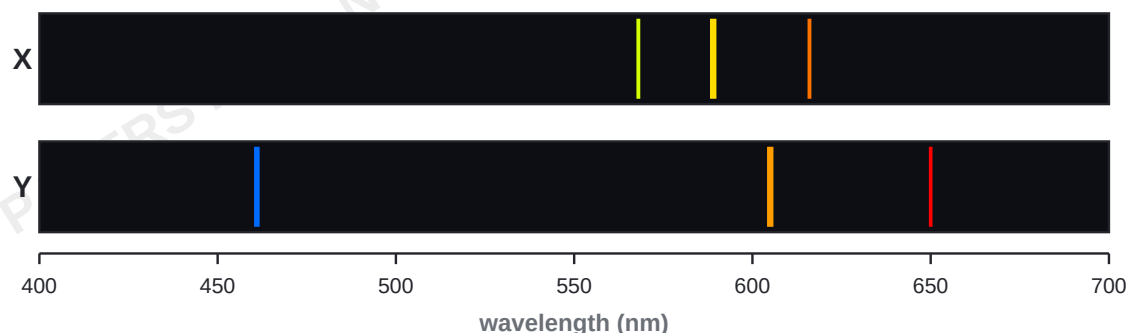
Question 1 (50 marks)

(a) The visible line emission spectra of salts of two elements, X and Y, are shown. The table gives the wavelengths of the three strongest visible lines for each of three elements.

- Explain how a line in an atomic emission spectrum is produced. Refer to the relationship $E_m - E_n = hf$ in your answer.
- Identify element X. Give a reason for your answer.
- Identify element Y. Give a reason for your answer.
- Calculate the frequency of the light of wavelength 589 nm. ($c = 3.00 \times 10^8 \text{ m s}^{-1}$)
- Calculate the energy of one photon of this light, and hence the energy of one mole of these photons in kJ mol^{-1} . ($h = 6.63 \times 10^{-34} \text{ J s}$; Avogadro constant = $6.02 \times 10^{23} \text{ mol}^{-1}$)

Element	Wavelengths of the three strongest lines (nm)
Sodium	568, 589, 616
Strontium	461, 605, 650
Copper	510, 522, 578

[10 marks]



(b) Enthalpy changes of combustion are measured in the laboratory, and Hess's law is used to find enthalpy changes that cannot be measured directly. The standard enthalpy changes of combustion are: carbon, $-393.5 \text{ kJ mol}^{-1}$; hydrogen, $-285.8 \text{ kJ mol}^{-1}$; ethanol, $-1367 \text{ kJ mol}^{-1}$.

(i) State Hess's law.

(ii) Use the data to calculate the enthalpy change of formation of ethanol, $\text{C}_2\text{H}_5\text{OH(l)}$, from its elements: $2\text{C(s)} + 3\text{H}_2\text{(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{C}_2\text{H}_5\text{OH(l)}$

(iii) A student burned ethanol in a spirit burner under a metal can of water and calculated a heat of combustion that was smaller (less exothermic) than the value above. Give two reasons for this difference.

[10 marks]

(c) The table gives the first ionisation energies (kJ mol^{-1}) of the elements in the third period.

(i) What is meant by the first ionisation energy of an element?

(ii) Identify the two elements in period 3 that do not follow the general increase in first ionisation energy across the period, and explain each exception.

(iii) The first four ionisation energies of an element Z are 738, 1451, 7733 and $10\,543 \text{ kJ mol}^{-1}$. In which group of the periodic table is Z? Explain your answer.

Na	Mg	Al	Si	P	S	Cl	Ar
496	738	578	786	1012	1000	1251	1521

[10 marks]

(d) Ethanol (boiling point 78 °C) and methoxymethane, CH₃OCH₃ (boiling point –24 °C), are structural isomers with the molecular formula C₂H₆O.

(i) Explain, in terms of intermolecular forces, why ethanol has a much higher boiling point than methoxymethane.

(ii) Draw a diagram to show the hydrogen bonding between two molecules of ethanol. Show the dipoles and the position of the hydrogen bond.

(iii) Ethanol mixes completely with water but is only slightly soluble in hexane. Explain both observations.

[10 marks]

(e) A student prepared a standard solution of sodium carbonate, Na₂CO₃ (relative molecular mass 106.0), to use in a titration.

(i) What is meant by a standard solution?

(ii) Calculate the mass of sodium carbonate needed to make 250.0 cm³ of a 0.100 mol L⁻¹ solution.

(iii) Describe how the student should prepare this solution accurately, naming the apparatus used.

