Answers to examination questions

1 Quantitative chemistry

Paper 1 IB questions and IB style questions

Q1 D One mole of a chemical element is the relative atomic mass expressed in grams.

Q2 D Mass of one molecule $= \dfrac{(12 \times 3) + (8 \times 1)\,\text{g mol}^{-1}}{6 \times 10^{23}\,\text{mol}^{-1}} = 7.33 \times 10^{-23}\,\text{g}$

Q3 B $28\,\text{g mol}^{-1} \times 2\,\text{mol} = 56\,\text{g}$

Q4 D A: 0.5 mol H_2 molecules; 1 mol H atoms, B: 0.03125 mol O_2 molecules; 0.0625 mol oxygen atoms; C: 3.9×10^{-3} mol S_8 molecules; 0.0312 mol S atoms; D: 6.258×10^{-3} mol Br_2 molecules; 0.0125 mol Br atoms

Q5 C Amount $= \dfrac{\text{mass(g)}}{\text{molar mass (g mol}^{-1})} = \dfrac{180\,\text{g}}{18\,\text{g mol}^{-1}} = 10\,\text{mol}$

Number of molecules $= 6 \times 10^{23} \times 10 = 6 \times 10^{24}$

Q6 D Multiply the equation through by four: 12 moles of hydrogen peroxide would produce 8 moles of water.

Q7 C $2H_2S + 3O_2 \rightarrow 2SO_2 + 2H_2O$

Q8 D M_r $BH_3 = 14$, molecular formula could be B_4H_{12} ($M_r = 56$); B: $M_r = 56$; C: $M_r = 56$; D: $M_r = 36.5$

Q9 C The molecular formulas simplify to NF, CH_2O, C_2H_4O (unchanged) and CN.

Q10 A $2Cl_2$ represents two chlorine molecules: Cl–Cl and Cl–Cl.

Q11 C $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$ (balanced with correct formulas).

Q12 D The charge is not balanced.

Q13 A $Fe_2(SO_4)_3$ is composed of three SO_4^{2-} and two Fe^{3+}.

Q14 B Amount of M atoms $= \dfrac{(52 - 32)}{40} = 0.50\,\text{mol}$;

amount of S atoms $= \dfrac{32}{32} = 1.0\,\text{mol}$.

Multiplying by two gives MS_2.

Q15 D $M + O_2 \rightarrow MO_2$

Amount of oxygen molecules $= \dfrac{16}{32} = 0.5\,\text{mol}$ O_2 molecules;

0.5 mol M atoms.

Hence relative atomic mass of M $= \dfrac{36}{0.5} = 72\,\text{g mol}^{-1}$.

Q16 B $2H_2 + O_2 \rightarrow 2H_2O$; 8 mol H_2 and 2 mol O_2; stoichiometric molar ratio is 2 : 1, hence the gas mixture contains unreacted hydrogen and water – no oxygen, since it is the limiting reagent.

Q17 B $3Mg(s) + Fe_2O_3(s) \rightarrow 2Fe(s) + 3MgO(s)$

Amount of magnesium $= \dfrac{2.4\,\text{g}}{24\,\text{g mol}^{-1}} = 0.1\,\text{mol}$;

amount of iron $= \dfrac{0.1\,\text{mol} \times 2}{3}$.

Mass of iron $= 0.066\,\text{mol} \times 56\,\text{g mol}^{-1} = 3.733\,\text{g}$.

Q18 A Molecular mass $= \dfrac{63 \times 12}{55.74 \times 100} = 1356.3\,\text{g mol}^{-1}$

Q19 D Amount of $NH_3 = \dfrac{8.5}{17} = 0.5\,\text{mol}$;

amount of NO $= \dfrac{15}{30} = 0.5\,\text{mol}$;

percentage yield $= \dfrac{0.5}{0.5} \times 100 = 100\%$

Q20 D There is range of kinetic energies for gas molecules in a sample at fixed temperature.

Q21 C Avogadro’s hypothesis: the number of particles in a gas is directly proportional to volume at the same temperature and pressure.

Q22 D Average kinetic energy of gas particles is directly proportional to the absolute temperature in kelvin.

Q23 D High pressure will bring the gas molecules very close together. This causes more collisions and also allows the weak attractive forces to operate. With low temperatures, the gas molecules do not have enough kinetic energy to continue on their path to avoid that attraction.

Q24 A *P* and *V* vary inversely at constant temperature. This relationship is termed Boyle’s law.

Q25 D The average kinetic energy of the gas molecules is directly proportional to absolute temperature (a result of thermodynamics). At a given temperature the molecules of all species of gas, no matter what size, shape or mass, have the same average kinetic energy.

Q26 A $V_2 = \dfrac{350\,\text{cm}^3 \times 99.3\,\text{kPa} \times 273\,\text{K}}{101.3\,\text{kPa} \times 295\,\text{K}}$

$V_2 = 317.5034\,\text{cm}^3 = 318\,\text{cm}^3$

Q27 B Amount of $CO_2 = \dfrac{88.0}{44} = 2\,\text{mol}$, hence amount of carbon atoms = 2 mol; amount of $H_2O = \dfrac{27}{18} = 1.5\,\text{mol}$ hence amount of hydrogen atoms = 3 mol; empirical formula is C_2H_3, hence possible molecular formulas include C_4H_6, C_6H_9, etc.

Q28 B 4 vol + 5 vol → 10 vol;
experimental data: 1 vol + 1.25 vol → 2.5 vol;
multiplying through by four gives: 4 vol + 5 vol → 10 vol

Q29 C $M_1 \times V_1 = M_2 \times V_2$

$0.600\,\text{mol dm}^{-3} \times 40.0\,\text{cm}^3 = 0.100\,\text{mol dm}^{-3} \times V_2$;
$V_2 = 240\,\text{cm}^3$; $V_2 - V_1 = 240\,\text{cm}^3 - 40\,\text{cm}^3$
$= 200\,\text{cm}^3$

Q30 D $Fe(NO_3)_3(aq) \rightarrow Fe^{3+}(aq) + 3NO_3^-(aq)$;
nitrate concentration $= (0.6 \times 3) = 1.8\,\text{mol dm}^{-3}$.

Q31 B $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$; amount of NaOH = 0.01 mol; amount of $H_2SO_4 = 5 \times 10^{-3}\,\text{mol}$;

amount $= \dfrac{\text{volume (cm}^3)}{1000} \times \text{molarity}$;

molarity $= 1000 \times 0.01 = 0.2817\,\text{mol dm}^{-3}$.

Q32 A $AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$

Amount of KCl(aq) $= \dfrac{55}{1000} \times 0.200 = 0.011\,\text{mol}$;

hence yield of AgCl = 0.011 mol (1 : 1 molar ratio).

Mass of AgCl $= 0.011\,\text{mol} \times 143.32\,\text{g mol}^{-1} = 1.57\,\text{g}$

Q33 C $V_1/T_1 = V_2/T_2$; $V_2/V_1 = T_2/T_1$; 473 K/373 K = 1.268

Q34 A $n = \dfrac{pV}{RT}$

$$= \frac{101\,000\,\text{Pa} \times 1.15 \times 10^{-3}\,\text{m}^3}{8.31\,\text{J mol}^{-1}\,\text{K}^{-1} \times 273\,\text{K}} = 0.0512\,\text{mol.}$$

$$\text{Molar mass} = \frac{3.96\,\text{g}}{0.0512\,\text{mol}} = 77.34\,\text{g mol}^{-1}.$$

Q35 D All properties will increase in value except density which remains constant since the mass and volume of gas remain constant.

Paper 2 IB questions and IB style questions

Q1 a $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$ [2]

b Carbon dioxide produced = 20 cm^3 [1]

Oxygen remaining = 10 cm^3 [1]

Q2 a $Zn(s) + I_2(s) \rightarrow ZnI_2(s)$ [1]

b Amount of zinc $= \dfrac{100\,\text{g}}{65.37\,\text{g mol}^{-1}} = 1.530\,\text{mol}$ [1]

Amount of iodine $= \dfrac{100\,\text{g}}{253.8\,\text{g mol}^{-1}} = 0.3940\,\text{mol}$ [1]

The reacting molar ratio is 1 : 1 therefore the zinc is present in excess. [1]

c Amount of zinc iodide = amount of iodine used

Amount of zinc iodide $= \dfrac{100}{253.8} = 0.3940\,\text{mol}$

Mass of zinc iodide

$= 0.3940 \times (65.37 + 253.8) = 125.8\,\text{g}$ [1]

Q3 a $\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}$; $\dfrac{955\,\text{cm}^3}{278\,\text{K}} = \dfrac{V_2}{303\,\text{K}}$; $V_2 = 1041\,\text{cm}^3$

The balloon will burst. [3]

b The increase in temperature increases the average kinetic energy of the nitrogen molecules and hence increases their average speed. They collide more frequently and energetically with each other and the walls of the balloon. [2]

Q4 a Amount of hydrochloric acid $= \dfrac{13.95\,\text{dm}^3}{1000} \times 2.00\,\text{mol dm}^{-3} = 0.0279\,\text{mol}$ [1]

b Amount of sodium carbonate in 20.00 cm^3 of solution = 0.01395 mol [1]

c Concentration of sodium carbonate $= \dfrac{1000\,\text{dm}^{-3}}{20} \times 0.01395\,\text{mol}$

$= 0.6975\,\text{mol dm}^{-3}$ [1]

d Molar mass of hydrated sodium carbonate $= \dfrac{200.00\,\text{g}}{0.6975\,\text{mol}}$

$= 286.74\,\text{g mol}^{-1}$ [1]

e Molar mass of anhydrous sodium carbonate, Na_2CO_3 = $(22.99 \times 2) + 12.01 + (3 \times 16.00) = 105.99\,\text{g mol}^{-1}$

Molar mass of water $= 286.74\,\text{g mol}^{-1} - 105.99\,\text{g mol}^{-1}$

$= 180.75\,\text{g mol}^{-1}$ [1]

Hence, $n = \dfrac{180.75\,\text{g mol}^{-1}}{18.02\,\text{g mol}^{-1}} = 10$ [1]

Q5 a $3Ag^+(aq) + XO_4^{3-}(aq) \rightarrow Ag_3XO_4(s)$ [1]

b i Amount of silver ions = $0.2040\,\text{mol dm}^{-3} \times 0.04118\,\text{dm}^3 = 8.401 \times 10^{-3}\,\text{mol}$ [1]

ii Amount of $Ag_3XO_4 = \frac{1}{3} \times 8.401 \times 10^{-3}\,\text{mol}$

$= 2.800 \times 10^{-3}\,\text{mol}$ [1]

iii 2.800×10^{-3} mol of Ag_3XO_4 weighs 1.172 grams.

Hence one mole weighs $\dfrac{1.172\,\text{g}}{2.800 \times 10^{-3}\,\text{mol}}$ $= 418.6\,\text{g mol}^{-1}$. [2]

iv $(3 \times 107.87) + x + 4\,(16.0) = 4\,18.6$; hence $x = 30.99$ and X is phosphorus (P) [2]

2 Atomic structure

Paper 1 IB questions and IB style questions

Q1 D Isotopes are two or more atoms of the same element with different numbers of neutrons (and therefore different relative isotopic masses). They have the same numbers of protons and hence the same number of electrons.

Q2 C H_2S: $[2H^+\ X^{2-}]$

Q3 A

i	1 p	1 e	0 n
ii	17 p	18 e	18 n
iii	19 p	18 e	20 n

Q4 D Emission spectra provide evidence that electrons exist in fixed energy levels within an atom.

Q5 C The atomic number of sulfur is 16; it is in group 6 and therefore has six valence electrons.

Q6 D A discontinuous spectrum is produced.

Q7 C 9 protons (from periodic table); 10 electrons in atom (due to the single negative charge) and 10 neutrons (19 – 9).

Q8 B $1 \rightarrow \frac{1}{2} \rightarrow \frac{1}{4} \rightarrow \frac{1}{8}$

Q9 C To account for the emission or line spectrum of hydrogen atoms.

Q10 C Relative atomic mass is a weighted average. The more abundant isotope has a mass number of 85.

Q11 A The element is silicon with the electron arrangement 2,8,4.

Q12 B $^{91}_{39}Y$: 52 neutrons $^{87}_{37}Rb$: 52 neutrons $^{89}_{38}Sr$: 51 neutrons $^{84}_{36}Kr$: 48 neutrons.

Q13 D They converge at high energies; lines are produced when electrons move from higher to lower energies.

Q14 D Relative atomic mass of chlorine $= \dfrac{(35 \times 3) + (37 \times 1)}{3 + 1}$

$= \dfrac{142}{4} = 35.5$

Q15 D Electrons release energy as they move from high to low energy levels: light is released.

Q16 A Energy levels rapidly converge with an increase in n.

Q17 A The thyroxine secreted by the thyroid is an iodine-containing compound.

Q18 B I. 157 neutrons; II. 156 neutrons; III. 158 neutrons and IV. 156 neutrons.

Q19 A Atoms of isotopes have the same number of protons have the same number of protons but different numbers of neutrons. The mass number is the number of protons and neutrons.

Q20 A The sample is first vaporized by lowering the surrounding pressure: a mass spectrometer can only accept gaseous atoms or molecules. The molecules or atoms are then subjected to bombardment by high energy electrons which remove a valence electron to form unipositive ions. The ions are then accelerated by an electric field before entering a powerful magnetic field where they experience deflection according to their mass-to-charge ratio.

Paper 2 IB questions and IB style questions

Q1 **a** [3]

	An atom of ^{79}Br	An ion of $^{81}Br^-$
Protons	35	35
Neutrons	44	46
Electrons	35	36

b ^{79}Br because its relative isotopic is closer to 79.90 [1]

Q2 **a** Atoms of the same element but with different mass numbers and hence different number of neutrons. [1]
b 47 protons, 46 electrons and 60 neutrons [1]
c Carbon-12 (^{12}C). [1]

Q3 **a** **i** Ionization, acceleration [2]
ii The gaseous atoms formed from the sample are bombarded with electrons travelling at high speed with large kinetic energies. [2]
iii $^{191}Ir^+$ [1]
It has the lowest mass-to-charge ratio (or lowest mass since the charges are identical). [1]
b **i** The weighted average of relative isotopic masses of all the stable isotopes of an element relative to the mass of one atom of carbon-12. [2]
ii $(191 \times 0.371) + (193 \times 0.629) = (70.861 + 121.397) = 192.258 = 192.26$ (2 d.p.) [2]
iii An unstable isotope of an element that decays or disintegrates spontaneously, emitting nuclear or ionizing radiation. [1]
Cobalt-60, molybdenum-99, chromium-51, iron-59 and iodine-125 [1]

Q4 The emission spectrum consist of a series of sharp or discrete coloured lines on a black background. [1]

The lines converge together at high energy (high frequency). The lines are generated when excited electrons move from high energy levels to lower energy levels. [1]

Light of a particular frequency is released during this process.[1]

3 Periodicity

Paper 1 IB questions and IB style questions

Q1 **A** Strontium and calcium are both in group 2.

Q2 **D** Electronegativity increases across a period from left to right. Reactivity decreases down group 7.

Q3 **A** Reactive metals form basic oxides; reactive non-metals form acidic oxides.

Q4 **C** Chlorine oxidizes the iodide ions:
$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$

Chlorine is a more powerful oxidizing agent than iodine.

Q5 **C** The elements in the periodic table are arranged in order of increasing atomic number.

Q6 **B** Element 20 is calcium and is in group 2. It has the electron arrangement 2,8,8,2.

Q7 **B** Atomic radii decrease across a period due to the increase in nuclear charge. Additional electrons enter the same shell and only a small increase in shielding occurs.

Q8 **A** Reactivity increases down group 1.

Q9 **A** Removal of successive electrons increases the nuclear charge experienced by the remaining electrons. Addition of successive electrons increases the electron–electron repulsion and decreases the nuclear charge experienced by the electrons.

Q10 **A** Most non-metals have low melting points because they are simple molecular substances.

Q11 **C** Each member of a group has the same number of valence electrons and hence form ions with the same charge or the same number of covalent bonds.

Q12 **C** Atomic radii decrease across a period. Each subsequent electron enters the same shell, the small increase in shielding is more than outweighed by the increase in nuclear charge.

Q13 **C** The removal of an electron from an iodine atom results in an increase in the nuclear charge experienced by the remaining electrons and hence the radius decreases. The addition of an electron increases the electron–electron repulsion and decreases the nuclear charge experienced by the electrons.

Q14 **A** Each member of a group has one more electron shell than the element above it.

Q15 **B** Atomic radii increase down group 7 due to the presence of additional electron shells. Melting points increase down the group due to an increase in van der Waals' forces caused by additional electrons.

Q16 **B** Shielding increases slightly across a period, but this is more than outweighed by the increase in nuclear charge.

Q17 **D** Aluminium oxide and silicon dioxide are insoluble; sodium oxide forms sodium hydroxide solution (basic). Sulfur trioxide forms sulfuric acid solution.

Q18 **D** Atomic radii increases down any group due to the presence of additional electron shells.

Q19 **B** Ionization energies decrease down a group and increase a period.

Q20 **C** The atomic radius of barium is larger than that of strontium. It has an additional shell of electrons.

Paper 2 IB questions and IB style questions

Q1 **a** **i** The minimum energy required for the removal of a mole of electrons from a mole of gaseous atoms to form a mole of unipositive gaseous ions. [2]
ii $Al(g) \rightarrow Al^+(g) + e^-$ [1]
b The magnesium atom has a greater nuclear charge due a higher number of protons. Consequently, the atomic radius is smaller and there is a stronger attraction for the electrons by the nucleus. [3]
c $2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$
There will be effervescence: release of gas at the surface. Bubbles of a colourless gas will be observed. The lithium moves around the surface and is converted into a soluble product. Heat and sound energy are released. [3]

Q2 a i The bromine atom has a greater number of protons and hence a greater nuclear charge: 35+ versus 34+. Hence the outer or valence electrons are attracted more strongly. The bromide ion is formed by the addition of one electron to the bromine atom; the selenide ion is formed by the addition of two electrons to the selenium atom. The additional electrons cause electron-electron repulsion which increase radii. The effect is highest in the selenide ion. [2]

ii The electronegativity is the ability of atom to attract bonding pair of electrons in a covalent bond. Fluorine has a smaller radius and hence the electrons are closer to a nucleus in a lower energy level. The electrons are repelled by fewer inner electrons and there is a decreased shielding effect. [3]

b i Brown solution formed; $Cl_2 + 2KI \rightarrow I_2 + 2KCl$; $Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-$ [2]

ii A cream or off-white precipitate (solid) is formed: $Ag^+ + Br^- \rightarrow AgBr$ [2]

c The reaction in **i** is redox. [1]
The chlorine is reduced/gains electrons/decreases its oxidation number. [1]
The iodide ion is oxidized/loses electrons/increases its oxidation number. [1]
The reaction in **ii** is not redox; there is no electron transfer/change in oxidation number. [1]

Q3 a The size of an atom is determined by the attraction of the nucleus for electrons and the shielding of outer electrons from the nucleus by inner shells of electrons. [3]

b i This is because the sodium ion, Na^+, has one less shell of electrons compared to the sodium atom. Consequently, there is a decrease in shielding and an increase in the nuclear charge experienced by each electron. [2]

ii This is because with increasing proton number, each successive positive ion contains an additional shell of electrons. The additional screening effect outweighs the effect of increased nuclear charge. [2]

c The magnesium ion has 12 protons and 10 electrons; the sodium ion has 11 protons and 10 electrons. The nuclear charge of the magnesium ion (+12) is greater than that of the sodium ion (+11). Consequently, all the electrons in the magnesium ion will experience a greater force and be located nearer to the nucleus. [2]

d i $N < N^{3-}$ [1]

ii $Fe^{3+} < Fe^{2+} < Fe$ [1]

Q4 The ions formed across period 3 would be Na^+, Mg^{2+}, Al^{3+} and P^{3-}, S^{2-} and Cl^-.

There is a decrease in ionic radii from Na^+ to Al^{3+}: all the ions have the electron arrangement of 2,8 (that is they, are isoelectronic), however, there is a progressive increase in the nuclear charge due to the additional protons: the sodium ion has 10 protons, the magnesium ion has 11 protons and the aluminium ion has 13 protons. All of the electrons experience a higher effective nuclear charge and the electrons in the two shells are pulled progressively towards the nucleus. [3]

There is a decrease in ionic radii from P^{3-}, S^{2-} to Cl^-: all the ions have the electron arrangement of 2,8,8 (that is, they are isoelectronic), however, there is a progressive increase in the nuclear charge due to the additional protons: the phosphide ion has 15 protons, the sulfide ion has 16 protons and the chloride ion has 18 protons. (All of the electrons experience a higher effective nuclear charge and the electrons in the three shells are pulled progressively towards the nucleus). [3]

4 Bonding

Paper 1 IB questions and IB style questions

Q1 A Magnesium is a metal; chlorine is a non-metal. Typically, a metal and a non-metal react to form an ionic compound.

Q2 A All four molecules are based upon a tetrahedral arrangement of four regions of high electron density. However, lone pairs cause more repulsion than bonding pairs. The oxygen in water is surrounded by two lone pairs (maximum repulsion), the nitrogen atom in ammonia contains one lone pair, but the carbon in methane has no lone pairs (least repulsion).

Q3 B H_2S has a higher molecular mass than H_2O and hence more extensive van der Waals' forces. However, water is hydrogen bonded. Hydrogen bonds are significantly stronger than van der Waals' forces (for molecules of similar molecular mass). This more than outweighs the increase in van der Waals' forces from H_2O to H_2S.

Q4 B X will form X^+ and Y will form Y^-. Hence the formula will be XY.

Q5 D The oxygen molecule (O=O) contains one double bond; the carbon dioxide molecule (O=C=O) contains two double bond and the tetrafluoroethene molecule contains one double bond.

Q6 A B is trigonal planar (bond angles 120°); A, C and D are based upon a tetrahedral arrangement with four regions of high electron density. However, the water molecule has two lone pairs around the central atom, compared with one for ammonia and none for methane. Lone pair repulsion is greater than bonding pair repulsion.

Q7 A Hydrogen bonding would only be expected to occur in CH_2CH_3COOH since it contains hydrogen bonded directly to oxygen. The –OH group is able to engage in hydrogen bonding. In the other two compounds the hydrogen is bonded to carbon.

Q8 C Both ethane and neon are non-polar. Hence, the predominant intermolecular and inter-atomic forces are van der Waals' forces whose size depends on the number of electrons present in the atom or molecule. A greater number of electrons in the ethane molecule results in a stronger temporary dipole which makes neighbouring molecules or atoms attract each other more strongly. This means more heat is required to overcome the intermolecular or inter-atomic forces in ethane molecules, resulting in a higher boiling point.

Q9 B Fluorine is non-polar: the two fluorine atoms have equal electronegativity values. Tetrafluoromethane is tetrahedral and hence non-polar. Hydrogen iodide and hydrogen fluoride are both polar owing to the presence of a polar bond and linear shape. However, fluorine is the most electronegative element and hence the bond has the greatest polarity.

Q10 C D is a description of covalent bonding; A is a description of ionic bonding.

Q11 **C** Ethane and propane are hydrocarbons and hence unable to form hydrogen bonds with water. Propan-1-ol and hexan-1-ol are both alcohols and contain a hydroxyl group, –OH, capable of forming hydrogen bonds with water molecules. However, propan-1-ol is the more soluble since its hydrocarbon chain is smaller: three carbon atoms compared with six in hexan-1-ol.

Q12 **B** B is a description of ionic bonding.

Q13 **B** X will form X^{3+}; Y will form Y^{2-} and hence the formula will be X_2Y_3 [$2X^{3+}$ $3Y^{2-}$].

Q14 **C** Tetrafluoroethene, C_2F_4, contains a carbon–carbon double bond.

Q15 **D** All of the hydrogen atoms are attached to carbon atoms and hence are unable to engage in hydrogen bonding.

Q16 **D** Chlorine has the lower boiling point because it has the lower molecular mass. A lower molecular mass means less electrons are available for temporary dipole formation. Hence, the van der Waals' forces of attraction are weaker.

Q17 **B** C_2H_4 has a trigonal planar arrangement around the two carbon atoms.

Q18 **A** Metals conduct electricity in the solid state. Dissolving releases mobile ions from soluble ionic compounds allowing them to conduct electricity and undergo electrolysis.

Q19 **A** Each oxygen atom has two lone pairs, so there are eight bonds: 4 × C–H; 2 × C–O, 1 × C=O.

Q20 **B** The molecule has a carbon–carbon double bond. There are three regions of high electron density around the two carbons atom. Hence, the molecule is trigonal planar and has bond angles of 120°.

Q21 **A** Pentane, C_5H_{12}, will have the lowest boiling point. It is a non-polar substance and its molecules are associated in the liquid state via van der Waals' forces – the weakest of the three intermolecular forces. $CH_3OCH_2CH_2CH_3$ is associated in the liquid state via dipole–dipole forces. Propan-1-ol has the highest boiling point since it is capable of hydrogen bonding owing to the presence of a hydroxyl group, –OH.

Q22 **D** Strontium forms Sr^{2+} and nitrogen form N^{3-}. Hence the formula of strontium nitride is Sr_3N_2 [$3Sr^{2+}$ $2N^{3-}$].

Q23 **D** Ethane, C_2H_6, has the lowest boiling point owing to the presence of van der Waals' forces of attraction – the weakest of the three intermolecular forces. Fluoromethane, CH_3Cl, has an intermediate boiling point owing to the presence of dipole–dipole forces. Methanol, CH_3OH, has the highest boiling point owing to the presence of hydrogen bonding – the strongest of the three intermolecular forces.

Q24 **C** Carbon dioxide is a non-polar molecule due to its linear shape: the two dipoles cancel each other out. Hence, the forces operating in the solid are van der Waals' forces.

Q25 **D** Carbon dioxide is a linear molecule. The central carbon atom is surrounded by two regions of high electron density: two double bonds.

Q26 **A** Water is hydrogen bonded. Hydrogen bonds are significantly stronger than van der Waals' forces of attraction, present in H_2Se and H_2S, and dipole–dipole forces of attraction present in HCl.

Q27 **A** Carbon-60 is a simple covalent substance.

Q28 **D** The relative order of strength is: van der Waals' forces, hydrogen bonds, covalent bonds.

Q29 **A** The largest difference in electronegativity values is for oxygen and hydrogen atoms.

Q30 **B** The nitrogen atom in nitrogen trichloride is surrounded by three bonding pairs and one lone pair.

Paper 2 IB questions and IB style questions

Q1 Sodium, magnesium and aluminium are metallic and hence are good electrical conductors due to the presence of delocalized valence electrons. From sodium, magnesium to aluminium, the atomic size decrease and the number of valence electrons available for delocalization increases from one to three. Thus, the metallic bonding strength increases and hence the melting point increases since more heat energy is required to break the metallic bonding. In addition, with more valence electrons for delocalization, the electrical conductivity also increases. [3]

Silicon has a giant covalent structure with strong silicon–silicon single bonds. A large amount of heat energy is required to break all these bonds and hence it has a high melting point. Phosphorus (P_4), sulfur (S_8) and chlorine (Cl_2) are simple molecular covalent substances and hence are held together in the solid state by van der Waals' forces. Argon is a simple atomic substance and is also held together in the solid state by van der Waals' forces. A small amount of energy is required to break these intermolecular forces and hence the melting points are low. The strength or extent of intermolecular forces increases with molecular size. Thus, sulfur has the highest melting point followed by phosphorus, chlorine and then argon. [3]

Q2 **a** Electron dot structures for N_2 and F_2: [2]

The triple bond in the nitrogen molecule is stronger and hence harder to break. [1]

b The C–F bond is more polar than the N–F bond. Nitrogen is more electronegative than carbon. NF_3 is polar; CF_4 is polar. The NF_3 molecule is not symmetrical; CF_4 is symmetrical. [5]

Q3 Ethanol and water are miscible in each other due to the formation of hydrogen bonding. The ethane molecule is non-polar and is not able to form hydrogen bonds with water molecules. [2]

Ethane molecules can only form weaker van der Waals' forces with water molecules. Cholesterol has a polar alcohol group; however, it behaves as a non-polar molecule owing to the relatively large size of the hydrocarbon group. [2]

Q4 **a** Potassium metal has lattice of unipositive ions (cations) in fixed positions with a 'sea' of delocalized or mobile valence electrons that are free to move through the lattice. The metal is held together by attraction between the nuclei of the positive ions and the delocalized electrons. [4]

b Lewis structure for fluorine:

[2]

The bonding within a fluorine molecule is a single covalent bond. This consists of a shared pair of valence electrons. [1]

The bonding between molecules arises from van der Waals' forces. These involve the attraction between dipoles temporarily set up by electron polarization. [1]

c The potassium atom has the configuration 2,8,8,1. [1]

The fluorine atom has the configuration 2,7. [1]

Potassium's outer electron is transferred to the outer shell of the fluorine atom. [1]

Oppositely charged ions are formed (with complete octets of electrons) which attract each other due to electrostatic forces of attraction. [1]

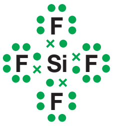
d The ions are not free to move in solid. However, they can move in the molten state under the influence of a voltage. [1]

Q5 a i AC [A^{3+} C^{3-}] [2]

ii BD_2 [1]

b 360 protons, 360 neutrons and 366 electrons in the ion $^{12}C_{60}{}^{6-}$. [2]

c Covalent bonds: Si–F



[3]

d Find the number of electron pairs/charge centres in the valence shell of the central atom. The electron pairs/charge centres in the valence shell of the central atom repel each other to positions of minimum energy/repulsion/maximum stability.
Electron pairs forming a double or triple bond act as a single bond. Non-bonding pairs (lone pairs) repel more than bonding pairs. [3]

e i SBr_2 has two bonding pairs and two non-bonding pairs (lone pairs).
The molecule has an angular/bent/non-linear or V-shape. The bond angle will lie between 90° and 107°.
C_2Br_2 has a triple carbon–carbon bond and hence there are two charge centres or regions of high electron density around each carbon atom. The molecule is linear and the bond angle will be 180°. [6]

ii SBr_2 is polar; C_2Br_2 is non-polar. In SBr_2 the two bond dipoles do not cancel and there is an overall or net dipole moment. There is no net or overall dipole moment for C_2Br_2 since the two bond dipoles cancel. [3]

5 Energetics

Paper 1 IB questions and IB style questions

Q1 C The negative sign by convention indicates that heat is released to the surroundings

Q2 D A positive enthalpy change means that heat flows from the surroundings into the system.

Q3 A Heat energy = mass × specific heat capacity × temperature change
$560\,J = 64.0\,g \times c \times (310\,K - 295\,K)$; $c = 0.583\,J\,g^{-1}\,K^{-1}$

Q4 D The equation is for 2 mol CaO(s).

Q5 D Sublimation of ice is endothermic – work needs to be done to overcome the hydrogen bonds. Bond breaking is endothermic – work needs to be done to break covalent bonds. Vapour deposition is exothermic – energy is released to the surroundings.

Q6 B Bond breaking is always endothermic.

Q7 D Changing a liquid to a gas is endothermic (eliminating A and B). D is more exothermic than C since condensation of a gas to a liquid is an exothermic change.

Q8 D The heat is used to overcome the hydrogen bonds operating between water molecules.

Q9 D Energy = $13.7 \times 209.2 = 2866.04\,J$; enthalpy of reaction per mole of KCl = $(\frac{1}{0.05}) \times 2866.04\,J = 57\,328\,J\,mol^{-1}$;
sign is negative because energy is given out

Q10 A The vertical axis represents potential energy, energy or enthalpy. The reactants are high in energy and hence unstable; the products are lower in energy and hence more stable. The difference is released in the form of heat energy.

Q11 B $Q = mc\Delta T$, $\frac{Q}{mc} = \Delta T$, hence if c is increased then ΔT will decrease. A high heat capacity means more energy is required to raise the temperature.

Q12 $\Delta H = \Sigma$ (bonds broken) – Σ (bonds made)
$= [(615 + 158)] - [(1 \times 348) + (2 \times 484)]$

Q13 B Bond breaking involves work being done – energy is required to pull the bonded atoms away from each other. Bond formation involves the release of energy resulting in a more stable lower energy state.

Q14 B The H–F bond enthalpy is the amount of energy (in kJ) required to break one mole of a hydrogen fluoride covalent bond into gaseous hydrogen and fluorine atoms (under standard thermodynamic conditions).

Q15 A During freezing intermolecular forces (van der Waals' forces) are formed and heat energy is released to the surroundings.

Q16 D Amount of $C_6H_{12}O_6 = \frac{100.0\,g}{180\,g\,mol^{-1}} = 0.55\,mol$
Enthalpy change = $0.55 \times 2824 = +1568\,kJ$

Q17 A $2NO(g) \rightarrow O_2(g) + N_2(g)$; $\Delta H = -180.4\,kJ$ (reversing the first equation)
$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$; $\Delta H = +66.4\,kJ$
Adding the two equations – cancelling the N_2 and simplifying the oxygen:
$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$; $\Delta H = -114\,kJ$ (adding the two enthalpy values together)
Dividing the equation through by two and the enthalpy value by two:
$NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g)$; $\Delta H = -57\,kJ$

Q18 A 800 J to melt 10 g; 3200 J to melt 40 g; 400 J to raise 10 g from 10 to 30 °C and 1600 J to raise 40 g from 10 to 30 °C

Q19 D Enthalpy change $= \Sigma$(bonds broken) – Σ(bonds made)
$= [(432 + 149)] - [(295 \times 2)] = -9\,kJ$

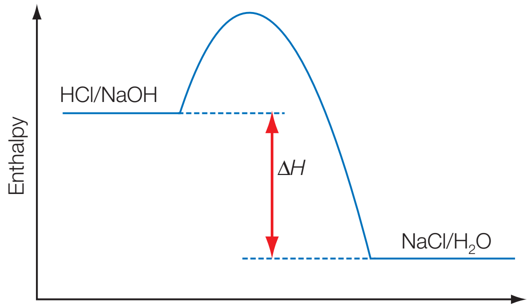
Q20 A Enthalpy change $= \Sigma$(bonds broken) – Σ(bonds made)
$= (N{\equiv}N + 3\ H{-}H) - (6 \times N{-}H)$.

Paper 2 IB questions and IB style questions

Q1 a i Enthalpy of products – enthalpy of reactants [1]
Specified temperature (e.g. 298 K (25 °C)) [1]
Specified pressure (e.g. 1 atm or 1.013×10^5 Pa) [1]

ii Take a known volume of sodium hydroxide solution [1] of known concentration. [1]
Place the solution in an insulated vessel. [1]
Add a solution of hydrochloric acid of known concentration, such that an equal or excess number of moles is added. [1]
Stir and mix the acid and alkali. [1]
Record the initial and final temperatures. [1]
The temperature rise is needed for the calculation. [1]
ΔH = total mass (or volume) × temperature rise × specific heat capacity [1]
Divide by the number of moles (of limiting reactant if excess of other reactant used). [1]

iii Overall diagram: enthalpy label and two different levels [1]
Labelled enthalpy levels – NaCl/H_2O must be lower [1]
Enthalpy change shown [1]
Products more stable than reactants/reaction exothermic [1]



b Hess's law enables enthalpy changes which cannot be found experimentally to be calculated from other experimental results. [1]
Suitable example [1]
Intermediate stage shown [1]
Show which enthalpies can be measured, hence which deduced [1]

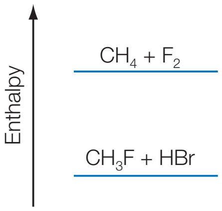
$$\begin{tikzcd}
C(s) + 2H_2(g) \arrow[r, "\text{deduced}"] \arrow[rr, bend right, "\text{measured}"'] & CH_4(s) \arrow[r, "\text{measured}"] & CO_2(g) + 2H_2O(l)
\end{tikzcd}$$

Q2 a i The energy needed to break one bond in a molecule in the gaseous state.
The value is averaged using those from similar compounds. [3]

ii It is an element and is the only species with a F–F bond. [1]

b i Sum of bonds broken = 412 + 158 = 570
Sum of bonds formed = 484 + 562 = 1046
$\Delta H = -476\,\text{kJ mol}^{-1}$ [3]

ii [2]



iii About the same since the same number and type of bonds are being broken and formed. [2]

Q3 a It is exothermic because heat is released to the surroundings and the temperature rises. [1]

b To make any heat loss as small as possible, so that all the heat will be given out very rapidly. [1]

c Heat released = mass × specific heat capacity × temperature increase
Amount of LiOH or HCl used = $0.500\,\text{dm}^3 \times 0.050\,\text{mol dm}^{-3} = 0.025\,\text{mol}$
Heat released = $100\,\text{g} \times 4.18\,\text{J g}^{-1}\,°\text{C}^{-1} \times 3.5\,°\text{C} = 1463\,\text{J} = 1.463\,\text{kJ}$
$\Delta H = (1.463\,\text{kJ}/0.025\,\text{mol}) = -58.5\,\text{kJ mol}^{-1}$ [4]

d Significant heat loss to the surroundings. Insulate the reaction vessel and use a lid. Draw a temperature versus time graph and extrapolate to calculate the rise in temperature. [2]

e The temperature change would be the same, namely, 3.5 °C. This occurs because the amount of lithium hydroxide reacted would be the same. The excess hydrochloric acid would not react; lithium hydroxide is the limiting reagent. [2]

6 Kinetics

Paper 1 IB questions and IB style questions

Q1 A The magnesium carbonate is present in excess – hence there is little change in surface area.

Q2 A Zinc is the limiting reactant and hence determines the volume of the hydrogen. The two samples of zinc have the same mass.

Q3 D Very few colliding molecules have kinetic energies equal to or in excess of the activation energy.

Q4 A Catalysts decrease the activation energy for the forward and backward reactions. They provide a new mechanism or pathway.

Q5 C Colliding particles must have a kinetic energy higher than the activation energy is necessary for a successful collision.

Q6 C Decreasing the temperature reduces the frequency of the collision rate and reduces the number of colliding species with combined kinetic energies greater or equal to the activation energy.

Q7 B The addition of hydrochloric acid of higher concentration means that the reaction will go faster and hence the loss in mass will be faster. In addition, since the acid is the limiting reagent then a greater total volume of gas is lost and hence a greater mass loss results.

Q8 C Activation energy does not vary with temperature.

Q9 A Rates of reaction are the highest at the start of the reaction when the concentrations of the reactants are highest.

Q10 D This combination has acid of the highest concentration and the greatest surface area for the calcium carbonate. Both these factors maximize the collision rate.

Q11 B The bromine atom is an intermediate and behaves as both a reactant and a product.

Q12 B An increase in particle size will reduce the rate of the reaction.

Q13 C Increasing the pressure increases the concentration of reacting gases and hence generally increases the rate of reaction.

Q14 B Tripling the volume of the acid will not alter the initial rate of reaction since there is no change in concentration and hence there will be no effect on the collision rate between magnesium atoms and hydrogen ions from the methanoic acid.

Q15 D An increase in temperature does increase the collision rate, but the rise in the proportion of molecules colliding with kinetic energies equal to or in excess of the activation energy increases is the major underlying reason for the rapid increase in rate.

Q16 A The function of a catalyst is to create a new reaction pathway with a lower activation energy. Catalysts are not chemically consumed during a reaction.

Q17 D Neutralization reactions, between hydrogen and hydroxide ions, are very rapid since oppositely charged ions are involved.

Q18 C Rate is defined as change in concentration (mol dm^{-3}) with time (s).

Q19 C 2 volumes of gas form 2 volumes of gas. Hence there is no change in pressure.

Q20 C The reaction takes place on the surface. The collision rate will be directly proportional to the rate.

Paper 2 IB questions and IB style questions

Q1 a Rate is the increase in product concentration per unit time or the decrease in reactant concentration per unit time. Rate is also change in reactant or product concentration per unit time. [1]

b i The reaction is slow because the reaction has a high activation energy barrier and few molecules have the necessary kinetic energy and correct collision geometry required for a successful reaction. [2]

ii The increase in temperature increases the collision rate and the combined kinetic energy of the colliding molecules. [1]
More pairs of colliding molecules have combined kinetic energies equal to or greater than the activation energy. [1]

iii One mark for any two of the following (max [2]):

Add a catalyst [1]; increase the total pressure which is equivalent to decreasing the volume of the container [1]; increase the concentration of C or D [1]. [2]

Q2 a i $900\,\text{cm}^3$ [1]
ii $500\,\text{cm}^3$ [1]
iii $320\,\text{cm}^3$ [1]

b As the reaction proceeds the surface area of the magnesium (the limiting reagent) decreases and hence the rate decreases. The acid concentration remains relatively constant. [2]

c The reaction has stopped; all of the magnesium has been consumed. [1]

d i Increased [1]
ii Decreased [1]
iii Increased [1]

e i Increased [1]
ii Unchanged [1]

f Hydrogen has a very low molar mass and the mass change observed will be relatively small and hence difficult to measure accurately. [2]

g The rate of formation of hydrogen is half the rate of consumption of hydrochloric acid. [1]
The coefficient in front of the hydrochloric acid is twice the coefficient in front of the hydrogen. [1]

h $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$ [1]
Reactions usually proceed by more than one step; a simultaneous collision between three particles is highly unlikely. [1]

Q3 a i Carbon dioxide, CO_2 [1]
ii $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$ [2]

b $800\,\text{cm}^3$ [1]; 8 minutes [1] [2]

c i The collision frequency between carbonate and hydrogen ions is greatest at the start, but decreases with time. The concentrations of the two reactants is highest at the start of the reaction. However, as the chemicals react the number of reacting ions decreases the concentrations and hence the collision rate decreases. [2]

ii The mass of calcium carbonate may be less in Y and hence greater in X. [1]
The concentration of hydrochloric acid may be less in Y and hence higher in X. [1]

7 Equilibrium

Paper 1 IB questions and IB style questions

Q1 A I and II are correct as the equilibrium involved is a dynamic process; III is not necessarily true, the balance, i.e. the position of equilibrium, depends on the conditions.

Q2 C The concentrations do not change unless the conditions are changed.

Q3 C I is correct as the equilibrium is dynamic; III is correct as the equilibrium can be reached from either direction.

Q4 C Catalysts speed up both forward and reverse reactions equally as they lower the activation energies of these reactions to the same extent.

Q5 A The equilibrium is dynamic with the rates of forward and reverse reactions equal.

Q6 D Forward reaction is endothermic; forward reaction favoured by increased temperature. So the equilibrium shifts to the right. K_c is a measure of how far to the right an equilibrium lies.

Q7 C II is correct as removing oxygen shifts the reaction to the right; III is correct as reduced pressure favours the side of the equation involving fewer moles of gas, in this case the products side.

Q8 B A catalyst speeds up both the forward and the reverse reactions equally; the composition of the equilibrium mixture remains unchanged.

Q9 D This is the correctly constructed expression which fits with the equilibrium law.

Q10 B Production of SO_3 favoured by lower temperatures (as forward reaction is exothermic) and high pressures as the products occupy less volume.

Q11 B The pressure is lowered favouring a shift to the products side of the equilibrium as this involves fewer moles of gas; so the amount of SO_2Cl_2 is decreased.

Q12 D In this case there are the same number of moles of gas on both sides of the equation.

Q13 B This is the correctly constructed expression which fits with the equilibrium law.