[10 marks]

QUESTIONS 2 TO 7

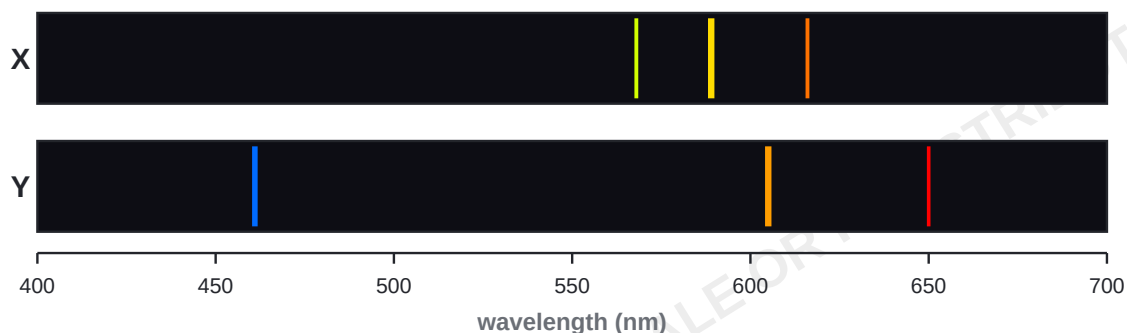
50 marks each · in the real paper you answer any FIVE

This practice paper prints all six so you can choose your best five, or attempt more than the real paper allows.

THE WORKED ANSWER TO THAT QUESTION

Question 1 (50 marks) [Atomic structure, energetics, bonding and a standard solution]

(a) The visible line emission spectra of salts of two elements, X and Y, are shown. The table gives the wavelengths of the three strongest visible lines for each of three elements. (i) Explain how a line in an atomic emission spectrum is produced. Refer to the relationship $E_m - E_n = hf$ in your answer. (ii) Identify element X. Give a reason for your answer. (iii) Identify element Y. Give a reason for your answer. (iv) Calculate the frequency of the light of wavelength 589 nm. ($c = 3.00 \times 10^8 \text{ m s}^{-1}$) (v) Calculate the energy of one photon of this light, and hence the energy of one mole of these photons in kJ mol^{-1} . ($h = 6.63 \times 10^{-34} \text{ J s}$; Avogadro constant = $6.02 \times 10^{23} \text{ mol}^{-1}$)



Line emission spectra of salts of X and Y.

(a) Answer: (i) Heat or an electric current excites electrons to higher energy levels; when an electron falls from a higher level (E_m) to a lower level (E_n) it emits a photon of energy $hf = E_m - E_n$, so each line has one fixed frequency. (ii) X is sodium: its lines are at 568, 589 and 616 nm. (iii) Y is strontium: its lines are at 461, 605 and 650 nm. (iv) $f = c/\lambda = 5.09 \times 10^{14} \text{ Hz}$. (v) $E = hf = 3.38 \times 10^{-19} \text{ J}$ per photon; per mole $3.38 \times 10^{-19} \times 6.02 \times 10^{23} = 2.03 \times 10^5 \text{ J} = 203 \text{ kJ mol}^{-1}$.

- Approach: link each spectral line to one electron transition, then match the line positions to the reference table, then use $f = c/\lambda$ and $E = hf$.
- (i) Excited electrons (heat or electricity supplied) occupy higher energy levels. When an electron drops from level m to a lower level n the energy difference is released as one photon, $E_m - E_n = hf$, where f is the photon's frequency. Only certain energy differences exist, so only certain frequencies, and therefore a few sharp lines, appear. 3 marks.
- (ii) The lines in spectrum X are at 568, 589 and 616 nm, which match the sodium row exactly. (iii) The lines in spectrum Y are at 461, 605 and 650 nm, which match strontium. 1 mark for each element, 1 mark for each reason.
- (iv) $\lambda = 589 \text{ nm} = 589 \times 10^{-9} \text{ m}$. $f = c/\lambda = 3.00 \times 10^8 \div 5.89 \times 10^{-7} = 5.09 \times 10^{14} \text{ Hz}$. 1 mark.
- (v) $E = hf = 6.63 \times 10^{-34} \times 5.09 \times 10^{14} = 3.38 \times 10^{-19} \text{ J}$ per photon (1). For one mole multiply by Avogadro's constant: $3.38 \times 10^{-19} \times 6.02 \times 10^{23} = 2.03 \times 10^5 \text{ J mol}^{-1} = 203 \text{ kJ mol}^{-1}$ (1).

KEY TAKEAWAY

Each emission line is one electron transition: $E_m - E_n = hf$, so a set of line positions is a fingerprint of the element.