Q14 A Forward reaction is endothermic and favoured by an increase in temperature, so more NO_2 is present in the equilibrium mixture (so the mixture is browner).

Q15 C The forward reaction is endothermic and favoured by an increase in temperature. I and II would have no effect on the equilibrium position.

Q16 B This is the correctly constructed expression which fits with the equilibrium law.

Q17 A I would shift the equilibrium to the right and so increase the concentration of ammonia; II would shift the equilibrium to the left.

Q18 D The forward reaction is endothermic and favoured by an increase in temperature (II). Increased pressure would shift the equilibrium to the left.

Q19 C Ethanol production is favoured by lower temperatures and increased pressure.

Q20 B The higher the value of K_c, the greater the value of the numerator (the top figures) in the equilibrium expression; so the greater the concentration(s) of the products. There is no connection between K_c and the rate at which equilibrium is achieved.

Paper 2 IB questions and IB style questions

Q1 a Less product is present at higher temperatures [1]; therefore the forward reaction is exothermic [1]. [2]

b i Forward and reverse reactions still occurring/forward and reverse rates equal. [1]
Concentrations of reactants and products unchanged at constant temperature. [1]

ii Reactants and products in the same phase/state. [1]

iii $K_c = [NH_3]^2/[N_2][H_2]^3$
Use of [ ] and formulas in correct position [1]; powers correct [1]. [2]

c i Rate increased [1]; equilibrium constant decreased [1]. [2]

ii Rate increased [1]; equilibrium constant unchanged [1]. [2]

iii Too expensive/greater cost of energy or pipes/more safety precautions/thicker pipes. [1]

iv Shifted to right [1]; to replace the ammonia removed [1]. [2]

d Iron/Fe [1]
Equilibrium constant unchanged [1]; only temperature affects K_c/catalyst speeds up forward and reverse reactions equally/activation energy reduced by same amount for both forward and reverse reactions [1]. [2]

e One mark for any three of the following (max [3]):
Successful collisions need minimum/activation energy/correct geometry [1]
Increasing temperature causes increase in kinetic energy of particles [1]
So increased proportion of successful collisions [1]
So increased frequency of collisions/more collisions per unit time [1]
(Simply writing ‘more collisions’ is not acceptable.)

Q2 a $K_c = [HI]^2/[H_2][I_2]$ [1]

b Concentration of product/HI greater (than $[H_2]$ and $[I_2]$) [1]

c It will have no effect. [1]

d As the reaction is endothermic, increasing the temperature will shift the equilibrium position to the right (more product formed). [1]

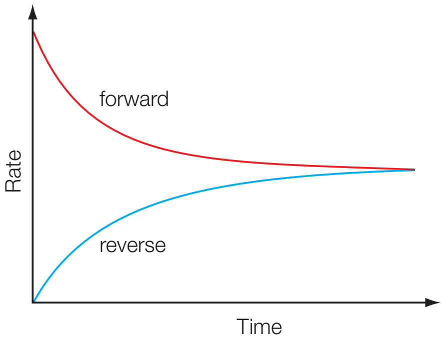
Q3 a $K_c = [NO_2]^2/[N_2O_4]$
Horizontal line concentration of reactant and product remains constant/equilibrium reached.
Magnitude of K_c greater than 1 as product concentration greater than reactant concentration. [4]

b Increased temperature shifts equilibrium position to right. [1]
Forward reaction is endothermic/absorbs heat. [1]

c Increased pressure shifts equilibrium to the left. [1]
Fewer (gas) moles/molecules on the left. [1]

d Both/forward and reverse rates increased/increase in forward reverse rates are equal; activation energy reduced; position of equilibrium unchanged.
Concentration/amount of reactants and products remain constant.
Value of K_c unchanged.
K_c only affected by changes in temperature. [6]

Q4 a



Two curves – one labelled ‘forward’ starting up high up y-axis and one labelled ‘reverse’ starting from zero. The curves merge and become horizontal.

i Forward reaction – highest concentration, thus rate high to begin with. As reaction proceeds, concentrations decrease, so does rate.

ii Reverse reaction – zero rate initially/at $t = 0$ (since no products present). Rate increases as concentration of products increases.
Equilibrium established when rate of forward reaction = rate of reverse reaction. [7]

b Reaction is endothermic [1]
K_c increases with (increasing) temperature [1]
Forward reaction favoured/heat used up [1]

Q5 a 200 °C [1] and 600 atm [1]. [2]
(Allow the ‘highest pressure and the lowest temperature’.)

b i Yield increases/equilibrium moves to the right/more ammonia [1]
4 (gas) molecules → 2/decrease in volume/fewer molecules on right-hand side [1]

ii Yield decreases/equilibrium moves to the left/less ammonia [1]
Exothermic reaction [1]

c High pressure expensive/greater cost of operating at high pressure/reinforced pipes, etc. needed
Lower temperature – greater yield, but lowers rate [2]
(Do not award a mark for simply saying 'compromise'.)

d $K_c = [NH_3]^2/[N_2][H_2]^3$ [1]

e **i** Artificial fertilizers/increasing crop yields; production of high explosives for mining [1 max] [1]
ii Fe/iron (allow magnetite/iron oxide) [1]
The claim is not valid since catalysts do not alter the yield/position of equilibrium/only increase the rate of reaction. [1}

8 Acids and bases

Paper 1 IB questions and IB style questions

Q1 **D** A strong acid is almost completely dissociated in water: $HA(aq) \rightarrow H^+(aq) + A^-(aq)$.

Q2 **C** Metals that react with acids form hydrogen and a salt.

Q3 **A** Strong acid < weak acid < weak base < strong base

Q4 **C** Pure distilled water is essentially molecular but contains low but equal concentrations of hydrogen and hydroxide ions.

Q5 **D** Metal hydroxides react with acids to form a salt and water.

Q6 **A** The conjugate acid is formed by protonation of the amine group.

Q7 **A** The most concentrated acid has the greatest amount of hydrogen ions.

Q8 **D** A conjugate acid–base pair differ by a proton (H^+).

Q9 **B** Acid behaviour is observed in water since the water molecule is a strong base (proton acceptor) and forms oxonium ions, H_3O^+.

Q10 **D** Sodium hydroxide is a strong alkali.

Q11 **B** The low concentration of ions accounts for its low conductivity.

Q12 **A** Alkaline solutions have pH values of greater than 7.

Q13 **D** Weaker acids have higher pH values than strong acids of the same concentration.

Q14 **B** A decrease in pH correlates to an increase in hydrogen ion concentration.

Q15 **D** pH and $[H^+]$ are inversely related. An increase in the pH by a value of one corresponds to a ×10 decrease in the H^+ concentration.

Q16 **B** $HSO_4^- \rightarrow SO_4^{2-} + H^+$; $HSO_4^- + H^+ \rightarrow H_2SO_4$

Q17 **D** A change in pH of one unit corresponds to a change of ×10 in the concentration of hydrogen ions.

Q18 **D** The sulfuric acid molecule is deprotonated.

Q19 **C** The ammonium ion can only act as an acid: $NH_4^+ \rightarrow NH_3 + H^+$

Q20 **A** $H_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + 2H_2O$
0.1 mol of sulfuric acid reacts with 0.1 mol of barium hydroxide.

Q21 **A** Hydrogen chloride reacts with water to form a solution of hydrochloric acid.

Q22 **C** Carbonic acid is a weak acid.

Q23 **A** A monoprotic acid releases one hydrogen ion per molecule in water.

Q24 **C** A Lewis base is any molecule or ion that can form a new coordinate covalent bond, by donating a pair of electrons.

Q25 **D** The ammonia molecule has an available lone pair of electrons and is able to accept a hydrogen ion and form the ammonium ion (via dative bond formation).

Paper 2 IB questions and IB style questions

Q1 **a** **i** Acid X [1]
ii Y > Z by a factor of 10 [1]
iii Y, Z and X [1]; conductivity increases with ion concentration. As the pH value decreases the concentration of hydrogen ions increases [1]. [2]

Q2 **a** A strong acid is fully ionized or dissociated in solution. [1]
$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$ [1]
A weak acid is partly ionized or dissociated in solution [1]
$H_2CO_3(aq) \rightleftharpoons 2H^+(aq) + CO_3^{2-}(aq)$
or
$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$ [1]

b Universal indicator: hydrochloric acid – red; carbonic acid – yellow/orange [1]
Electrical conductivity: hydrochloric acid – high; carbonic acid – low
or
Reaction with metal/metal carbonate: hydrochloric acid – vigorous; carbonic acid – slow. [1]

c 10000 : 1 or $10^4 : 1$ or 0.1 : 0.00001 or $10^{-1} : 10^{-5}$ [2]

d **i** Base: HCO_3^-; conjugate acid: H_2CO_3 [1]
Acid: HCl; conjugate base: Cl^- [1]
ii Brønsted–Lowry [1]

9 Oxidation and reduction

Paper 1 IB questions and IB style questions

Q1 **A** The products of the reduction of the manganate(VII) ion depend on pH. However, under acidic conditions, manganese(II) ions are produced:
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

Q2 **B** The formation of chlorine from hydrochloric acid is an oxidation process: the oxidation number changes from 0 to –1.

Q3 **D** $6H^+(aq) + 6NO_3^-(aq) + I^-(aq) \rightarrow IO_3^-(aq) + 6NO_2(g) + 6H_2O(l)$

Q4 **A** Precipitation reactions do not involve a change in oxidation number.

Q5 **D** The anode is positive and attracts negative ions. Oxidation occurs at the anode of an electrolytic cell.

Q6 **B** The nickel complex is neutral and consists of a nickel atom and four carbon monoxide molecules.

Q7 **A** X is less reactive than Y; Z is more reactive than Y.

Q8 **D** Oxidizing agents undergo reduction during a redox reaction: the oxidation number of oxygen changes from 0 to –2; the oxidation number of chlorine changes from 0 to –1.

Q9 **A** The iodine undergoes a change in oxidation number from 0 to –1.

Q10 B The chromium remains in oxidation state +6.

Q11 B Chloride ions can only act as a reducing agent: $2Cl^- \rightarrow Cl_2 + 2e^-$. All the other chemicals are in an intermediate oxidation state.

Q12 D Metals that are weak reducing agents (electron donors) form ions that are strong oxidizing agents (electron acceptors).

Q13 C Iodine is below bromine in group 7 and hence not sufficiently powerful to oxidize bromide ions to bromine molecules.

Q14 C Electrons consumed at the cathode by reduction are supplied by the anode by the action of oxidation.

Q15 B $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ (multiply through by 3)
$3H_2O + Y^{2+} \rightarrow YO_3^- + 6H^+ + 3e^-$ (multiply through by 5)
$3MnO_4^- + 24H^+ + 15e^- + 15H_2O + 5Y^{2+} \rightarrow 3Mn^{2+} + 12H_2O + 5YO_3^- + 30H^+ + 15e^-$
$3MnO_4^- + 3H_2O + 5Y^{2+} \rightarrow 3Mn^{2+} + 6H^+$
Amount of $MnO_4^- = \frac{3}{5} \times$ amount $Y^{2+} = 3.216 \times 10^{-3}$ mol

Q16 B $2Cr + (-14) = -2$; $2Cr = +12$; $Cr = +6$

Q17 D Ions move to electrodes opposite in sign to their charges.

Q18 D The manganese changes oxidation number from +6 to +7. The chromium remains in oxidation number +6. The iron changes from oxidation number +2 to +3.

Q19 A Precipitation reactions do not involve a change in oxidation number.

Q20 C The oxidation number is +5.

Paper 2 IB questions and IB style questions

Q1 a i Magnesium (Mg), iron (Fe), copper (Cu), gold (Au) [1]

ii The loss or removal of electrons from a chemical species [1]
Iron atoms lose electrons during their reaction with copper(II) ions:
$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ [1]

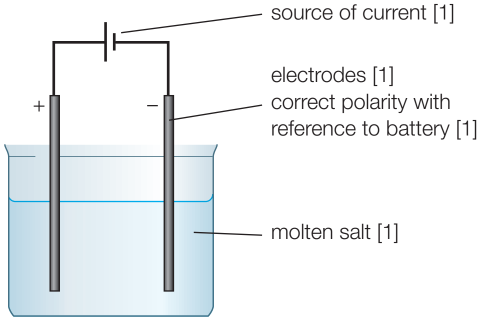
iii Reduction involves a decrease in oxidation number. [1]
Copper(II) ions gain electrons during their reaction with iron atoms: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ [1]

iv Magnesium (Mg) is the strongest reducing agent [1]; It reduces Fe^{2+}; Fe reduces Cu^{2+}; Cu reduces Au^+ [1]. [2]

v Gold ions, Ag^+ [1]; Every metal present can reduce gold ions to gold atoms [1]. [2]

vi Does not react [1]; It is too low in the reactivity series and hence is less reactive than magnesium [1]. [2]

b [4]



c At the cathode (negative electrode):
electrons are given to the cations (positive ions) [1]
electrons are removed from the anions (negative ions). [1]
At the anode (positive electrode):
copper atoms formed (at the cathode) [1]
chlorine molecules formed (at the anode). [1]

Q2 a At the negative electrode (cathode):
$K^+ + e^- \rightarrow K$ [1]
At the positive electrode (anode):
$2F^- \rightarrow F_2 + 2e^-$ [1]
Electrons flow through the external circuit or wires and the ions move through the electrolyte to the electrodes where they gain or lose electrons. [2]

b 0.2 mol since the K to F_2 molar ratio is 2 : 1. [2]

c i Sodium atoms lose electrons and undergo reduction. Aluminium ions (in aluminium chloride) gain electrons and undergo oxidation. [2]

ii Reactants: sodium metal: 0; aluminium in aluminium chloride: 3; aluminium metal: 0; sodium in sodium chloride: +1. [2]

Q3 a The voltage of the cell decreases as the metal is closer to the reactivity series. [1]

b The metal is the negative electrode because it is higher than silver in the reactivity series and it will oxidize (lose electrons) in preference to silver. [2]

c $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$; $Ag^+(aq) + e^- \rightarrow Ag(s)$ [2]

10 Organic chemistry

Paper 1 IB questions and IB style questions

Q1 C See page 266 in Chapter 10.

Q2 B Aldehydes contain the –CHO functional group.

Q3 D Members of an homologous series differ by adding $-CH_2-$ units to the chain.

Q4 C Molecule C has the formula C_5H_{12}; all other answers incorrect.

Q5 B The product is an ester, ethyl ethanoate.

Q6 A All structures are valid and have the molecular formula C_4H_{10}.

Q7 A C_6H_{14} and $C_{13}H_{28}$ are both alkanes. C_2H_4 and $C_{10}H_{20}$ are both alkenes, but only C_2H_4 is a gas at room temperature.

Q8 A Light is a form of energy and is absorbed by chlorine molecules to dissociate the covalent bond between the two chlorine atoms. Homolytic fission occurs: $Cl_2(g) \rightarrow 2Cl\bullet(g)$

Q9 D Alkenes rapidly react with bromine water via an addition reaction.

Q10 D Carbon dioxide and water are formed by the complete combustion of a hydrocarbon.

Q11 A All the other compounds are possible products of repeated free radical substitution. B is the product of a possible termination reaction.

Q12 B It is a tertiary alcohol; $C(CH_3)_3-OH$ or $CH_3C(CH_3)_2-OH$.

Q13 A There are four carbon atoms in the longest unbranched chain. There is a methyl group and an –OH group on the second carbon atom.

Q14 B The bond angle in ethane is smaller, 109.5° compared to 120°; bond length in ethane is longer as the bond is just a single C–C bond, whereas there is a C=C bond in ethene.

Q15 D It contains an ester linkage: –CO–O–.

Q16 A Propan-2-ol is a secondary alcohol and secondary alcohols are oxidized to ketones.

Q17 D Repeated condensation to form a polyamide could occur.

Q18 B The compound is an ester formed from butanoic acid (see the first part of the structure) and methanol (see the second part of the structure); therefore it is methyl butanoate.

Q19 C The members of an homologous series are not isomers of each other; other statements are correct.

Q20 A These molecules will only have van der Waals' forces acting between them; other examples have the capability of hydrogen bonding between the molecules.

Paper 2 IB questions and IB style questions

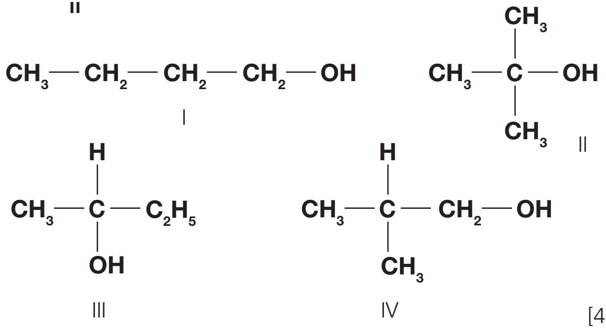
Q1 a Molar ratio = 1 : 2.2 [1]
Therefore empirical formula is C_5H_{11} [1]
b Molar mass of empirical formula = 71 [1]
Molecular mass = 142; therefore molecular formula = 2 × empirical formula
Molecular formula of hydrocarbon = $C_{10}H_{22}$ [1]
c i $2C_{10}H_{22} + 31O_2 \rightarrow 20CO_2 + 22H_2O$ [2]
ii $2C_{10}H_{22} + 21O_2 - 20CO + 22H_2O$ [2]
Could also have an equation where C is produced.
d i C_4H_{10} [1]
ii C_5H_{10} [1]

Q2 a One mark for any three of the following (max [3]):
Same general formula/C_nH_{2n} [1]
Formulas of successive members differ by CH_2 [1]
Similar chemical properties/same functional group; gradation/gradual change in physical properties. [1]
b But-2-ene [1]
Strongest intermolecular/van der Waals' forces [1]
Largest (molecular) mass/size/surface area/area of contact [1]
c $CH_2CHCH_2CH_2CH_3$ / $CH_3CHCHCH_2CH_3$ /any correct branched structure [1]
Pent-1-ene/pent-2-ene [1]
d $C_4H_8 + HBr \rightarrow CH_3CH_2CHBrCH_3$ [2]
Addition [1]
e Oxidation/redox [1]
(Potassium or sodium) dichromate(VI)/$Cr_2O_7^{2-}$ [1]
(Sulfuric) acid [1]
Distilling off propanal as it is formed [1]
Heating under reflux with excess oxidizing agent (to obtain propanoic acid) [1]
f (Propan-1-ol) hydrogen bonding; (propanal) dipole–dipole attractions; (propanoic acid) hydrogen bonding
Propanoic acid > propan-1-ol > propanal [4]

Q3 a i % oxygen = 36.4
Empirical formula is C_2H_4O [3]
ii Molar mass of empirical formula = 44
Molecular formula = 2 × empirical formula = $C_4H_8O_2$ [2]
b Empirical formula is CH_2O
Molar mass of empirical formula = 30
Molecular formula = 2 × empirical formula = $C_2H_4O_2$ [2]

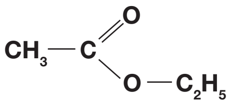
Q4 a Full or condensed structures: $CH_3CH_2CH_2CH_3$, butane/ $CH_3CH(CH_3)CH_3$, 2-methylpropane [4]
b A; CH_3CH_2COOH/$CH_3CH_2CO_2H$ propanoic acid
B or C; CH_3COOCH_3 methyl ethanoate
C or B; $HCOOCH_2CH_3$ ethyl methanoate [6]
c i A forms hydrogen bonds with water [1]
A ionizes/dissociates to give H^+ ions [1]
ii $C_2H_5COOH + NaOH \rightarrow C_2H_5COONa + H_2O$ [1]
iii No C=C bond [1]
d i Esters [1]
Flavourings/plasticizers/solvents/perfumes [1]
ii Carboxylic acid/alkanoic acid [1]; alcohol/alkanol [1]; water [1]. [3]
e Any feasible formula containing C=C (e.g. $CH_3CHC(OH)_2$ or $HOCH{=}C(OH)CH_3$) [1]
Addition [1]
Mixture is decolorized/colour change is from yellow/orange to colourless [1]

Q5 a i Same general formula/differ by $-CH_2-$/similar chemical properties/gradual change in physical properties. [3]
ii



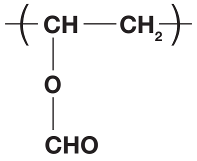
[4]

b i Esterification/condensation
$CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$
Ethyl ethanoate



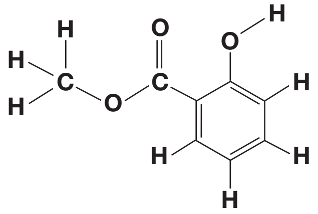
[4]

ii Catalyst; lowers activation energy (by providing an alternate pathway) [2]
iii Flavouring agents/in plasticizers/in solvents/in perfumes [1]
c i II reacts with bromine/contains C=C/II is an alkene/has unsaturated R group.
I contains only saturated R groups. [2]
ii Addition polymerization



[2]

Q6 **a** **i** Empirical formula = $C_8H_8O_3$



[2]

ii The bond at 0.1373 nm is a double bond and the bond at 0.1424 nm is a single bond. [1]
In $CO_2(g)$ both bonds are double bonds and would have a value around 0.137 nm. [1]

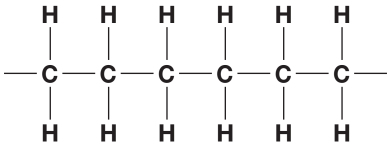
iii Ester/arene or benzene ring/alcohol. [2]

b **i** Boiling point increases as the number of carbons increases. [1]
Greater M_r *and* hence greater van der Waals'/London/dispersion forces present. [1]

ii $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$ needs UV light
Initiation step:
$Cl_2 \rightarrow 2Cl\bullet$ needs uv light
Propagation step:
$CH_4 + Cl\bullet \rightarrow CH_3 + HCl$
$CH_3 + Cl_2 \rightarrow CH_3Cl + Cl\bullet$
Termination step:
$Cl\bullet + Cl\bullet \rightarrow Cl_2$ or $Cl\bullet + \bullet CH_3 \rightarrow CH_3Cl$
or $\bullet CH_3 + \bullet CH_3 \rightarrow CH_3CH_3$ [5]

c **i** A = $CH_3(CH_2)_7CHO$
B = $CH_3(CH_2)_7COOH/CH_3(CH_2)_7CO_2H$
C = $(CH_3)_3COH$
D = $(CH_3)_2CO$
E = $BrCH_2CH_2Br$ [5]

ii Addition



$$\begin{array}{ccccccccccccc} & H & & H & & H & & H & & H & & H & \\ & | & & | & & | & & | & & | & & | & \\ - & C & - & C & - & C & - & C & - & C & - & C & - \\ & | & & | & & | & & | & & | & & | & \\ & H & & H & & H & & H & & H & & H & \end{array}$$

[2]

Q7 **a** **i** Propan-2-ol [1]
ii Propan-1-ol [1]
iii 2-methylbutan-2-ol [1]
iv Ethane-1,2-diol [1]

b **i** Secondary [1]
ii Primary [1]
iii Tertiary [1]
iv Primary [1]

c **i** Propanone, CH_3COCH_3 [2]
ii Propanoic acid, CH_3CH_2COOH [2]
iii No oxidation [1]
iv Ethane-1,2-dioic acid, $HOOCCH_2CH_2COOH$ [2]

11 Measurement and data processing

Paper 1 IB questions and IB style questions

Q1 **C** Zeros to the right of the decimal point are significant.

Q2 **B** Zeros to the right of the decimal point and between non-zero numbers are significant.

Q3 **A** Zeros to the right of the decimal point are significant.

Q4 **B** 100.5094 = 100.5 (When adding the final result should be reported to the same number of decimal places as the value with the least number of decimal places, i.e. 48.2 m.)

Q5 **A** 435.54 = 436 (When adding the final result should be reported to the same number of decimal places as the number with the least number of decimal places, i.e. 451.)

Q6 **D** Percentage uncertainty = ±3 (0.1/40) = ± 0.0075 = ±0.75

Q7 **C** $1\ \mu g = 10^{-6} g = 10^{-6} \times 10^{-3} kg = 10^{-9} kg$; percentage random uncertainty = $(10^{-9} kg/1\ kg) \times 100\% = 10^{-7}\ \%$

Q8 **B** A small systematic error occurs when all the readings are consistently higher or lower than true values. Imprecise readings implies there is a spread of readings about the true value. The spread for balance B is more than balance A.

Q9 **C** Value of error = $327.66\ kJ\ mol^{-1} \times 0.03 = 9.83\ kJ\ mol^{-1}$ = $10\ kJ\ mol^{-1}$ (1 significant figure). Enthalpy of combustion = $330\ kJ\ mol^{-1}$ (rounded to the same tens place as the error).

Q10 **A** A zero error is a systematic error.

Q11 **B** Precision refers to how close the titre data points are from one another. Accuracy is a measure of how close the mean value is from the true value. Answer's B's titre data points are close to one another, but their mean value is very from the far true value x_0. Hence the results in answer B are precise but not accurate.

Q12 **A** Percentage error decreases as the size of the measurement increases.

Q13 **B** Random uncertainty is half the smallest division.

Q14 **A** There is an implied random uncertainty of ±0.05 cm.

Q15 **D** $\frac{0.005}{6.25} \times 100 = 0.0769\% = 0.08\%$

Q16 **C** Systematic error: I, III, VI; Random uncertainties: II, IV, V

Q17 **B** Low accuracy – they are far from the centre; but they are close together (high precision).

Q18 **C** The measuring cylinder has the most imprecise scale.

Q19 **A** There are three significant figures, equal to the number of significant digits in each number.

Q20 **B** Percentage error $= \frac{\text{literature} - \text{experimental value}}{\text{literature value}}$

$$= \frac{(278 - 258)}{278} \times 100 = 7\%$$

Paper 2 IB questions and IB style questions

Q1 **a** **i** The actual volume of sodium hydroxide will be lowered, and the calculated concentration of hydrochloric acid will be higher. [2]

ii If there was water in the burette, the sodium hydroxide would be diluted, reducing the amount of sodium hydroxide and the calculated concentration of hydrochloric acid. [2]

iii This will make the concentration of the hydrochloric acid appear greater than it really is. [1]

iv The end-point will not be accurate, causing random errors. [1]

b **i** Systematic error [1]

ii Measure the pH of a buffer solution of known pH value. [1]

Q2 **a** Mass of copper, time and current [3]

b Directly proportional [1]

c Rate = $\frac{\text{mass of copper/g}}{\text{time/min}} = \frac{1.24\,\text{g}}{15\,\text{min}} = 0.08\,\text{g min}^{-1}$ [2]

d Interpolation, area under the graph and extrapolation [2]

Q3 **a** Percentage uncertainty in length = $(\frac{1}{298}) \times 100 = 0.336\%$ [1]

b Percentage uncertainty in width = $(\frac{1}{210}) \times 100 = 0.4\%$ [1]

c Area = $298\,\text{mm} \times 210\,\text{mm} = 62\,580\,\text{mm}^2$ [1]

$$\text{Uncertainty in area} = (\frac{1}{298} + \frac{1}{210}) \times 62\,580$$
$$= 508\,\text{mm}^2$$

Area = $62\,580\,\text{mm}^2 \pm 508\,\text{mm}^2$ [2]

Q4 $\frac{\Delta d}{2.50} = \frac{0.1}{25.0} + \frac{0.01}{5.00} + \frac{0.01}{2.00} + \frac{0.01}{1.00}$; $\Delta d = 0.05\,\text{g cm}^{-3}$

Hence the density is $2.50\,\text{g cm}^{-3} \pm 0.05\,\text{g cm}^{-3}$ [2]

Q5 $(120.2\,\text{g} - 119.0\,\text{g}) \times (\frac{1\,\text{cm}^3}{2.05\,\text{g}}) = 0.59\,\text{cm}^3$ [2]

12 Atomic structure

Paper 1 IB questions and IB style questions

Q1 **D** Titanium has an atomic number of 22. Its electron configuration is $1s^22s^22p^63s^23p^64s^23d^2$.

Q2 **A** The removal of an electron from a sodium ion would involve the removal of a core electron from the second shell close to the nucleus. This electron would be strongly held and hence would have a large ionization energy. This electron would have the highest effective nuclear charge.

Q3 **C** The Fe^{2+} ion has the configuration $[Ar]3d^6$. Hence there is one spin pair and four unpaired electrons.

Q4 **B** 1s sub-shell: 1 occupied orbital; 2s sub-shell: 1 occupied orbital; 2p sub-shell: 3 occupied orbitals; 3s sub-shell: 1 occupied orbital; 3p sub-shell: 3 occupied orbitals: two filled and one half-filled.

Q5 **C** The two elements are both in the same group (group 1). Elements in groups, especially those at the extremes of the periodic table, have very similar chemical properties.

Q6 **A** Ionization energies decrease down a group and generally increase across a period.

Q7 **C** A tin atom has the following electron configuration: $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^2$.

Q8 **A** Ionization energies decrease down a group and generally increase across a period. The valence electron in sodium is experiencing the lowest effective nuclear charge.

Q9 **D** The electron configuration of the cobalt(III) ion is: $1s^22s^22p^63s^23p^63d^6$. The 4s electrons are lost from a transition metal atom before the 3d electrons.

Q10 **C** The electron configuration of the iron(III) ion is $1s^22s^22p^63s^23p^63d^5$. The 4s electrons are lost from a transition metal atom before the 3d electrons.

Q11 **D** The electron arrangements of the species are all 2,8, that is, they are iso electronic. However, the atomic (proton) numbers are 8, 9 and 11. The effective nuclear charge increases left to right and hence radii decrease.

Q12 **B** The electron configuration for a manganese atom is $4s^23d^5$.

Q13 **D** Na 2,8,1 (11 protons), Na^+ 2,8 (11 protons), Mg^{2+} 2,8 (12 protons) and Al^{3+} 2, 8 (13 protons). The radii decrease due to the increasing effective nuclear charge.

Q14 **B** Na^+ 2,8 (11 protons), N^{3-} 2,8(7 protons), F^- 2,8 (9 protons), Ne 2,8 (10 protons). The electrons in the nitride ion are experiencing the lowest effective nuclear charge.

Q15 **D** A cobalt atom has the following electron configuration: $1s^22s^22p^63s^23p^63d^74s^2$. The 3d sub-shell has two spin pairs and three unpaired electrons.

Q16 **A** Ionization energies generally increase across periods. This results in a decrease in metallic behaviour, that is, a decrease in tendency to form cations (positive ions).

Q17 **C** A carbon atom has the following electron configuration $1s^22s^22p^2$. The p sub-shell has two unpaired electrons in accordance with Hund's rule.

Q18 **B** All d sub-energy levels have 10 orbitals and hence can hold to a maximum of 10 electrons.

Q19 **A** Ionization energies decrease down a group and generally increase across a period. The electrons in argon are experiencing the highest effective nuclear charge.

Q20 **A** X is in group 1 and Y is in group 2 or 4.

Q21 **D** Al^{3+} 2,8 (13 protons), Mg^{2+} 2,8 (12 protons), F^- 2,8 (9 protons), S^{2-} 2,8,8 (16 protons). The electrons in the sulfide ion are experiencing the lowest effective nuclear charge.

Q22 **D** Ni $3d^84s^2$; Mn $3d^54s^2$; Cu^{2+} $3d^9$; Ni^{2+} $3d^8$

Q23 **D** Elements in the first transition series (with the exception of copper and chromium) have a $3d^x4s^2$ configuration.

Q24 **D** The atomic number is the number of protons in the nucleus. The proton number determines the number of electrons and hence the electron configuration.

Q25 **B** Oxygen 2,6 (8 protons), fluorine 2,7 (9 protons), neon 2,8 (10 protons). Effective nuclear charge is increasing from left to right.

Q26 **C** From lithium to neon the atomic number (proton number) increases by one. This is accompanied by the addition of an extra electron.

Q27 **C** I has two unpaired 2p electrons; II has three unpaired 2p electrons and III has two unpaired 2p electrons.

Q28 **C** The chromium(III) ion has the electron configuration $3d^3$. The chromium atom has the electron configuration $3d^54s^1$.

Q29 **C** X will have the electron configuration is $[Ar]3d^54s^2$. Transition metal ions ionise via loss of 4s and then 3d electrons. The total number of electrons in the atom is 18 + 5 + 2 = 25.

Q30 **D** A germanium atom has the following configuration: $1s^22s^22p^63s^23p^64s^2\ 3d^{10}4p^2$. Three 2p orbitals, three 3p orbitals and two unfilled 4p orbitals.

Paper 2 IB questions and IB style questions

Q1 **a** The first ionization energy is the minimum energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of unipositive gaseous ions (under standard thermodynamic conditions). [2]

b In a group 2 element (C) the electron (for the first ionization energy) is removed from a spin pair in a s sub-shell. For D, a group 3 element, the electron (for the first ionization energy) is removed from a 3p sub-shell further away from the nucleus. The 3p sub-shell also experiences more shielding. Hence there is a decrease in first ionization energy when moving from C to D. The increase in nuclear charge accounts for the increase from E (np^2 to F np^3). In addition C cannot be in group 5 (the other group after which there is a slight decrease) because there is a steady rise for the next three elements (D to E to F) indicating the filling of a p sub-shell). [4]

c As you move down group 2 the first ionization energies decrease. As you descend the group the nuclear charge increases due to the presence of additional protons, but the shielding effect progressively increases due to the presence of an extra electron shell as you move from one period to the next. The effect of the extra protons is compensated for by the effect of the extra electrons. The only factor left is the additional distance between the outer electron and the nucleus. That lowers the first ionization energy. [3]

d C is in period 3 since in period 2 the group 2 element would have the higher ionization energy while in period 4 the group 2 element would have a lower ionization energy. This can be accounted for by the progressively increasing average distance the valence electron is from the nucleus. C cannot be in the first period (He) as p-orbitals are being filled, it cannot be in the fourth period (Ca) as ($x + 8$) would not bring you back to group 2 again. [3]

e There is a slight decrease from F to G due to the presence of a spin pair in the *n*p sub-shell for G. The resulting electron–electron repulsion is greater than the effect of the increase in nuclear charge and reduces the ionization energy of G^+. [2]

Q2 i 4p [1]

ii $Ge^{+}(g) \rightarrow Ge^{2+}(g) + e^{-}$ [1]

iii 5th electron removed from 3rd energy level and 4th electron from 4th energy level. The attraction by the protons in the nucleus is greater since the electrons are closer to the nucleus. [2]

Q3 a Across a period the atomic number increases progressively by one unit, thus resulting in an additional proton being added to the nucleus and thus an increase in the attraction for all the electrons, including the outermost electron. [1]

Since the number of electron shells for all elements in period 3 is constant, namely, three, there is very little increase in shielding. [1]

Hence, the increase in nuclear charge across the period will bring all the electrons progressively closer to the nucleus, causing the removal of the second electron from the atom to require progressively more energy. [1]

b The electron configurations of the S^+ and Cl^+ ions are $1s^22s^22p^63s^23p^3$ and $1s^22s^22p^63s^23p^4$. [2]

c The Cl^+ ion has a spin pair of electrons in the 3p sub-shell. This electron-electron repulsion counteracts, to some extent, the increase in nuclear charge. Hence, the increase in ionization energy is relatively small. [1]

13 Periodicity

Paper 1 IB questions and IB style questions

Q1 D Ligands may be anions or molecules. They must contain at least one non-bonding pair of electrons for dative bond formation.

Q2 A The outer configuration is $3d^34s^2$, indicating that it could lose or share five electrons.

Q3 B d-block metals show only a slight decrease in atomic radius across the first row.

Q4 A Group 1 and 2 metals are more reactive than transition metals.

Q5 B Magnesium reacts to form magnesium hydroxide, a weak alkali; the oxides of sulfur both form acidic solutions.

Q6 B $[Fe(CN)_6]^{4-}$ is composed of Fe^{2+} and $6CN^-$ and $[Fe(CN)_6]^{3-}$ is composed of Fe^{3+} and $6CN^-$.

Q7 C P_4O_{10} forms H_3PO_4; SO_2 forms H_2SO_3; NO is a neutral oxide.

Q8 D The nitrogen atom does not have a lone pair.

Q9 B MgO is the correct formula for magnesium oxide.

Q10 C d-block metals are relatively unreactive and have relatively low standard electrode potentials.

Q11 D Na_2O and Al_2O_3 are both ionic; SiO_2 is giant covalent.

Q12 B Transition metal compounds are coloured.

Q13 D Al_2O_3 and SiO_2 are both insoluble; Na_2O forms NaOH. Sulfur dioxide forms sulfurous acid: $SO_2 + H_2O \rightarrow H_2SO_3$.

Q14 B The +2 oxidation number corresponds to the loss of the $4s^2$ electrons.

Q15 A d–d transitions are generally responsible for the colours of d-block compounds.

Q16 B Aluminium chloride has a low melting point and sublimes, which indicates covalent character.

Q17 C Magnesium oxide is a basic oxide and hence would not be expected to react with sodium hydroxide.

Q18 D The chromium atom has the configuration $[Ar]3d^54s^1$.

Q19 A The potassium dichromate(VI) acts as an oxidizing agent and is converted to chromium(III) ions.

Q20 D Aluminium and lead form Al^{3+} and Pb^{4+} ions; group 4 metals can form complex ions. Most non-transition metal compounds are colourless.

Q21 A Titanium can lose up to a maximum of four electrons (from the 4s and 3d orbitals) in the formation of compounds, i.e., titanium can exhibit oxidation numbers of +1, +2, +3 and +4. In A, titanium has an oxidation number of +6.

Q22 B Magnesium oxide is basic, aluminium oxide is amphoteric and phosphorus(V) oxide is acidic.

Q23 D There is no change in oxidation state: both ions contain chromium in the +6 oxidation state.

Q24 A Fe^{3+} has the configuration $3d^5$ and hence has five unpaired electrons.

Q25 **C** Catalysts often operate via the formation of intermediates where the catalyst undergoes a temporary change in oxidation state.

Paper 2 IB questions and IB style questions

Q1 Chlorine is the most electronegative element in period 3 and forms ionic compounds with the reactive metals (with the exception of aluminium) and non-metals.
Sodium and magnesium form ionic chlorides by the transfer of electrons; [1]
aluminium and the non-metals share electrons to form covalent bonds. [1]
Sodium and magnesium chloride dissolve in water to release hydrated ions. [1]
Neutral solutions are formed. [1]
$NaCl(s) + (aq) \rightarrow Na^+(aq) + Cl^-(aq)$ [1]
$MgCl_2(s) + (aq) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$ [1]
(Magnesium chloride undergoes slight hydrolysis with water and a slightly acidic solution is formed.)
Anhydrous aluminium chloride and the non-metallic chlorides are covalent; [1]
and undergo hydrolysis with water to form hydrochloric acid: [1]
$2AlCl_3 + 3H_2O \rightarrow Al_2O_3 + 6HCl$
(or $Al(OH)_3$ + HCl or $[Al(H_2O)_5(OH)]^{2+}$) [1]
$SiCl_4 + 2H_2O \rightarrow SiO_2 + 4HCl$ (or $Si(OH)_4$ + HCl) [1]
$PCl_3 + 3H_2O \rightarrow H_3PO_3 + 3HCl$ [1]
$PCl_5 + 4\ H_2O \rightarrow H_3PO_4 + 5HCl$ [1]

Q2 Across period 3 of the periodic table, the oxides of the elements change from basic to amphoteric to acidic in nature; [1]
due to the increasingly electronegativity of the elements, which gradually changes the bonding from ionic to covalent. [1]
Sodium and magnesium oxide react to form hydroxide ions:
$Na_2O + H_2O \rightarrow 2NaOH$
$MgO + H_2O \rightarrow Mg(OH)_2$ [1]
Aluminium oxide is amphoteric and reacts with both acids and bases:
$Al_2O_3 + 6H^+ \rightarrow 2Al^{3+} + 3H_2O$
$Al_2O_3 + 2OH^- + 3H_2O \rightarrow 2Al(OH)_4^-$ [1]
Silicon dioxide is an acidic oxide and reacts with hot, concentrated alkali:
$SiO_2 + 2OH^- \rightarrow SiO_3^{2-} + H_2O$ [1]
The highest oxides of the remaining non-metallic elements react with water to form strongly acidic solutions:
$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$ [1]
$SO_3 + H_2O \rightarrow H_2SO_4$ [1]
$Cl_2O_7 + H_2O \rightarrow 2HClO_4$ [1]

Q3 **a** **i** +2 and +3/$Fe^{2+}$and Fe^{3+}; both 4s electrons are lost giving Fe^{2+} and one 3d electron is also lost to form Fe^{3+}. [2]
ii The presence of unpaired electrons; the 3d orbitals are split into two energy levels; electrons move between these energy levels; 3d electrons can absorb energy from light of visible wavelength. [3]

Q4 **a** A pale blue precipitate of hydrated copper(II) hydroxide, $Cu(OH)_2$. The precipitate reacts with excess ammonia to form a solution containing a deep blue complex ion:
$Cu^{2+} + 2OH^- \rightarrow Cu(OH)_2$
$Cu(OH)_2 + 4NH_3 \rightarrow [Cu(NH_3)_4{}^{2+}] + 2OH^-$ [4]
b **i** $4H^+ + 3MnO_4^{2-} \rightarrow 2MnO_4^- + MnO_2 + 2H_2O$ [2]
ii +6 and +7 and +4 [3]
iii Catalysis [1]

c The zinc ion, Zn^{2+}, has a $3d^{10}$ electron configuration (fully occupied 3d-orbitals). Hence it is not possible to have d–d transitions and it is thus colourless. [2]

Q5 **a** The chromium atom has the configuration $1s^22s^22p^63s^23p^63d^54s^1$.
Likely oxidation states are +3, + 6 and +1. [2]
b This is because the 3d and 4s orbitals have very similar energies in both the atoms and ions. Hence, transition metals can form ions of similar stability by losing different numbers of electrons. [2]
c $MnCO_3$, MnO_2 and $KMnO_4$. [3]
d In a complex, the ligand is bonded to the transition metal by a coordinate (or dative covalent) bond. The lone pair of electrons on the ligand is donated into the vacant (empty) orbitals in the transition metal ion to form the dative bond. [2]

Q6 **a** +3 and +2 [2]
b Six cyanide ligands are coordinated octahedrally to a central iron(II) ion. [2]
c A transition metal is a metal that forms compounds in which the metal has a partially filled 3d orbital. [1]
d Hemoglobin is an iron-containing protein found inside red blood cells. The iron forms a complex with oxygen, known as oxyhemoglobin, when the oxygen concentration is high. At low oxygen concentrations the oxygen is released and hemoglobin is re-formed. [2]

14 Bonding

Paper 1 IB questions and IB style questions

Q1 **A** Ethyne (C_2H_2) has a carbon–carbon triple bond; the carbon atoms in graphite form one carbon–carbon double bonds and two single bonds; in diamond each carbon atom is forming four bonds.

Q2 **D** The hydrogen atom does not undergo a process of hybridization during bond formation. The two s orbitals overlap to form a sigma bond.

Q3 **B** The fluoroethene molecule contains a carbon–carbon double bond and hence contains two sp^2 hybridized carbon atoms.

Q4 **D** Ethanal contains a doubly bonded oxygen which is sp^2 hybridized.

Q5 **C** The hydrogen peroxide molecule contains only single bonds. Only multiple bonds contain π bonds.

Q6 **A** End on overlap gives rise to σ bonds.