(b) Enthalpy changes of combustion are measured in the laboratory, and Hess's law is used to find enthalpy changes that cannot be measured directly. The standard enthalpy changes of combustion are: carbon, $-393.5 \text{ kJ mol}^{-1}$; hydrogen, $-285.8 \text{ kJ mol}^{-1}$; ethanol, $-1367 \text{ kJ mol}^{-1}$. (i) State Hess's law. (ii) Use the data to calculate the enthalpy change of formation of ethanol, $\text{C}_2\text{H}_5\text{OH(l)}$, from its elements: $2\text{C(s)} + 3\text{H}_2\text{(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{C}_2\text{H}_5\text{OH(l)}$ (iii) A student burned ethanol in a spirit burner under a metal can of water and calculated a heat of combustion that was smaller (less exothermic) than the value above. Give two reasons for this difference.

(b) Answer: (i) The enthalpy change for a reaction is independent of the route taken (it depends only on the initial and final states). (ii) $\Delta H_f = 2(-393.5) + 3(-285.8) - (-1367) = -787.0 - 857.4 + 1367 = -277.4 \text{ kJ mol}^{-1}$. (iii) Heat lost to the surroundings/air and the can; incomplete combustion (soot/CO formed); evaporation of ethanol before burning; heat not all transferred to the water.

- Approach: build a cycle in which the elements and ethanol both burn to the same products (CO_2 and H_2O).
- (i) Hess's law: the total enthalpy change of a reaction is the same whichever route is taken, because enthalpy is a state function. 2 marks.
- (ii) Burning the elements: $2\text{C} \rightarrow 2\text{CO}_2$ releases $2 \times (-393.5) = -787.0 \text{ kJ}$; $3\text{H}_2 \rightarrow 3\text{H}_2\text{O}$ releases $3 \times (-285.8) = -857.4 \text{ kJ}$. Burning the ethanol to the same products releases -1367 kJ , so going from elements to ethanol directly = (route through combustion of the elements) – (combustion of ethanol): $-787.0 - 857.4 - (-1367) = -277.4 \text{ kJ mol}^{-1}$. 5 marks (cycle 2, arithmetic 2, sign and units 1).
- (iii) Any two: some heat escapes to the air and to the metal can, so not all of it is absorbed by the water; incomplete combustion gives CO and soot, which release less energy; ethanol evaporates from the wick. 1.5 marks each, 3 in total.

KEY TAKEAWAY

Hess's law lets you add or subtract measured combustion enthalpies to find a formation enthalpy: $\Delta H = \Sigma(\text{combustion of reactants}) - \Sigma(\text{combustion of products})$.

(c) The table gives the first ionisation energies (kJ mol^{-1}) of the elements in the third period. (i) What is meant by the first ionisation energy of an element? (ii) Identify the two elements in period 3 that do not follow the general increase in first ionisation energy across the period, and explain each exception. (iii) The first four ionisation energies of an element Z are 738, 1451, 7733 and $10\,543 \text{ kJ mol}^{-1}$. In which group of the periodic table is Z? Explain your answer.

(c) Answer: (i) The energy needed to remove one mole of electrons from one mole of gaseous atoms (the outermost electron from each atom), to form one mole of gaseous $1+$ ions. (ii) Aluminium ($578 < 738$ for Mg): the 3p electron is higher in energy and better shielded by the 3s electrons than the 3s electron of magnesium. Sulfur ($1000 < 1012$ for P): in sulfur two electrons are paired in one 3p orbital, and the repulsion between them makes one easier to remove than the unpaired 3p electron of phosphorus. (iii) Group 2: there is a big jump between the 2nd (1451) and 3rd (7733) ionisation energies, so the atom has two outer electrons.

- Approach: look for places where the trend falls, then use sublevel and electron-pairing arguments.
- (i) The minimum energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of gaseous positive ions ($1+$). 2 marks.
- (ii) The values go Na 496, Mg 738, Al 578 (drop), Si 786, P 1012, S 1000 (drop), Cl 1251, Ar 1521. Al: its outer electron is in a 3p sublevel, which has higher energy than 3s and is shielded by the filled 3s sublevel, so it is easier to remove than Mg's 3s electron (3 marks). S: its fourth 3p electron must pair up in an orbital with another electron; the electron–electron repulsion makes it easier to remove than the unpaired 3p electron in P (3 marks).

- (iii) The successive values are 738, 1451, 7733 and 10543 kJ mol⁻¹. The first two are low, then the third is more than five times the second, so the third electron is being removed from a complete inner shell. The atom therefore has two outer electrons: Group 2 (in fact magnesium). 2 marks.