Q7 **D** The cyanide group contains a triple bond: one σ and two π bonds.

Q8 **A** The SF_6 molecule is octahedral because the sulfur atom is surrounded by six bonding pairs of electrons. The repulsion between these six negative charge centres is minimized by the adoption of an octahedral geometry.

Q9 **B** A tetrahedral geometry is observed.

Q10 **B** BeF_2 and CO_2 are both linear molecules.

Q11 **C** The tri-iodide ion is a linear ion with three lone pairs on the central iodine.

Q12 C The carbon–carbon bonds in benzene are intermediate in length between carbon–carbon single bonds and carbon–carbon double bonds. This is a consequence of resonance or π delocalization.

Q13 D The sulfur hexafluoride molecule is non-polar because the effects of the polar bonds cancel out vectorially.

Q14 A Three equivalent resonance structures can be drawn for the nitrate ion.

Q15 D The sulfur atom of the sulfur hexafluoride molecule has 10 electron pairs in its outer shell. All the noble gases (except helium) have eight electrons in their outer shells.

Q16 B It is represented by three equivalent resonance structures; the central nitrogen atom has a lone pair and forms one double bond and one single bond.

Q17 B In isolated carbon–carbon double bonds the hybridization is sp^2. Each carbon atom forms one π bond and two σ bonds.

Q18 C There is no equilibrium; the resonance structures do not exist. The molecule or ion is a hybrid.

Q19 B The molecule contains only single bonds: N–N and N–H.

Q20 D Carbon–carbon triple bonds are shorter than carbon–carbon double bonds which in turn are shorter than carbon–carbon single bonds.

Q21 C A difluoroethyne molecule has two C–F bonds. These are σ bonds. The triple bond contains two π bonds and one σ bond.

Q22 C The carbon atom is making four bonds: a double bond to the oxygen atom and two single bonds to the singly charged oxygen atoms.

Q23 D The total number of electrons around the chlorine atom is 10: seven from the chorine atom and one for each of the covalent bonds formed with the three fluorine atoms. The number of electron pairs is five: three bonding pairs and two lone pairs.

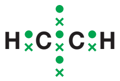
Q24 D The propanoate ion has two equivalent resonance structures involving the carboxylate group.

Q25 B The total number of electrons in the outer shell of chlorine is eight: seven electrons from the chlorine atom; one for each of the covalent bonds formed with the two fluorine atoms and the removal of one electron for the formation of the positive charge. The number of bonding pairs is four: two bonding pairs and two lone pairs.

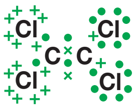
Paper 2 IB questions and IB style questions

Q1 a Each correct structure (with non-bonding electron pairs where relevant)

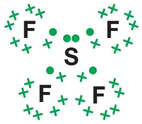
C_2H_2 linear/straight
Angle 180°



C_2Cl_4 (trigonal) planar
Angle 120°



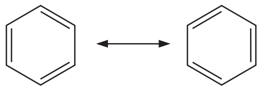
SF_4 K-shaped/based on trigonal bipyramid/seesaw
Angle 90°
Angle 120°



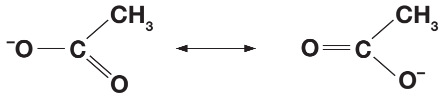
[10]

b sp (C_2H_2)
sp^2 (C_2Cl_4) [2]

c

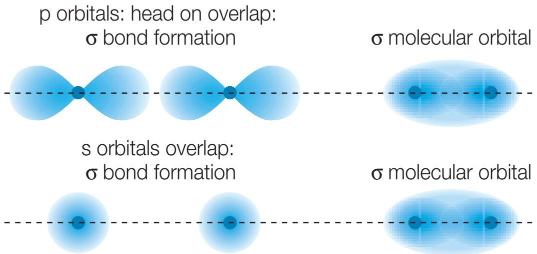


[2]

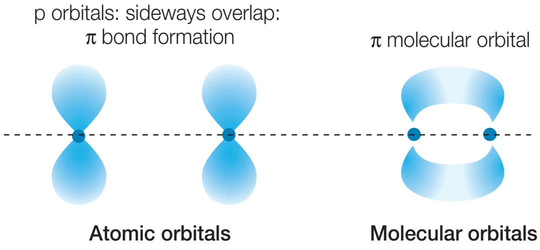


[2]

Q2 a Sigma bonds involve overlap of orbitals end-on/along the molecular axis (it has axial symmetry around axis joining the two nuclei)



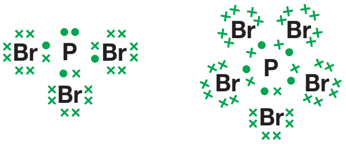
Pi bonds result from overlapping of parallel p orbitals/ sideways overlap



Double bond: a σ bond and a π bond
Triple bond: a σ bond plus two π bonds [4]

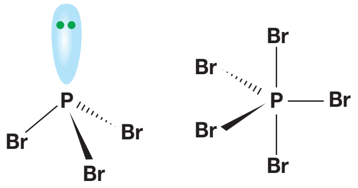
b Delocalization: when π electron pairs are not confined to two adjacent bonding atoms but extend over three or more atoms. [2]

Q3 a Lewis electron dot structures for PBr_3 and PBr_5:



[2]

Molecular shapes for PBr_3 and PBr_5:



[2]

b The bond angles present in PBr5 are 90°, 120° and 180°, respectively.
The bond angle in PBr_3 is approximately 107° since a lone pair of electrons (due to its more diffuse nature) produces more repulsion than a bonding pair, hence the bond angle is reduced below the tetrahedral bond angle of 109.5°. [4]

c

Ion	$[PBr_4]^+$
Number of electrons in valency shell of central phosphorus atom	(5 + 4 – 1) = 8 (adjusting for ionic charge)
Number of electron pairs contributing to basic molecular shape	4
Molecular shape	Tetrahedral

[1]

d The hybridizations of the phosphorus in phosphorus(III) bromide and the oxygen in water are sp^3 and sp^3, respectively. [2]

15 Energetics

Paper 1 IB questions and IB style questions

Q1 C Two moles of gas are converted into four moles of gas. Gases have very large values of absolute entropy relative to solids and liquids.

Q2 C Add the reverse of first equation with the second equation and simplify oxygen and cancel the phosphorus:

$P_4O_6(s) \rightarrow P_4(s) + 3O_2(g)$; $\Delta H^{\ominus} = 1600\ \mathrm{kJ\,mol^{-1}}$

$P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$; $\Delta H^{\ominus} = -3000\ \mathrm{kJ\,mol^{-1}}$

$P_4O_6(s) + 2O_2(g) \rightarrow P_4O_{10}(s)$; $\Delta H^{\ominus} = (1600\ \mathrm{kJ\,mol^{-1}}) + (-3000\ \mathrm{kJ\,mol^{-1}}) = -1400\ \mathrm{kJ\,mol^{-1}}$

Q3 D The lattice enthalpy is the energy needed to break one mole of an ionic compound into its component gaseous ions (under standard thermodynamic conditions).

Q4 A $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$, hence $\Delta G^{\ominus}$ will vary with temperature since T represents the absolute temperature. This reaction is not spontaneous at 298 K. The reaction involves an increase in entropy.

Q5 D A bond enthalpy refers to the breaking of one mole of covalent bonds in the gaseous state.

Q6 C Cooling the flask will decrease the average speed and hence kinetic energy of the hydrogen molecules. The degree of disorder will decrease and hence entropy decreases (only in the system).

Q7 B The formation of a lattice from its ions is a highly exothermic process. It is the reverse of lattice enthalpy. The process is more exothermic if small and highly charged ions are involved. Fluoride ions are smaller than bromide ions; lithium ions are smaller than sodium ions.

Q8 B A mole of gas is formed from a mole of a solid. Gases have relatively large values of absolute entropy.

Q9 B The charge on the lithium ion is +1; the charge on the strontium ion is +2. The higher the charge on the ions the greater the electrostatic forces of attraction and hence the higher the value of the lattice enthalpy.

Q10 D A negative value for the Gibbs free energy change means that the reaction is spontaneous under standard conditions. The reaction may be fast, slow, endothermic or exothermic.

Q11 A This is a decomposition reaction and hence endothermic. Two moles of gas are produced from one mole of solid and hence involves an increase in entropy.

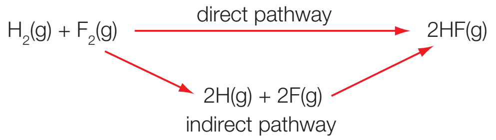
Q12 B The first electron affinity of fluorine is exothermic: the attraction between the incoming electron and the nucleus is greater than the electron–electron repulsion. The formation of an ionic lattice is exothermic: heat is released to the surroundings and the potential energy is lowered.

Q13 D A gas is much more disordered than a liquid and hence has a relatively high entropy value. Larger and more complex molecules have higher entropy values than smaller and less complex molecules (at the same temperature).

Q14 A Oxygen is not in its standard state. It is a gas a under standard thermodynamic conditions.

Q15 A The negative sign by convention indicates that heat is released to the surroundings.

Q16 D



Enthalpy of direct pathway = enthalpy of indirect pathway;
(2 × –268) = (–1130 + 435 + F–F);
–536 = –695 + F–F; F–F = 159

Q17 B Maximum entropy and minimum enthalpy both help to 'drive' the reaction forward. $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$ – to attain a negative value for $\Delta G^{\ominus}$, $\Delta H^{\ominus}$ must be negative and $\Delta S^{\ominus}$ must be positive.

Q18 D The production of a gas, a very disordered state of matter, results in an increase in entropy.

Q19 D The enthalpy of formation refers to the formation of one mole of a compound from its elements in their standard states under standard conditions. Sodium is a solid under these conditions and fluorine is a diatomic gas. All ionic substances are solids at room temperature.

Q20 C $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$. For $\Delta G^{\ominus}$ to be negative at high temperatures, $\Delta H^{\ominus}$ must be positive and $\Delta S^{\ominus}$ must be positive.

Q21 A The enthalpy of formation refers to the formation of one mole of a compound from its elements in their standard states under standard conditions. Carbon is a solid under these conditions and oxygen is a diatomic gas.

Q22 A $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$, hence if T changes then $\Delta G^{\ominus}$ changes. There is a decrease in entropy, 3 moles of gas is converted to 2 moles of gas – this favours the backward reaction. The forward reaction is exothermic.

Q23 B Bond breaking is always endothermic since forces of attraction need to be overcome.

Q24 C 2 moles of gas are converted to 2 moles of gas.

Q25 A $\Delta G^{\ominus}$ for the forward reaction is negative; $\Delta G^{\ominus}$ for the backward reaction is positive. At equilibrium the rates of the backward and forward reactions are equal and $\Delta G = 0$ (not $\Delta G^{\ominus}$).

Paper 2 IB questions and IB style questions

Q1 a i $\Delta H_f^{\ominus}$ is the standard enthalpy change of formation of a substance. [1]

It is the heat change (absorbed or released under constant pressure) when a mole of a compound is formed from its elements in their standard states. [1]

$S^{\ominus}$ represents the standard or absolute entropy. [1]

It is related to the disorder or randomness of particles. [1]

Related to standard conditions or 298 K (or 25 °C) and one atmosphere pressure. [1]

ii Δ was not included because $S^{\ominus}$ has absolute values; $S^{\ominus}$ values can be experimentally measured. [1]

iii $\Delta H_f^{\ominus}$ (Cu) = 0 (by definition an element is in its standard state) [1]

b i $\Delta H_f^{\ominus}$ reaction = $\Sigma\Delta H_f^{\ominus}$ products − $\Sigma\Delta H_f^{\ominus}$ reactants [1]

$\Delta H_f^{\ominus}$ reaction =[4 × (−242) + (−1084)] − (−2278) [1]

= +226 kJ mol⁻¹ [1]

An endothermic process or reaction needs heat energy. [1]

ii $\Delta S^{\ominus}$ reaction = $\Sigma\Delta S^{\ominus}$ products − $\Sigma\Delta S^{\ominus}$ reactants

= [4 × (189) + (150)] − (305) [1]

ΔS reaction = 601 J K⁻¹ mol⁻¹ or 0.601 kJ K⁻¹ mol⁻¹ [2]

(+) sign/value, thus products more disordered than reactants. [1]

iii Gibbs free energy (ΔG). [1]

Units: kJ mol⁻¹ [1]

c i At 25 ° Celsius, $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$

$\Delta G^{\ominus} = (+226)\,\text{kJ mol}^{-1} - (25 + 273)\,\text{K} \times \left(\frac{601}{1000}\right)\text{kJ mol}^{-1}\,\text{K}^{-1}$ [2]

$\Delta G^{\ominus} = +46.9\,\text{kJ mol}^{-1}$ [1]

As the value of the Gibbs free energy change, $\Delta G^{\ominus}$, is positive at 25 °C, the forward reaction is not spontaneous. However, this implies that the backward reaction is spontaneous. Hence, at 25 °C $CuSO_4.5H_2O(s)$ is the thermodynamically more stable compound. [2]

ii When both compounds have equal thermodynamic stabilities, the Gibbs free energy change will be zero, i.e. $\Delta G = 0$.

$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$

$0 = (+226\,\text{kJ mol}^{-1}) - \left(\frac{T \times 601}{1000}\right)\text{kJ mol}^{-1}\,\text{K}^{-1}$ [1]

$T = \frac{226\,\text{kJ mol}^{-1}}{0.0601\,\text{kJ K}^{-1}\,\text{mol}^{-1}} = 376\,\text{K}\ (103\,°\text{C})$ [1]

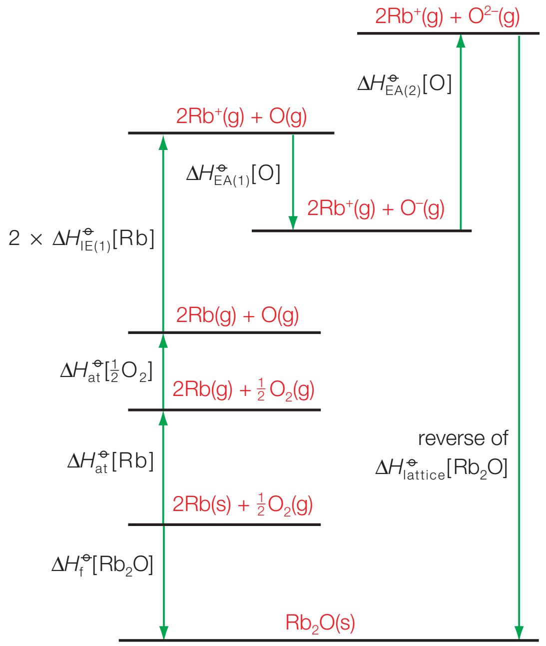
Hence, $CuSO_4.4H_2O(s)$ is more thermodynamically stable above 103 °C. [1]

Q2 a The lattice enthalpy is the energy needed to break down one mole of ionic solid into gaseous ions (separated to a infinite distance) under standard thermodynamic conditions (1 atm and 25 °C). [2]

b The greater the charge on the ions, the higher the value of the lattice enthalpy. Larger charges produce greater electrostatic forces of attraction between oppositely charged ions. [1]

The smaller the ionic radius, the higher the value of the lattice enthalpy. Smaller ions can approach each other more closely and hence experience greater electrostatic forces of attraction. [1]

c Born–Haber cycle for rubidium oxide, Rb_2O



[2]

(2 × 80.9) + 249.2 + (2 × 403.0) − 146.1 + 795.5 +

$-\Delta H_{lattice}\,(Rb_2O) = -339.0$

$-\Delta H_{lattice}\,(Rb_2O) = -339.0 - (1866.4) = -2205.4\,\text{kJ mol}^{-1}$

$\Delta H_{lattice}\,(Rb_2O) = -2205.4\,\text{kJ mol}^{-1}$ [2]

d The ionic model is a good model for rubidium oxide – the bonds have very little covalent character. The ionic model is not a good model for silver bromide – the bonds are polar covalent. [2]

e The sodium ion is smaller than that of rubidium. It can therefore approach the oxide more closely and exert higher electrostatic forces of attraction. [2]

Q3 a The enthalpy change that occurs when one mole of a pure compound is formed under standard thermodynamic conditions from its elements in their standard states. [2]

b The enthalpy change that occurs when one mole of a pure compound undergoes complete combustion in the presence of excess oxygen under standard conditions. [2]

c Hess's law states that the total enthalpy change for a reaction is independent of the route taken. It depends only on the initial and final states. [1]

d

$$\begin{tikzcd}
3C(s) + 4H_2(g) + 5O_2(g) \arrow[rr, "\Delta H_1"] \arrow[dr, "\Delta H_2"'] & & C_3H_8(g) + 5O_2(g) \\
& 3CO_2(g) + 4H_2O(l) \arrow[ur, "\Delta H_3"'] &
\end{tikzcd}$$

According to Hess's law, $\Delta H_1 = \Delta H_2 - \Delta H_3$

$\Delta H_1 = \Delta H_f^{\ominus}[C_3H_8(g)]$

$\Delta H_2 = 3 \times \Delta H_c^{\ominus}[C_{graphite}] + 4 \times \Delta H_c[H_2(g)]$

$\Delta H_2 = 3 \times (-393\,\text{kJ mol}^{-1}) + 4 \times (-286\,\text{kJ mol}^{-1})$

$= -2323\,\text{kJ mol}^{-1}$

$\Delta H_3 = [C_3H_8(g)] = -2220\,\text{kJ mol}^{-1}$

Hence:

$\Delta H_f^{\ominus}[C_3H_8(g)] = (-2323\,\text{kJ mol}^{-1}) - (-2220\,\text{kJ mol}^{-1})$

$= -103\,\text{kJ mol}^{-1}$ [4]

16 Kinetics

Paper 1 IB questions and IB style questions

Q1 A The negative sign indicates that the concentration decreases with time. The half indicates the stoichiometric relationship between the rate of disappearance of N_2O_5 and the rate of appearance of O_2.

Q2 C A first-order reaction is characterized by a directly proportional relationship between initial rate and concentration. The reaction is zero-order with respect to hydroxide ions: a change in concentration has no effect on the rate.

Q3 D A catalyst provides an alternative energy pathway with a different activation energy.

Q4 D The reaction is zero-order with respect to iodine. It does not appear in the rate equation, hence its concentration has no effect on the rate.

Q5 D First-order reactions have a constant half-life. Two half-lives are required for $\frac{3}{4}$ of the N_2O_5 to decompose in both containers.

Rate = $[N_2O_5]$; $t_{1/2} = \frac{\ln 2}{k}$;

from the formula, the half-life $t_{1/2}$ is independent of the concentration of reactant (N_2O_5) (cf. radioactive decay).

Q6 D The overall order is the sum of the individual orders: 1 + 2.

Q7 C Since A is first order when the concentration is tripled, the rate is also tripled. A first-order reaction has a directly proportional relationship between initial rate and concentration. Since B is second order when the concentration is doubled, the rate is quadrupled. (×3) × (×4) = ×12.

Q8 A Rate constants increase with temperature, generally in an exponential manner. Rate constants are unaffected by changes in concentrations of reactants.

Q9 D A mechanism cannot be deduced from a balanced equation for the overall reaction.

Q10 C I and IV are answered by kinetics; II and II are answered by thermodynamics. There is no correlation between rate and spontaneity.

Q11 D Rate = $k[A]^2$; $0.30 = k \times 0.20^2$; $k = 7.5$

Q12 B When the concentration of A is halved, the rate is also halved (from $0.04\,s^{-1}$ to $0.02\,s^{-1}$). A first-order reaction has a directly proportional relationship between rate and concentration. Rate is defined for this reaction as the reciprocal of time.

Q13 B Response A has the units of a zero-order reaction. Response D has the units of a first-order reaction. However, for an equation with an overall order of two.

Rate = $k[A][B]$ and hence $k = \frac{\text{rate}}{[A] \times [B]}$

Substituting the units for rate and concentration gives $\text{mol dm}^{-3}\,\text{s}^{-1}/(\text{mol dm}^{-3})^2$ which simplifies to $\text{mol dm}^{-3}\,\text{s}^{-1}/\text{mol}^2\,\text{dm}^{-6}$ before cancelling to $\text{mol}^{-1}\,\text{dm}^3\,\text{s}^{-1}$.

Q14 D A bimolecular step results in second-order kinetics. Rate = $k[I] \times [I]$. Bond formation is exothermic: energy is released; the iodine molecule is more energetically stable than iodine atoms.

Q15 D *A* is called the Arrhenius constant (or pre-exponential factor), and relates to the geometric requirements of the collisions.

Q16 D The powers in the rate expression are both two, so the orders are both two. The overall order is the sum of the individual orders (2 + 2). Since they both have the same orders then the changes in concentration have the same effect on the rate.

Q17 B All graphs are characteristic of a first-order reaction.

Q18 D The pressure increases by a factor of ×4. Since the reaction is second-order the overall rate increases by (×4) × (×4) = ×16.

Q19 A A comparison of the first two experiments, where the NO concentration is doubled (and the hydrogen concentration is kept constant) reveals that the rate is increased by a factor of four, that is, quadrupled. Hence, the order with respect to NO is two. A comparison of the last two experiments, where the H_2 concentration is doubled (and the NO concentration kept constant) reveals that the rate is increased by a factor of two, that is, doubled. Hence the order with respect to H_2 is one.

Q20 B The reaction is second order with respect to A and zero order with respect to B. Mechanism B is bimolecular and second order with respect to A.

Paper 2 IB questions and IB style questions

Q1 a i $t_{1/2} = \frac{\ln 2}{k}$; $k = \frac{0.693}{1.62 \times 10^4\,\text{s}}$; $k = 4.28 \times 10^{-5}\,\text{s}^{-1}$ [1]

ii $\ln k = \ln A - \frac{E_a}{RT}$; $\frac{E_a}{RT} = \ln A - \ln k$; $E_a = (\ln A - \ln k)RT$ [2]

$E_a = [3.219 - (-10.059)] \times 8.31 \times 1107 = 122\,\text{kJ}$

Q2 a Step 1 since it is the slowest. [1]

b Step 1 is the slowest step, therefore it has a higher activation energy relative to step 2. [2]

c Rate = $k[NO_2][F_2]$ [1]

Q3 a Two [1]

b One [1]

c Rate = $k[A(aq)]^2 \times [B(aq)]$. [1]

d $1.00 = k(0.4)^2 \times (0.4)$; $k = 15.625$ [1]

17 Equilibrium

Paper 1 IB questions and IB style questions

Q1 C There is no change in the proportions of X and Y after 10 minutes.

Q2 D $\frac{(5.0)^2}{0.5 \times 0.5} = 100$

Q3 D The only change that will alter the position of this physical equilibrium is a change in temperature.

Q4 B The mercury is depressed lowest in tube III, followed by I and then II.

Q5 C The curve is steepest at 0.5 atmospheres for substance 1.

Q6 A K_c increases with increased temperature; so the forward reaction must be endothermic.

Q7 C A catalyst lowers the activation energy of both the forward and reverse reactions. ΔH for the reaction is unchanged.

Q8 B A catalyst increases the rates of the forward and reverse reactions equally.

Q9 C The amount of water vapour is greater at the higher temperature as evaporation is endothermic.

Q10 B This statement is an essential characteristic of a dynamic equilibrium.

Q11 D Removal of ammonia shifts the equilibrium to the right so that more ammonia is produced to replace what has been removed.

Q12 D A catalyst lowers the activation energy of both the forward and reverse reactions. ΔH for the reaction is unchanged.

Q13 A For a reaction that goes almost to completion K_c is very large.

Q14 D The value of K_c for an equilibrium reached from the opposite direction is the reciprocal of the original value.

Q15 B Feed data into equilibrium expression.

$$K_c = 5.0 \times 10^{-3} = \frac{[NO_2]^2}{[N_2O_4]} = \frac{[NO_2]^2}{5 \times 10^{-1}}$$

So $[NO_2]^2 = 25 \times 10^{-4}$ and $[NO_2] = 5.0 \times 10^{-2}\,mol\,dm^{-3}$

Q16 A Initially the rate of evaporation is higher until equilibrium is set up.

Q17 B A high boiling point and ΔH_{vap} are consequences of strong intermolecular forces between molecules in a liquid.

Q18 B Strong intermolecular forces in a liquid result in a low vapour pressure. Strong intermolecular forces would also result in high surface tension, high viscosity and a high value of ΔH_{vap}.

Q19 D The equilibrium constant is greater at the higher temperature, so there is more ionization at equilibrium. Therefore the forward reaction (the ionization of water) is endothermic.

Q20 D The addition of extra H_3O^+ ions (a product of the reaction) would result in the equilibrium shifting to the left; so there would be a decrease in [HBrO].

Q21 B The vapour pressure exerted by a liquid is independent of its surface area. An increase in surface area increases the rate of condensation and evaporation equally.

Q22 A The vapour pressure exerted by a liquid depends only on the identity of the liquid and its temperature. The presence of a non-volatile solute will lower the vapour pressure of the methylbenzene.

Q23 D Entropy (disorder) increases with temperature; the rate of evaporation and hence vapour pressure increases with temperature; the average kinetic energy of gas particles increases with temperature.

Q24 A

	$2HI(g)$	$\rightleftharpoons$	$H_2(g)$	+	$I_2(g)$
Initial:	1 mol				
At equilibrium:	0.76 mol		0.12 mol	+	0.12 mol

(NB: 0.76 + 0.12 + 0.12 = 1 mol; the hydrogen and iodine are in a 1 : 1 molar ratio.)

$$K_c = \frac{[H_2] \times [I_2]}{[HI]^2} = \frac{0.12 \times 0.12}{(0.76)^2} = 0.0249$$

Q25 A

$$K_c = \frac{[PCl_3] \times [Cl_2]}{[PCl_5]}$$

$$5 \times 10^{-12} = \frac{[Cl_2]^2}{0.050}$$

$$[Cl_2] = \sqrt{0.050 \times 5 \times 10^{-12}} = 5.0 \times 10^{-7}$$

Paper 2 IB questions and IB style questions

Q1 a i There will be no change in pressure. [1]
ii The pressure will decrease. [1]
b i The temperature will increase. [1]
ii The methanol concentration will increase. [1]
c i $K_c = \frac{[CH_3OH]}{[CO][H_2]^2}$
ii

	$CO(g)$	+	$2H_2(g)$	$\rightleftharpoons$	$CH_3OH(g)$	
Initial amount:	1.00		2.00		–	
Amount at eq./mol	1.00 – 0.85		2.00 – 1.70		0.85	[1]
Eq. conc.	$\frac{0.15}{0.45}$		$\frac{0.30}{0.45}$		$\frac{0.85}{0.45}$	[1]
	= 0.333 mol		= 0.667 mol		= 1.889 mol	

$K_c = \frac{1.888}{0.333 \times 0.667^2} = 12.7$ [1]

iii Side reactions *or* leaks in the system *or* not operating under equilibrium conditions *or* operating at a higher temperature *or* the product might be collected before equilibrium is reached. [1]
iv No effect on K_c (it just speeds up the reaction). [1]

Q2 a i One mark for any two of the following (max [2]):
Some mention of a reversible reaction/$\rightleftharpoons$ etc. [1]
Indication that rate of forward reaction = rate of reverse reaction [1]
At equilibrium no net change in concentration of reactants and products [1]
Closed system [1]
ii $K_c = \frac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]}$ [1]
iii $[C_2H_5OH] = [CH_3COOH] = 0.33\,mol\,dm^{-3}$
$K_c = 4.1$ [2]
b K_w increases with temperature rise/$[H^+]$ or $[OH^-]$ rises as temperature rises [1]; equilibrium shifts to right (as temperature increases) [1]; endothermic reaction [1]. [3]

Q3 a A; E [2]
b i After 15 s (product) = 0.37 ($mol\,dm^{-3}$);
rate = $0.37\,mol\,dm^{-3}/15\,s = 0.025\,mol\,dm^{-3}\,s^{-1}$ [3]
ii At equilibrium/rates of forward and reverse reactions are equal/$\Delta G = 0$ [1]

Q4 a $K_c = [I_{2\,(water)}]/[I_{2\,(hydrocarbon)}]$ [1]
b No change [1]
c $[I_{2\,(hydrocarbon)}]$ will decrease [1]
I_2 combining with I^- ions removes I_2 [1]
So equilibrium in a shifts to the right, so less I_2 dissolved in the hydrocarbon layer. [1]

Q5 a

	C_5H_{10}	+	CH_3COOH	$\rightleftharpoons$	$CH_3COOC_5H_{11}$	
Initial no. of mol	0.02		0.01		–	
Amount at eq./mol	(0.02 – 0.009)		(0.01 – 0.009)		0.009	
	0.011 mol		0.001 mol			[2]

b $K_c = \frac{[CH_3CO_2C_5H_{11}]}{[C_5H_{10}]\,[CH_3CO_2H]}$ [1]

c $[C_5H_{10}] = \frac{0.011}{0.6} = 0.0183\,\text{mol dm}^{-3}$

$[CH_3COOH] = \frac{0.001}{0.6} = 1.67 \times 10^{-3}\,\text{mol dm}^{-3}$

$[CH_3COOC_5H_{11}] = \frac{0.009}{0.6} = 0.015\,\text{mol dm}^{-3}$ [3]

d $K_c = \frac{0.015}{0.0183} \times 1.67 \times 10^{-3} = 491$ [2]

18 Acids and bases

Paper 1 IB questions and IB style questions

Q1 C $[H^+(aq)] = 10^{-pH}$; $[H^+(aq] = 10^{-10}$; $K_w = [H^+(aq)] \times [OH^-(aq)]$
$10^{-14} = 10^{-10} \times [OH^-(aq)]$; $[OH^-(aq)] = 10^{-4}$

Q2 B Sodium carbonate can be made by titrating sodium hydroxide (a strong alkali) with carbonic acid (a weak acid). Salts of a strong base and a weak acid are alkaline. The carbonate ions undergo hydrolysis with water molecules to release hydroxide ions:
$CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + HCO_3^-(aq)$

Q3 A The pK_a is the negative logarithm to the base 10 of the acid dissociation constant, K_a.

Q4 B The solution prepared by I will contain sodium ethanoate, the salt of a weak acid, and a weak acid, ethanoic acid. There is no chemical reaction between a salt and dilute acid. The solution prepared by II will contain sodium ethanoate (formed by neutralization) and excess ethanoic acid.

Q5 A The concentrations of HIn(aq) and In^-(aq) will depend on the value of pK_a. No indicator is suitable for a titration involving weak acids and bases since there is no sharp end-point. A strongly acidic solution has a high concentration of hydrogen ions and hence will shift the equilibrium to the left and favour the backward reaction.

Q6 D Ammonium chloride solution is acidic:
$NH_4^+(aq) + OH^-(aq) \rightleftharpoons NH_3(aq) + H_2O(l)$;
ammonium ethanoate is close to being neutral:
$NH_4^+(aq) + CH_3COO^-(aq) + 2H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq) + CH_3COOH(aq) + OH^-(aq)$
and potassium ethanoate is alkaline:
$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$

Q7 B $pH = -\log_{10}[H^+(aq)]$; $pH = -\log_{10}(2.62 \times 10^{-3}) = 2.58$

Q8 A $K_a = \frac{[H^+](aq)]^2}{[HA(aq)]}$; $K_a = \frac{(6.3 \times 10^{-5})^2}{0.1} = 3.98 \times 10^{-8}$

Q9 C $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
Amount of NaOH(aq) = 2.2×10^{-4} mol; amount of HCl(aq) = 3.5×10^{-4} mol. Amount of HCl(aq)/H^+(aq) after neutralization (in 45 cm^3 of solution) = 1.3×10^{-4} mol; amount of H^+(aq) in 1 dm^3 = 2.8×10^{-3}; $pH = -\log_{10}[H^+(aq)] = 2.54$

Q10 D Potassium carbonate and potassium cyanide are both formed by the neutralization of a strong alkali by a weak base, hence their solutions will be alkaline. Sodium nitrate is formed by the neutralization of a strong acid and strong alkali, hence its solution will be neutral. Iron(III) ions from iron(III) chloride react with water to release hydrogen ions:
$[Fe(H_2O)_6]^{3+}(aq) \rightarrow H^+(aq) + [Fe(H_2O)_5OH]^{2+}$

Q11 C $[\text{pyridineH}^+] \times \frac{[OH^-(aq)]}{[\text{pyridine}]}$;
$\frac{[OH^-(aq)]^2}{0.05\,\text{mol dm}^{-3}} = 1.4 \times 10^{-9}$;
$pOH(aq) = -\log_{10}(8.36 \times 10^{-6}) = 5.08$; pH = 8.92

Q12 C Alcohols do not react with water to release ions. They are molecular covalent substances and dissolve in water to release hydrated molecules.

Q13 C An alkaline buffer is formed by a salt and a weak base.

Q14 A Weak acids have small K_a values, but large values of pK_a. The conjugate bases of weak acids are strong bases. They are poor conductors due to the presence of a low concentration of ions.

Q15 D The concentration of hydrogen and hydroxide ions remain in a 1 : 1 relationship as the temperature changes.

Q16 B $K_b = \frac{[CH_3CH_2NH_3^+(aq)] \times [OH^-(aq)]}{[CH_3CH_2NH_2(aq)]}$

Q17 C K_a is an equilibrium constant and hence will vary with temperature.

Q18 A $pH = pK_a + \log_{10}\frac{[\text{salt}]}{[\text{acid}]}$; $pH = 5.045 + \log_{10}\frac{[0.002]}{[0.01]} = 4.346$

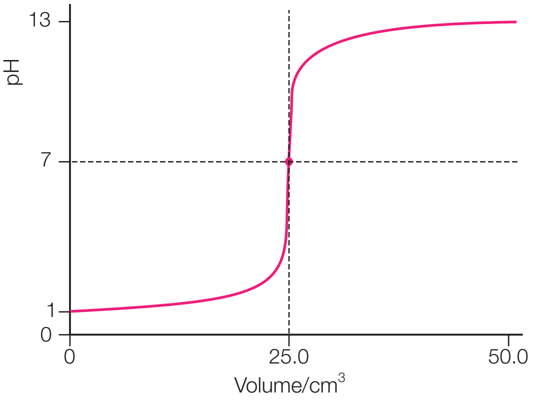
Q19 D A and B form an acidic buffer; C is not a buffer.

Q20 C A titration curve for a strong acid/weak base starts at a low pH and will level off to a pH above 7 but well below 14. There is a vertical region centred below 7.

Paper 2 IB questions and IB style questions

Q1 a i $pH = -\log[H^+(aq)]$ [1]

ii The curve should include the following: starting pH = 1; equivalence point: 25.0 cm^3 of NaOH: pH at equivalence point = 7; pH to finish =12–13 [4]



iii $K_a = 10^{-4.76} = 1.74 \times 10^{-5}$

$K_a = \frac{[H^+(aq)]^2}{[CH_3COOH]}$

$1.74 \times 10^{-5} = \frac{[H^+(aq)]^2}{0.100}$

$[H^+(aq)] = \sqrt{1.74 \times 10^{-6}} = 1.32 \times 10^{-3}$ (mol dm^{-3})
starting pH = 2.88;
pH at equivalence point: 8–9 [5]

b i HIn is a weak acid:
$HIn \rightleftharpoons H^+ + In^-$ and two colours indicated
In acid the equilibrium moves to the left or vice versa. [3]

ii Phenolphthalein/phenol red/bromothymol blue; the colour change of indicator occurs within the range of pH at the equivalence point/on vertical part of graph. [2]

c i Specific examples of weak base and its salt/specific strong acid and weak base, e.g. NH_3 and NH_4Cl. [1]

ii pH changes very little/most of the acid is neutralized by base:
e.g. $NH_3 + H^+ \rightarrow NH_4^+$/$NH_4OH + H^+ \rightarrow NH_4^+ + H_2O$ [2]

d Brønsted–Lowry acid: a proton donor
Lewis acid: an electron pair acceptor
Brønsted–Lowry acid: any suitable equation
Lewis acid: BF_3/$AlCl_3$/transition metal ions that form complex ions with ligands: for example
$BF_3 + NH_3 \rightarrow BF_3NH_3$/$Cu^{2+} + 4NH_3$
$\rightarrow [Cu(NH_3)_4]^{2+}$/$AlCl_3 + Cl^- \rightarrow AlCl_4^-$
(or any suitable equation) [5]

e acidic
$[Al(H_2O)_6]^{3+}$ is (weak) due to the formation of H^+
$[Al(H_2O)_6]^{3+} \rightleftharpoons [Al(H_2O)_5(OH)]^{2+} + H^+$ [2]

Q2 a Brønsted–Lowry acids: CH_3CH_2COOH and H_3O^+ [2]
Brønsted–Lowry bases: $CH_3CH_2COO^-$ and H_2O [2]

b Bromophenol blue is blue and phenol red is yellow. [1]
The pH of 4.9 is above the range of bromophenol blue: the indicator shows its alkaline colour. [1]
The pH of 4.9 is below the range of phenol red: the indicator shows its acidic colour. [1]

c $K_a = \dfrac{[CH_3COO^-(aq)][H^+(aq)]}{[CH_3COOH(aq)]}$ [1]

$[H^+(aq)] = \dfrac{1.74 \times 10^{-5} \times 0.0500}{0.200} = 4.35 \times 10^{-5}\,\text{mol dm}^{-3}$ [1]

$\text{pH} = -\log_{10}[H^+(aq)] = 4.36$ [1]

Q3 a A buffer solution is an aqueous solution that resists a change in pH when a small amount of a strong acid or base is added.
A buffer contains a weak acid (HA) and its conjugate base (A^-) or a weak base and its conjugate acid.
If hydrogen ions, $H^+(aq)$, are added, these will react with the base of the buffer. The hydrogen ions are removed as they combine with A^- and the equilibrium shifts to the left. Hence, there is only a very small change in the concentrations of HA and A^-. The strong acid (H^+) is replaced by a weak acid, HA. [5]

b Amount of sodium benzoate $= \dfrac{7.2\,\text{g}}{144\,\text{g mol}^{-1}} = 0.05$ moles

The calculation assumes that all of the anion concentration is derived from the salt only (as that from the acid is too small to consider). It also assumes that the ionization of the acid is so small it can be ignored. Standard conditions (1 atm and 298 K) are also assumed. [2]

$6.3 \times 10^{-5} = \dfrac{[H^+(aq)] \times 0.05}{0.03}$; $[H^+(aq)] = 3.78 \times 10^{-5}\,\text{mol dm}^{-3}$

$\text{pH} = -\log_{10}[H^+(aq)] = 4.42$ [2]

c A weak acid is partially or incompletely ionized or dissociated. A monoprotic acid donates one hydrogen ion per molecule. [2]

Q4 a Acidic because it is a hydrogen ion donor and basic because it is a hydrogen ion acceptor. Such substances are described as amphoteric. [2]

b Oxide ion, O^{2-} [1]

c pH is the negative logarithm to the base of ten of the hydrogen ion concentration. $\text{pH} = -\log_{10}[H^+(aq)]$; pH of pure water (at 25 °C) is 7. [2]

d i $K_w = [H^+(aq)] \times [OH^-(aq)]$ [1]

ii Water is more dissociated or ionized at a higher temperature. Le Châtelier's principle implies that endothermic reactions are favoured by increases in temperature, hence the dissociation or ionization of water is endothermic. [3]

Q5 a Salt hydrolysis occurs when one or both ions from a salt react with water to release hydrogen or hydroxide ions. [2]

b i Hydrated iron(III) ions are acidic and release hydrogen ions from water:
$[Fe(H_2O)_6]^{3+}(aq) \rightarrow [Fe(H_2O)_5OH]^{2+}(aq) + H^+(aq)$
The solution will be acidic [2]

ii Potassium nitrate is formed by the neutralization of a strong alkali, potassium hydroxide, by a strong acid, nitric acid. Nitrate and potassium ions do not undergo hydrolysis with water molecules. The solution will be neutral. [2]

iii The carbonate ion is a weak base and reacts with water molecules to form hydroxide ions:
$CO_3^{2-}(aq) + H_2O(l) \rightarrow OH^-(aq) + HCO_3^-(aq)$
The solution will be alkaline. [2]

c $K_b = 10^{-4.75} = 1.78 \times 10^{-5}$

$K_b = \dfrac{[NH_4^+(aq)] \times [OH^-(aq)]}{[NH_3(aq)]}$

$[OH^-(aq)] = \sqrt{K_b \times [NH_3(aq)]} = \sqrt{1.78 \times 10^{-5} \times 0.120}$
$= 1.46 \times 10^{-3}$; $\text{pOH} = -\log[OH^-(aq)] = 2.84$ [4]

19 Oxidation and reduction

Paper 1 IB questions and IB style questions

Q1 D A change in the size of the anode will change the current, but not the voltage (which is determined by the nature of the electrodes).

Q2 C $E^\ominus_{\text{cell}} = (+0.14\,\text{V}) + (-0.13\,\text{V}) = +0.01\,\text{V}$

Q3 B $\Delta G^\ominus = -zFE^\ominus_{\text{cell}}$; the more positive the value of the cell potential, the more negative the value of the Gibb's free energy change.

Q4 A The greatest voltage is generated by the largest difference in standard electrode potentials.

Q5 C The equilibrium lies almost exclusively on the right-hand side. A positive value of $E^\ominus_{\text{cell}}$ is accompanied by a negative value for ΔG.

Q6 C $Ag^+ + e^- \rightarrow Ag$; $1 \times 108 = 108\,\text{g}$;

$Ni^{2+} + 2e^- \rightarrow Ni$; $\dfrac{1}{2} \times 59 = 29.5\,\text{g}$

$Cr^{3+} + 3e^- \rightarrow Cr$; $\dfrac{1}{3} \times 52 = 17.3\,\text{g}$

Q7 A $Ni(s) + Mg^{2+}(aq) \rightarrow Mg(s) + Ni^{2+}(aq)$;
$E^\ominus_{\text{cell}} = (+0.23\,\text{V}) + (-2.36\,\text{V}) = -2.13\,\text{V}$. Since $E^\ominus_{\text{cell}}$ is negative the reaction is not spontaneous.

Q8 B $E^\ominus_{\text{cell}} = (+0.141\,\text{V}) + (0.518\,\text{V}) = +0.659\,\text{V}$

Q9 D $E^\ominus_{\text{cell}} = (+0.40\,\text{V}) + (-0.23\,\text{V}) = +0.17\,\text{V}$

Q10 C $E^\ominus_{\text{cell}} = (+0.77\,\text{V}) + (0.23\,\text{V}) = +1.00\,\text{V}$

Q11 B These changes result in the equilibrium shifting to the right and increasing the cell voltage.

Q12 B The cell reaction is: $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

Q13 A No gas production occurs; the cathode gains copper; the anode loses copper, but these processes are in equilibrium and there is no change in the electrolyte concentration.

Q14 A It only deposits one mole of element if the element is formed from unipositive ions.

Q15 B Sodium ions remain in solution; they are not discharged.

Q16 A Standard electrode potentials are measured when both solutions in the voltaic cell have concentrations of 1.00 mol dm^{-3}.

Q17 B Positive and negative ions flow through the salt bridge. Electrons only flow through the external circuit (wires and voltmeter). The addition of hydrochloric acid will result in the formation of a precipitate of silver chloride and hence a decrease in the concentration of aqueous silver ions in the half-cell.