KEY TAKEAWAY

Dips in the first ionisation energy across a period come from a new sublevel (Al) or from electron pairing (S); a big jump in successive values shows how many outer electrons there are.

(d) Ethanol (boiling point 78 °C) and methoxymethane, CH₃OCH₃ (boiling point -24 °C), are structural isomers with the molecular formula C₂H₆O. (i) Explain, in terms of intermolecular forces, why ethanol has a much higher boiling point than methoxymethane. (ii) Draw a diagram to show the hydrogen bonding between two molecules of ethanol. Show the dipoles and the position of the hydrogen bond. (iii) Ethanol mixes completely with water but is only slightly soluble in hexane. Explain both observations.

(d) Answer: (i) Ethanol has an O–H group so its molecules form hydrogen bonds with each other (strong permanent dipole–dipole attractions, as well as London forces); methoxymethane has no O–H, so it has only weaker dipole–dipole and London forces. More energy is needed to separate ethanol molecules. (ii) Two ethanol molecules with δ+ on the H of the O–H group and δ– on the O of the neighbouring molecule, and a dotted line between the hydrogen of one molecule and the oxygen of the other (O–H···O). (iii) Ethanol forms hydrogen bonds with water (and the molecules are polar), which replace the hydrogen bonds between water molecules, so it dissolves. Hexane is non-polar and has only London forces with ethanol, so it does not dissolve well.

- Approach: identify the strongest intermolecular force in each liquid and link it to the energy needed to separate the molecules.
- (i) A hydrogen bond needs hydrogen bonded directly to a very electronegative atom (O). Ethanol has that O–H bond; methoxymethane has none. Hydrogen bonds are the strongest of the van der Waals forces here, so ethanol needs more energy to boil. 4 marks.
- (ii) Draw C₂H₅–Oδ––Hδ+ twice. Put a dotted line from the Hδ+ of molecule 1 to the Oδ– of molecule 2, marked 'hydrogen bond', with the O–H···O atoms in a line. 3 marks (dipoles 1, labelled bond 1, correct atoms 1).
- (iii) Like dissolves like. Ethanol's –OH group forms hydrogen bonds with water, so the energy needed to separate water molecules is paid back (2 marks). Hexane molecules are non-polar and cannot form hydrogen bonds, and the ethyl chain of ethanol only forms weak London forces with hexane (1 mark).

KEY TAKEAWAY

Hydrogen bonding (O–H···O) explains the higher boiling point of an alcohol than its ether isomer and its solubility in water.

(e) A student prepared a standard solution of sodium carbonate, Na₂CO₃ (relative molecular mass 106.0), to use in a titration. (i) What is meant by a standard solution? (ii) Calculate the mass of sodium carbonate needed to make 250.0 cm³ of a 0.100 mol L⁻¹ solution. (iii) Describe how the student should prepare this solution accurately, naming the apparatus used.

(e) Answer: (i) A solution whose concentration is known accurately. (ii) $n = 0.100 \times 0.2500 = 0.0250$ mol; mass = $0.0250 \times 106.0 = 2.65$ g. (iii) Weigh about 2.65 g of dry Na₂CO₃ accurately on a balance in a clean weighing bottle; dissolve it in a beaker in a little deionised water; transfer the solution and the rinsings into a 250 cm³ volumetric flask using a funnel; make up to the mark with deionised water (last drops with a dropper) so the bottom of the meniscus is on the line; stopper and invert several times to mix.

- Approach: moles from concentration and volume, then mass from moles and M_r ; then describe the careful transfer so no solute is lost.
- (i) A standard solution is one whose concentration is accurately known (for example made from a primary standard). 2 marks.
- (ii) Moles = concentration $\times$ volume = $0.100 \text{ mol L}^{-1} \times 0.2500 \text{ L} = 0.0250 \text{ mol}$. Mass = moles $\times M_r = 0.0250 \times 106.0 = 2.65 \text{ g}$. 3 marks.
- (iii) Weigh the solid on a balance (record the exact mass); dissolve in a beaker with deionised water; funnel into a 250 cm^3 volumetric flask; rinse the beaker, glass rod and funnel into the flask; add deionised water to the line with a dropper; stopper and invert to mix. 5 marks (weighing 1, dissolve 1, transfer and rinse 1, make up to mark 1, mix 1).

KEY TAKEAWAY

A standard solution is made by weighing a primary standard accurately, dissolving it, transferring every drop to a volumetric flask, and making up to the mark.

LOS: Atomic structure, energetics, bonding and a standard solution (Leaving Certificate Chemistry, Higher Level written paper, Strands 1, 2, 3)

QUESTIONS 2 TO 7

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