Q18 B $E^{\ominus}_{\text{cell}}$ = (–0.13 V) + (–0.20 V) = +0.07 V. Since $E^{\ominus}_{\text{cell}}$ is positive the reaction is spontaneous.

Q19 B The mass of gold deposited depends on the number of coulombs passed – this is a product of the current (A) and time (s).

Q20 A

Mg^{2+}	+	$2e^-$	→	Mg
0.5 mol	+	1.0 mol		0.5 mol (12 g)
Na^+	+	e^-	→	Na
1.0 mol	+	1.0 mol		1.0 mol (23 g)

Paper 2 IB questions and IB style questions

Q1 a Oxidation number of copper on left-hand side is 0; oxidation number of nitrogen on left-hand side is +5. [2]
Oxidation number of copper on right-hand side is +2; oxidation numbers of nitrogen on right-hand side are +5 and +2. [2]
$3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(l)$ [1]

b i The two electrodes must be dipping into aqueous solutions of their ions. The two electrodes must be connected via a salt bridge and an external circuit consisting of wire and a voltmeter. [3]
ii Hydrogen gas at 1 atm pressure passed over a platinum electrode. The temperature of all chemicals is 298 K and all solutions have a concentration of 1 mol dm^{-3}. [3]

c i $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$; $Mn(s) \rightarrow Mn^{2+}(aq) + 2e^-$ [2]
Manganese is oxidized. [1]
Nickel(II) ions are the oxidizing agent. [1]
ii Manganese is the anode. [1]
The electron flow is from manganese to nickel. [1]
iii $E^{\ominus}_{\text{cell}}$ = (–0.23 V) + (1.18 V) = +0.95 V [1]
ΔG is negative. [1]

d Gold is formed at the cathode (negative electrode). [1]
Oxygen is formed at the anode (positive electrode). [1]

Q2 a i Oxidation half-reaction: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$
Reduction half-reaction:
$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ or $2H^+ + 2e^- \rightarrow H_2$ [2]
ii Potassium has a highly negative value for its standard reduction electrode potential. Water molecules or hydroxide ions are preferentially reduced since smaller amounts of energy are required for their discharge. (Any potassium atoms formed would immediately react with water to form sodium ions.) [1]
iii Molten or fused state. [1]

b Hydrogen, H_2, and oxygen, O_2.
$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ or $4OH^- \rightarrow 2H_2O + O_2 + 4e^-$
Water or hydroxide ions are preferentially oxidized instead of fluoride ions.
This is because the standard reduction electrode potential is highly negative. Fluorine molecules are strongly oxidizing; fluoride ions are very weakly reducing. [4]

Q3 a Cathode (negative electrode): $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ [1]
Anode (positive electrode): $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ [1]
b There is no change in the intensity of colour since the rate of removal of copper(II) ions at the cathode is balanced by the rate of formation of copper(II) ions. at the anode. [1]
There is no change in the pH since hydrogen and hydroxide ions are not involved in the discharge reactions. [1]
c $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$ [2]
d Charge passed = 0.360 A × 600 s = 216 C [1]

Amount of electrons = $\dfrac{216\,\text{C}}{96\,485\,\text{C mol}^{-1}} = 2.24 \times 10^{-3}$ mol [1]

2 moles of electrons are required for the discharge of one mole of copper atoms. [1]

Mass of copper deposited $= 0.50 \times 63.55\,\text{g mol}^{-1} \times 2.24 \times 10^{-3}\,\text{mol}$
$= 0.0711\,\text{g}$ [1]

Q4 a An oxidizing agent gains electrons. [1]
b The current flow is from the aluminium electrode to the nickel. [1]
The aluminium electrode is negative; the nickel electrode is positive. [1]
c i $Mg(s) + Ni^{2+}(aq) \rightarrow Ni(aq) + Mg^{2+}(aq)$ [2]
ii $E^{\ominus}_{\text{cell}}$ = (2.36) + (–0.23) = +2.13 V [1]
d Time = 8.00 × 60 = 480 s [1]
Charge = 8.00 × 1.00 × 3600 = 28 800 C [1]

Number of moles = $\dfrac{28\,800\,\text{C}}{96\,480\,\text{C mol}^{-1}} = 0.2985$ mol [1]

$\dfrac{0.2985\,\text{mol}}{6} = 0.04975$ mol [1]

Amount of chromium atoms = 0.05 mol [1]

20 Organic chemistry

Paper 1 IB questions and IB style questions

Q1 A Tertiary halogenoalkanes hydrolyse via an S_N1 mechanism which implies first-order kinetics. Its hydrolysis rate is higher because it passes through a carbocation intermediate rather than a high-energy transition state.

Q2 D 2-Chlorobutane has a chiral carbon atom which is attached to an ethyl group, a hydrogen atom, a chlorine atom and a methyl group.

Q3 D Primary amides have the general formula $RCONH_2$. D is ethanamide.

Q4 A The *bromo-* group was replaced or substituted by an –OH group; the alcohol is then oxidized – it gains oxygen and is converted to an ester (general formula: RCOOR′).

Q5 A Ethanol can be oxidized to ethanal (distillation) and then to ethanoic acid (refluxing).

Q6 B Proteins are condensation polymers of amino acids.

Q7 B $CH_3COOH + HOCH_3 \rightarrow CH_3COOCH_3 + H_2O$

Q8 C A racemic mixture contains equal amounts of a pair of enantiomers (optical isomers) of the same molecule.

Q9 D The two methyl groups can be arranged *cis* and *trans* to each other.

Q10 C An ester with the general formula R–COO–R′ is named by naming R′ as an alkyl group and then R–COO as a carboxylate group.

Q11 D Nucleophiles react via the donation of a pair of electrons and formation of a dative bond with an electron-deficient centre.

Q12 D There are no chiral carbon atoms.

Q13 D $CH_3CHBrCH_3 \rightarrow H_3CHC{=}CH_2 + HBr$

Q14 D The iodine–carbon bond is long and weak and hence iodoalkanes are the most reactive. The chlorine–carbon bond is short and strong and hence chloroalkanes are the least reactive.

Q15 B $(CH_3)_2CHX + CN– \rightarrow (CH_3)_2CHCN + X–$

Q16 C Optical isomers rotate plane-polarized light in opposite directions.

Q17 B Propanenitrile undergoes reduction to give a primary amine:
$CH_3CH_2CN + 2H_2 \rightarrow CH_3CHCH_2NH_2$

Q18 C Water is eliminated during the polymerization process.

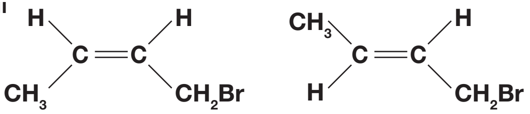
Q19 D Terylene is a polyester formed from ethane-1,2-diol and benzene-1,4-dicarboxylic acid.

Q20 B $(CH_3)_3CNH_2$ and $C_2H_5NH_2$ will be formed by reactions I and II respectively.

Paper 2 IB questions and IB style questions

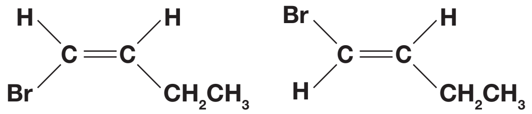
Q1 a i No rotation is possible due to the presence of a double bond or π bond. (There is restricted or hindered rotation.) [1]

ii



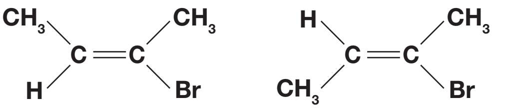
cis *trans*

or



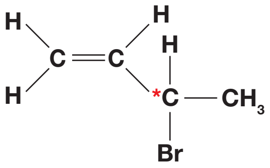
cis *trans*

or



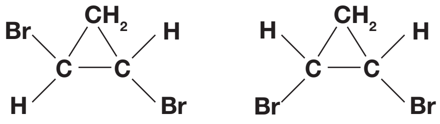
cis *trans* [2]

iii



[2]

b Restricted rotation because the carbon–carbon bond is part of a rigid cyclic system. [1]



cis *trans* [2]

Q2 a $CH_3CH(OH)CH_2CH_3$ [1]
Butan-2-ol [1]
The molecule contains a chiral/asymmetric carbon atom/ four different groups around one carbon atom. [1]
The plane of plane-polarized light is rotated in opposite directions. [1]

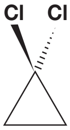
b $CH_2CHCH_2CH_3$/$CH_3CHCHCH_3$ [1]
But-1-ene/but-2-ene [1]
Dehydration/elimination [1]
Bromine or bromine water is decolorized. [1]

Q3 a The replacement of an atom or group of atoms in a molecule; [1]
by a species with a lone pair of electrons/species attracted to an electron-deficient carbon atom. [1]

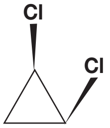
b One mark for any of the following (max [4]):
Correct structure of $(CH_3)_3CI$ [1]
Curly arrow showing C–I bond fission [1]
Correct structure of carbocation intermediate $(CH_3)_3C^+$ [1]
Curly arrow showing attack by OH^- on the electron-deficient carbon atom [1]
Correct structure of $(CH_3)_3COH$ [1]

c Correct structure of $CH_3CH_2CH_2CH_2I$ [1]
Correct structure of transition state showing charge and all bonds [1]
Curly arrow showing attack by OH^- on the correct carbon atom [1]
Correct structure of $CH_3CH_2CH_2CH_2OH$ [1]

Q4 a



and 1,1-dichlorocyclopropane [1]



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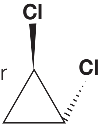

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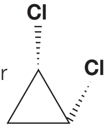

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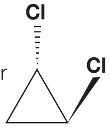
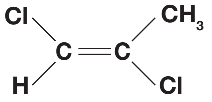


(*cis*- or *trans*-) 1,2-dichlorocyclopropane [1]
Award [1] for the correct name corresponding to the related isomer.
Accept diagrams that do not display 3-dimensional structure.
(Award [1] max for correct structures only, without the corresponding names.)

b



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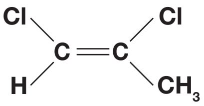

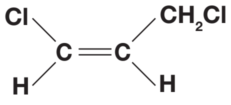

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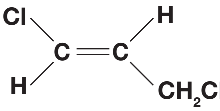
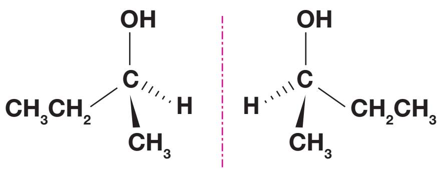


[2]

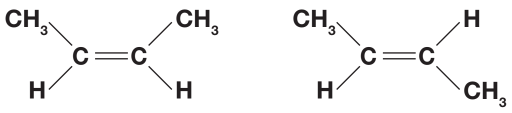
Q5 a



i Award [2] for both tetrahedral structures, or [1] if tetrahedral structure is not clear. [2]

ii Plane-polarized light [1]; rotation in opposite/different directions [1]. [2]

iii



[2]

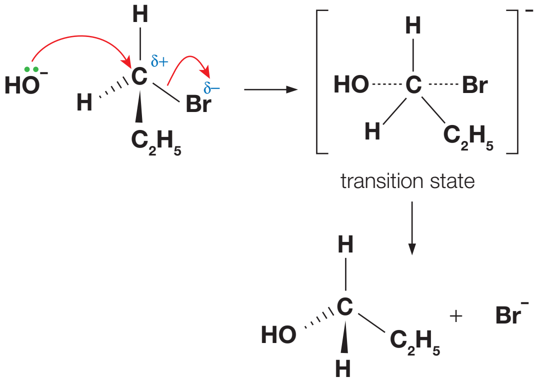
iv One mark for any of the following (max [3]):
Curly arrow showing attack by OH^- on end H [1]
Curly arrow showing C–Br bond fission [1]
Curly arrow showing formation of double bond [1]
H_2O and Br shown as products. [1]
(If but-2-ene formed, award [2] max.)

b $CH_3OH + HCOOH \rightarrow HCOOCH_3 + H_2O$
Award [1] for both reactants and [1] for both products (accept $C_2H_4O_2$). [2]
Methyl methanoate [1]

Q6 a **C** = CH_3CH_2CHO; **D** = CH_3CH_2COOH [2]

b Potassium dichromate(VI)/$K_2Cr_2O_7$ (accept other strong oxidizing agents); acidified/H^+/sulfuric acid; distil off the product as it forms/limiting the amount of oxidizing agent/reducing concentrations of acid or oxidizing agent. [3]

c Substitution; nucleophilic; bimolecular/two particles in rate-determining step. (Award [2] for three correct, [1] for two correct.) [2]



Suitable diagram with:
Curly arrow from O of OH^- to C joined to Br [1]
Curly arrow from C–Br bond to Br (this can be on the reagent or on the transition state) [1]
Transition state with negative charge and --- bonds to Br and OH group [1]
Correct products [1]

d Faster (for **A**)
C–Br bond weaker/easier to break than C–Cl bond [2]
Allow opposite argument for $CH_3CH_2CH_2Cl$ reaction being slower.

e Esterification/condensation [1]

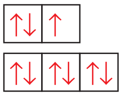
$CH_3CH_2COOCH_2CH_2CH_3$ [1]

f CH_3COOCH_3/$HCOOCH_2CH_3$ [1]; no OH group/**D** has OH group [1]; no hydrogen bonding possible with water/hydrogen bonding only possible with **D** [1]. [3]

21 Modern analytical chemistry

Q1 a Difference in metal ion (Zn^{2+} lacks a partially filled d sub-shell). [1]
Difference in oxidation number (V^{2+} has a d^3 configuration; V^{3+} has a d^2 configuration). [1]
Difference in type of ligands (this influences the splitting of the energy levels in the d sub-shell). [1]

b [2]



c $[Ni(NH_3)_4(H_2O)_2]^{2+}$ greater; $[NiCl_4]^{2-}$ less (based on the spectrochemical series). [2]

d i The colours shown by the two ions are complementary to the colours they absorb. [1]
The colour is caused by transitions between the split d orbitals. [1]
Different metals cause the d orbitals to split differently (due mainly to the different number of protons in the nucleus (nuclear charge), which determines the electron configuration. [1]

ii The oxidation state affects the size of the d orbital splitting due to the different number of electrons present and hence alters the effective nuclear charge of the nucleus of the transition metal ion. [2]
The more electron-dense (or more basic) the ligand the greater the splitting of the d orbitals. [1]

Q2 a The bond angle (H—S—H) changes as bending occurs. [1]
The lengths of the S—H bonds change as stretching occurs. [1]
The dipole moment of the molecule changes. [1]

b i Absorptions in NMR are due to transitions between different spin states in the nucleus (when an external magnetic field is applied). [1]
IR absorptions are due to bond vibrations. [1]
The nuclear transitions are at a much lower energy than bond vibrations. [1]
There is an inversely proportional relationship between frequency and energy. [1]

ii The bond in both molecules vibrates and stretches; only the stretching in the hydrogen bromide, H—Br, causes a change in dipole moment. [2]

c $\text{wavenumber} = \dfrac{1}{\text{wavelength (cm)}}$;

$\text{energy} = \dfrac{\text{Planck's constant} \times \text{velocity}}{\text{wavelength}}$

$\text{wavenumber} = \dfrac{\text{energy}}{\text{Planck's constant} \times \text{velocity}}$ [2]

d Atomic masses increase from chlorine to iodine. [1]
Heavier atoms vibrate more slowly and hence the stretching frequency decreases. [1]

e A O—H stretch (hydrogen bonded)
B C—H stretch
C C—H stretch and O—H bend
D C—H stretch

E C—N stretch
F C—H stretch
G O—H stretch (hydrogen bonded)
H C=O stretch
I C—H stretch
J C=O stretch [5]

f Top to bottom: ethanol, propanenitrile, ethanoic acid and ethyl butanoate. [3]

Q3 Determine the wavelength of the peak of maximum absorption, λ_{max}. Prepare at least five different solutions of known concentrations from the standard solution. Measure the absorbance for each concentration at λ_{max}. Plot a calibration curve and read off the value of unknown concentration from its absorbance. At a fixed wavelength the absorbance is directly proportional to the concentration provided the same path length is used. The Beer–Lambert law ($A = \varepsilon cl$) applies only to dilute solutions. [6]

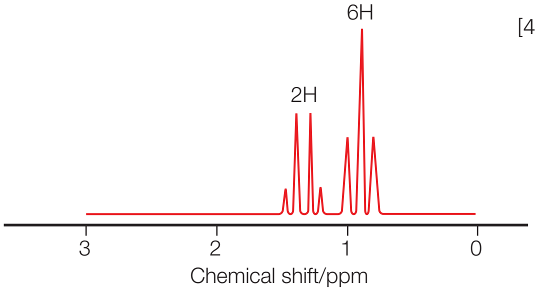
Q4 **a** Tetramethylsilane (TMS); used as reference standard. [2]

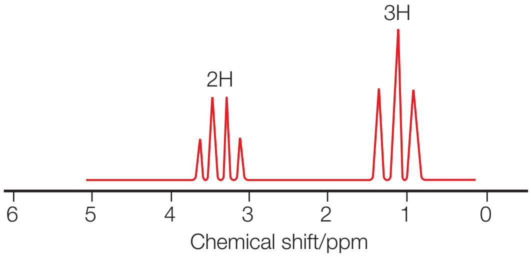
b **i** The number of different chemical environments. [1]

ii The number of hydrogen atoms in each chemical environment. [1]

iii The number of hydrogens on adjacent carbon atoms. [1]

c [4]





d The compound is ethylbenzene: $C_6H_5CH_2CH_3$. [1]
Five hydrogen atoms at a chemical shift greater than 7 ppm suggests the presence of five hydrogen atoms in a benzene ring with one substituent (attached functional group). [1]
There are two hydrogen atoms with a chemical shift corresponding to hydrogen atoms bonded to a carbon atom which is attached to a benzene ring. [1]
There are three hydrogen atoms which are part of a methyl group, $-CH_3$. [1]

Q5 **a** There is a stationary phase and a mobile phase. [1]
The components of the mixture dissolve, absorb, or partition to differing extents between the two phases. [1]

b **i** A [1]

ii B [1]

iii No, the two 'oranges' are different substances since they have different R_f values. [2]

iv R_f value = 9/22 = 0.41 [1]

c **i** R_f value of compound 1 = 1.6/12.6 = 0.13 [1]
R_f value of compound 2 = 9.1/12.6 = 0.72 [1]
Compound 1 has a greater affinity for the thin layer than compound 2.
Since the thin layer is composed of silica gel, compound 1 is more polar than compound 2. [1]

ii TLC is faster than paper chromatography. The thin layer may be made from a variety of different solids. Hence, a range of mixtures can be separated by using different stationary and mobile phases. [1]

Q6 **a** By measuring the difference in time between the injection of the sample and the centre of the peak for a component. [1]

b The areas under the peaks represent the amounts of the components in the mixture. [1]

c **i** Hexan-1-ol [1]

ii Ethanol [1]

d Two from: flow rate of the carrier gas; temperature of the column; length and diameter of the column; the nature of column surface and the volatility of the solute. [1]

e **i** The two compounds have relatively high melting points and are involatile. [2]

ii HPLC [1]

Q7 **a** **i** An electron will be promoted from a relatively low energy molecular orbital to a higher energy molecular orbital. (The molecular orbitals may be σ, π or lone pairs.) [2]

ii Complex molecules will have a greater number of available energy levels within their molecular orbitals. [1]

b **i** Two from: chlorophyll, hemoglobin, retinol and beta-carotene. [2]

ii The more extensive the π conjugation, the smaller the difference between lower and higher energy levels (energy gap) and the longer the wavelength. [2]

c Methylbenzene absorbs ultraviolet radiation because of the presence of delocalized π electrons. Ethanol has no π bonds and hence does not absorb ultraviolet radiation. [3]

Q8 **a** P: $^+CH_3$, Q: $^+C_2H_5$ and R: $^+C_3H_7$ [3]

b [3]

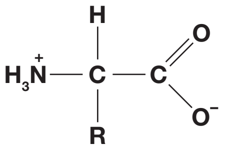
Information	Analytical technique
Relative atomic mass of an element	Mass spectrometry
Functional groups and strength of bonds present in an organic molecule	Infrared spectroscopy
Potassium ions in a sample of blood serum	Atomic absorption spectroscopy
Detailed structure of an organic molecule (without degrading the sample)	Nuclear magnetic resonance

22 Human biochemistry

Q1 **a** **i** An optically active molecule will rotate the plane of plane polarized light. [1]

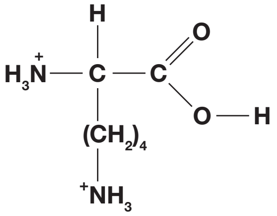
ii Glycine has two hydrogen atoms bonded to the central carbon atom. [1]
The molecule does not contain a chiral or asymmetric centre. [1]

b **i** [2]



ii The amine or amino group: $-NH_2$. [1]

iii [2]



iv Overall charge is +1. [1]

Q2 **a** Humans lack the enzyme cellulase. [1]

b **i** Any two from: diverticulosis, irritable bowel syndrome, constipation, obesity, Crohn's diease, hemorrhoids, reduction in blood pressure, risk of coronary heart disease and diabetes mellitus. [2]

ii One from: hemicellulose, dextrins, inulin, lignin, pectin and chitin. [1]

c **i** Amylose is a linear polymer of α-glucose residues linked by 1,4-glycosidic bonds. [2]

Cellulose is a linear polymer β-glucose residues linked by 1,4-glycosidic bonds. [2]

ii Amylose acts as a storage form of glucose in plants. [1]

Cellulose is the major component of cell walls in plants. [1]

Q3 **a** Components of cell membranes. [1]

b It is a digylceride since it contains two esterified fatty acids. [2]

c It contains a charged phosphate group and a charged ammonium group. [1]

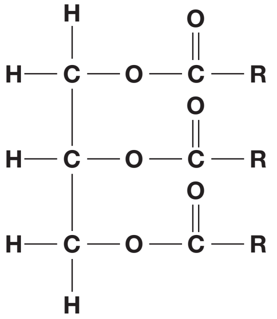
These charged groups are polar and attract water molecules. [1]

d Cholesterol. [1]

Synthesis of testosterone. [1]

Q4 **a** Release of energy during respiration; act as energy stores in the form of adipose (fat) tissue; and production of phospholipids for cell membranes. [3]

b [1]



c Triglycerides are non-polar and unable to form hydrogen bonds with water. [1]

d **i** Unsaturated; it contains a carbon–carbon double bond. [2]

ii Liquid; plants. [2]

Q5 **a** Four. [1]

b The iodine number is defined as the number of grams of iodine that adds to 100 g of a fat or an oil in an addition reaction. [1]

Mass of 1 mole of $C_{20}H_{32}O_2$ = 304 g; its four double bonds react with 1040 g of I_2. [1]

Hence 100 g of $C_{20}H_{32}O_2$ reacts with 342 g I_2. Iodine number of $C_{20}H_{32}O_2$ is 342. [1]

c Improve the blood lipid (cholesterol and triglyceride) levels and decrease the risk of coronary heart disease. [1]

d To transport cholesterol and fatty acids. [1]

High density lipoproteins and low density lipoproteins. [1]

Q6 **a** α-D-ribose – A; β-D-ribose – C. [2]

b No, because they are not mirror images: A and C are both D-ribose. [1]

c Carbon atoms 1 and 5. [1]

d RNA (ribose nucleic acid). [1]

e Heat produced = mass of water × specific heat capacity × temperature rise [1]

Heat produced = $500\,\text{g} \times 4.18\,\text{J g}^{-1}\,\text{K}^{-1} \times 67.5\,\text{K}$ = 141 075 J [1]

Calorific value of fruit bar = $\frac{40.0}{10.0} \times 141\,075\,\text{J}$ = 564.3 kJ (564 kJ). [1]

Q7 **a** **i** It is a chemical substance that interacts with a target cell. [1]

ii In the bloodstream. [1]

b Testes. [1]

c Use – treatment of wasting illness such as cancer or AIDS/to regain muscle tissue and appetite, treatment of eczema; bone marrow stimulation in patients with kidney failure; to increase libido and to induce male puberty. [1]

Abuse – increase muscle mass to enhance performance/increase strength. [1]

d Alcohol, alkene or ketone. [2]

e Any one from:

- cholesterol has an —OH group instead of the >C=O group in testosterone
- cholesterol has an alkyl/hydrocarbon side-chain instead of the —OH group intestosterone
- there is no carbonyl group present in cholesterol
- the position of the >C=C< bond is different in cholesterol compared to testosterone. [1]

f Testosterone and epitestosterone differ in the configuration of the hydroxyl group, —OH. In one isomer it is pointing down; in the other isomer it is pointing up. [1]

g Insulin; pancreas; control of blood glucose level. [3]

h Any three from:

- change release of hormones/FHS/LH (from hypothalamus/pituitary gland)
- prevent ovulation: release of ovum (egg)
- prevent attachment of ovum (egg) to uterus
- prevent the sperm from reaching ovum. [3]

Q8 **a** **i** Activity found by collecting and measuring the total volume of carbon dioxide gas per unit time. [1]

Temperature and pH (via use of a buffer) must be controlled during the reaction. [1]

ii Specific – a particular enzyme can catalyse only one reaction. [1]

Enzyme binds to/reacts with substrate to form an enzyme substrate complex/E + S → ES. [1]

After the reaction the product leaves the enzyme /EP → E + P; [1]

Mention of an active site - a cleft where the substrate binds. [1]

b **i** The enzyme would be irreversibly denatured and its biological activity would be lost. [1]
Loss of three-dimensional structure (secondary and tertiary structures) due to disruption of hydrogen bonds and other intermolecular forces. [1]

ii Mercury ions would irreversibly inhibit the reaction and the biological activity would be lost. [1]
The mercury ions would form covalent bonds with the —SH groups of cysteine residues. This would disrupt the catalysis and alter the shape of the enzyme. [1]
Mercury ions would react with charged carboxylate groups, —COO⁻, in variable side chains (*R* groups) disrupting ionic interaction. [1]

c Iron – oxygen transport in hemoglobin; anemia [2]
Iodine – thyroxine synthesis; goitre [2]

d **i** In competitive inhibition, the inhibitor binds to the same active site as the normal enzyme substrate, without undergoing a chemical reaction. [1]
The substrate molecule cannot enter the active site while the inhibitor is there, and the inhibitor cannot enter the site when the substrate is there. [1]
Correct line on graph – to right of original – approaches V_{max} at a higher substrate concentration. [1]

ii V_{max} is unchanged; K_m increased. [2]

Q9 **a** **i** Hydrogen bonding [1]

ii Complementary base pairing: adenine only hydrogen bonds to thymine; adenine only hydrogen bonds to cytosine. [1]
Complementary base pairing (and appropriate enzymes) allows an accurate copy of a DNA molecule to be synthesized. [1]

b **i** The ends of the DNA are chemically different: the 5′ end has a phosphate group and the 3′ end has a hydroxyl group. [1]
The backbones or strands of DNA have a polarity or direction. [1]

ii 3′-CTTAAGGTACTGTATGGTAACC…-5′ [2]

iii ATG = methionine; ACA = threonine; TAC = tyrosine and CAT = histidine. [2]

c Deoxyribose sugar and phosphate group. [2]
The nucleotides condense (under enzyme control) and form a phosphodiester bond; [1]
between carbon-3 of the deoxyribose sugar and a neighbouring phosphate group. [1]

d **i** DNA profiling is the name given to the technique that can identify an individual with a high degree of certainty from a sample of the individual's genetic material. [1]
The lengths of each repeated sequence of bases (minisatellites) can be made visible by the use of DNA probes. [1]
The probe is labelled with radioactive phosphorus-32, mixed with the individuals' DNA and exposed on photographic film. [1]
Uses – one from: forensic uses (e.g. identifying rapists from blood or semen), establishing genealogy (family trees) or paternity suits. [1]

ii Virtually zero [1]

iii The blood stain is likely to be that of the male householder and suspect A. [2]

Q10 **a** $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$ [2]

b Pyruvic acid or pyruvate [1]

c Lactate or lactic acid; ethanol [2]

d Carbon dioxide [1]

e Aerobic respiration [1]

f Cytochrome; copper [2]

Q11 **a** It is the maximum rate or velocity, V_{max} (active site saturation). [1]

b 0 to Y: free active sites can accommodate an increase in the substrate concentration since there are free active sites. [1]
At higher concentrations, the enzyme molecules are saturated and there are no free active sites. [1]

c K_m represents the substrate concentrations at which the reaction rate is half of the maximal rate, V_{max}. [1]
The higher the value of K_m, the lower the affinity of the affinity for substrate. [1]

d 30 μmol dm⁻³ [1]

e $H_2O + C_{12}H_{22}O_{11} \rightarrow 2C_6H_{12}O_6$ [2]
A large number of glucose molecules are involved in many enzyme controlled dehydration reactions leading to the formation of glycosidic bonds, —O—. [1]

Q12 **a** The retention factor (R_f) = distance moved by solute/distance moved by solvent [1]

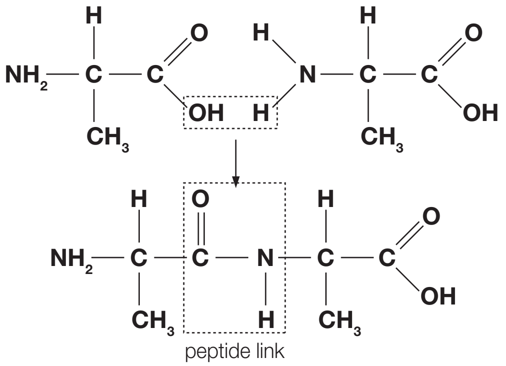
b 0.8; 0.3 and 0.9. [1]

c 0.2, 0.4 and 0.6. [1]

d Dimension 1: methionine and proline [1]
Dimension 2: glutamic acid, threonine, histidine [1]

e Electrophoresis (gel or paper) [1]

Q13 **a** [2]

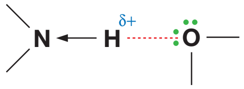


b Structural (e.g. collagen); enzymes (e.g. pepsin); substrate for respiration; oxygen transport (e.g. hemoglobin) or oxygen storage (e.g. myoglobin). [2]

c The mixture of amino acids is spotted on the electrophoretic paper or gel. [1]
A high voltage is applied and maintained across the ends of the paper or gel. [1]
The chromatogram is developed with ninhydrin. [1]
The distances moved by the stained amino acids are compared with known samples of pure amino acids run under the same conditions. [1]

d Six [1]

e The tertiary structure of a protein describes the way the chain of amino acids is folded into a precise compact globular structure. [1]
Lengths of α-helix (secondary structure) are brought into contact with each other by connecting lengths of the amino acid chain by taking up a more random arrangement. [1]
The tertiary structure is maintained by hydrogen bonds between hydrogen atoms attached to oxygen and nitrogen atoms.



When hydrogen bonds form between groups in the peptide chain, which are far apart in the primary structure, the chain must fold in such a way as to bring these groups close together. [1]
Ionic bonds (salt bridges) may also form between $—NH_3^+$ and $—COO^-$ groups. [1]
Disulfide bridges (—S—S—) may be formed when two —SH groups are close enough to be oxidized. [1]
(The folding process is also driven by the hydrophobic effect which involves the non-polar side chains of amino acid residues located in the interior of the protein.)

Q14 **a** Indicates the presence of one or more carbon–carbon double bond (>C=C<). [1]
b $C_{25}H_{51}COOH > C_{23}H_{47}COOH > C_{25}H_{48}COOH$ [1]
c van der Waals' forces [1]
d $C_{25}H_{51}COOH$ and $C_{25}H_{47}COOH$: difference in chain length, area of contact, surface area or number of electrons. [1]
$C_{25}H_{51}COOH$ and $C_{25}H_{48}COOH$: difference in bond angle, closeness of packing or area of contact. [1]

23 Chemistry in industry and technology

Q1 **a** Aluminium is too reactive for its oxide to be reduced by carbon; it is above carbon in the reactivity series. [1]
b Alumina is composed of two ions with a high charge density (Al^{3+} and O^{2-}); [1]
so the electrostatic attraction between the ions is large and the melting point high. [1]
c Cryolite is added to reduce the melting point of aluminium oxide; [1]
less energy is required and hence the cost of the process is decreased. [1]
d Cathode: $Al^{3+} + 3e^- \longrightarrow Al$ [1]
Anode: $2O^{2-} \longrightarrow O_2 + 4e^-$ [1]
e Aluminium metal becomes coated with a thin layer of aluminium oxide. [1]
This layer adheres to the surface of the metal protecting it from further reaction. [1]
Iron reacts with the air to produce an oxide, but this layer does not adhere to the metal surface so corrosion continues. [1]
f Less energy is needed to recycle aluminium than to extract it. [1]
Recycling conserves bauxite deposits and reduces the need to tip waste in landfill sites. [1]

Q2 **a** Percentage by mass of iron in Fe_2O_3

$$= 55.85 \times 2 \times \frac{100}{159.7} = 69.9\% \qquad [1]$$

Percentage by mass of iron in Fe_3O_4

$$= 3 \times 55.85 \times \frac{100}{231.55} = 72.4\% \qquad [1]$$

b Iron ore/haematite/magnetite [1], coke (carbon) [1] and limestone (calcium carbonate) [1] [3]
c Molten iron [1] and slag (calcium silicate) [1] (accept carbon dioxide, CO_2) [2]
d **i** $C(s) + CO_2(g) \longrightarrow 2CO(g)$ [1]
ii Reducing agent [1]
iii $Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$ [1]
e At high temperatures the calcium carbonate undergoes thermal decomposition:
$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$ [1]
The calcium oxide then reacts with sand (silicon dioxide) to form slag: [1]
$CaO(s) + SiO_2(s) \longrightarrow CaSiO_3(s)$ [1]
f Amount of iron(III) oxide

$$= \frac{80\,000\,000\,\text{g}}{159.7\,\text{g mol}^{-1}} = 5\,009\,398\,\text{mol} \qquad [1]$$

Mass of iron $= 500\,939\,\text{mol} \times 2 \times 55.85\,\text{g mol}^{-1}$
$= 55\,954\,915\,\text{g}$
$= 56$ tonnes [1]
g **i** Cast iron or pig iron [1]
ii The presence of carbon (2–4%) makes it brittle. [1]
h Oxygen is blown into molten pig iron at high pressure through molten iron. [1]
The oxidized impurities, sulfur and phosphorus, are removed as their gaseous oxides. [1]
i Steel can be made less brittle by tempering. [1]
This involves heating it to about 400–600 °C and letting it cool slowly. [1]
To make the steel more ductile a process known as annealing is used; this involves heating the steel to a higher temperature (about 1040 °C) followed by slow cooling. [1]
If hard steel is required then after annealing the steel must be quenched by rapid cooling. [1]

Q3 **a** Any two from:
- The products of electrolysis of brine are more useful/economically important.
- As well as chlorine, brine produces hydrogen and sodium hydroxide.
- Electrolysis of molten salt produces only sodium. [2]

b At anode: $2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$ [1]
At cathode: $2H^+(aq) + 2e^- \longrightarrow H_2(g)$ [1]
c The Castner–Kellner cell uses a flowing mercury cathode; [1]
mercury is toxic and dangerous if it leaks into the environment; [1]
where it enters the food chain in shellfish. [1]

Q4 **a** There is no lithium metal in lithium-ion cells. [1]
The anode is made of graphite and the cathode of a layered metal oxide (usually $LiCoO_2$). [1]
b Lithium ions (Li^+) flow through the electrolyte and become inserted (intercalated) between the layers of the two electrode materials when they reach them. [1]
The conversion between Li^+ ions and Li atoms takes place between the layers of the electrode material. [1]
c Similarity – one from:
- Both convert chemical energy to electrical energy.
- Both can function repeatedly (the fuel cell needs to be supplied continuously, a battery needs to be recharged). [1]

Difference – one from:
- Fuel cells have no 'memory effect' when they are refuelled and are low maintenance.
- The efficiency of fuel cells is approximately 80%, which is much higher than for conventional batteries. [1]

Q5 **a** **i** Cracking refers to the decomposition of large hydrocarbons [1] into smaller hydrocarbons [1]. [2]
ii The smaller hydrocarbons make better fuels. [1]
Alkenes, used in the manufacture of plastics, are also produced. [1]
b **i** … + C_8H_{18} [1]
ii Ethene [1]
propene [1]
and octane. [1]

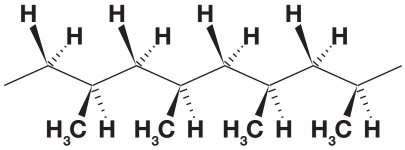
c **i** Modern cracking uses zeolites as the catalyst. These are complex aluminosilicates. Aluminium oxide and silicon dioxide can also be used as a catalyst. The alkane is brought into contact with the catalyst at a temperature of 500 °C and moderately low pressures. The zeolites used in catalytic cracking are chosen to give high percentages of hydrocarbons with between five and ten carbon atoms – particularly useful for petrol. It also produces high proportions of branched alkanes and aromatic hydrocarbons such as benzene. Any [3]

In thermal cracking, high temperatures (typically in the range of 450 to 750 °C) and pressures (up to about 70 atmospheres) are used to break the large hydrocarbons into smaller ones. Thermal cracking produces mixtures of products containing high proportions of alkenes. Any [2]

ii The feedstock is mixed with hydrogen at a pressure of about 80 atmospheres and cracked over a platinum or silica/alumina catalyst. This gives a high yield of branched alkanes, cyclic alkanes and aromatic compounds for use in unleaded gasoline (petrol). Major products from hydrocracking include kerosene and diesel. Any [2]

The feedstock, for example, napthha or ethane, is preheated, vaporized and mixed with steam and then converted at 850 °C to give high yields of low molecular mass alkenes. This is a very rapid process and a variety of products are produced. Ethane produces a range of light alkenes; naphtha produces products rich in aromatics. Any [2]

Q6 **a** [2]



b In the atactic form the methyl side chains are randomly orientated. [1]
In the isotactic form the methyl groups are arranged are on side of the polymer chain. [1]

c The isotactic form has a regular structure; [1]
and hence is strong and crystalline. [1]
The atactic form has an irregular structure; [1]
and hence is soft with little strength. [1]

d Carpet fibres or polymer banknotes; [1]
which need to have high tensile strength. [1]

e Plasticizers are small molecules that can fit between the long polymer chains. [1]
They weaken the attractive intermolecular forces operating between the chains, making the plastic more flexible. [1]
By varying the amount of plasticizer added polypropene can form a range of polymers with physical properties ranging from rigid to fully pliable. [1]

f Pentane is added during the formation of polystyrene and the product is heated in steam; the pentane vaporizes producing expanded polystyrene. [1]
This light material is a good thermal insulator; [1]
and is also used as packaging as it has good impact-absorbing properties. [1]

g Advantage – one from:
- light
- easy to mould
- resistant to corrosion [1]

Disadvantage – one from:
- depletion of natural, non-renewable resources
- the majority of polymers are derived from crude oil
- most plastics are non-biodegradable and release toxic fumes when incinerated [1]

Q7 **a** **i** Cathode: $Pb + SO_4^{2-} \longrightarrow PbSO_4 + 2e^-$ [1]
Anode: $PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \longrightarrow PbSO_4 + 2H_2O$ [1]

ii **Reduction** at anode or + electrode [1]
Oxidation at cathode or − electrode [1]

b $Pb + PbO_2 + 4H^+ + 2SO_4^{2-} \longrightarrow 2PbSO_4 + 2H_2O$ [1]

c The reactions are reversible. [1]

d Lead(II) ions react with sulfate ions to form lead(II) sulfate. [1]

e Electrolysis occurs and the hydrogen ions undergo discharge at the cathode; [1]
to form hydrogen. [1]

f One from: the acid is corrosive/lead is potentially polluting and toxic. [1]

Q8 **a** Nanotechnology is the ‘technology of the very small’ – 1 nanometre = 10^{-9} metres. [1]
Nanotechnology is concerned with building devices and manipulating materials at this size. [1]
Materials show novel and exceptional properties at this scale which nanotechnology sets out to exploit. [1]

b **i** Ratio = $5 \times 10^{-6}\,\text{m}/10^{-9}\,\text{m}$ [1]
= 5000 [1]

ii Nanotubes are strongly bonded and the rigid tubes with an exceptional surface area to size ratio – very long and narrow. [1]
One use from: as heterogeneous catalysts, as electrical wires (if coated internally with, say, silver), as transport mechanism for medical drugs. [1]

c The nature of nanotubes means that if there is a diverse population in terms of length, for instance, then the batch could contain a very wide range of properties. [1]
It is therefore difficult to standardize behaviour in an experimental context and control the results. [1]

Q9 **a** The formation of low-density poly(ethene) involves a free-radical mechanism. [1]
Plus two from:
- Initiation brought about using a catalyst such as benzoyl peroxide (ROOR)
  $ROOR \longrightarrow RO\bullet + \bullet OR$
  homolytic fission of the weak O–O bond
- Propagation: $RO\bullet + H_2C{=}CH_2 \longrightarrow RO{-}CH_2{-}CH_2\bullet$
  $RO{-}CH_2{-}CH_2\bullet + H_2C{=}CH_2$
  $\longrightarrow RO{-}CH_2{-}\,CH_2{-}CH_2{-}CH_2\bullet$
- Termination: $2RO{-}CH_2{-}CH_2\bullet$
  $\longrightarrow RO{-}CH_2{-}\,CH_2{-}CH_2{-}CH_2{-}OR$ [2]

b Catalyst is a Ziegler–Natta catalyst or titanium(IV) chloride, $TiCl_4$, with diethylaluminium chloride, $Al(C_2H_5)_2Cl$. [1]
Possibility of dative covalent bond formation between the pi (π) electrons of ethene and vacant orbitals on the Ti atom. [1]

Q10 a Any three from:

- Graphite has a giant molecular lattice structure whereas carbon-60 has a simple molecular structure with just relatively weak van der Waals' forces between the molecules.
- Graphite has a very high m.p., whereas carbon-60 has low m.p.
- Graphite is very strong in one orientation because of extensive covalent lattice, carbon-60 is weak.
- Graphite has delocalized electrons between the layers and conducts electricity; carbon-60 has no free electrons so does not conduct. [3]

b The walls of a nanotube represent a single sheet of graphite (graphene) rolled on itself; [1]
made up of hexagons. [1]
To close the tube need some pentagons in the arrangement to achieve the curved structure (as in carbon-60). [1]

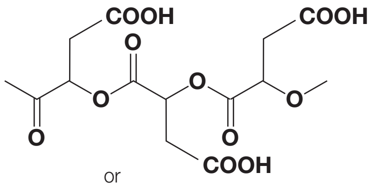
c Titanium dioxide particles reflect the harmful ultraviolet radiation away from the skin. [1]
They act as a 'sunblock'. [1]
The small nanoscale particles do not give a whitening effect because of their small size. [1]

Q11 a **i** Sphere with a diameter of the order of nanometres/in the nanometre range. [1]
ii It is (highly) acidic/has very low pH/contains HCl. [1]
iii Use hydrogels: [1]

- of different (wall) thickness/strength (to release drug over time) [1]
- of different chemical composition (for different breakdown times) [1]
- incorporating pores/holes (in their walls) [1]

Any [2]

b For the homopolymer, either using the amino acid the minimum is:
–CO–CHR–NH–CO–CHR–NH–CO–CHR–NH– [1]
or using the hydroxyacid the minimum is:



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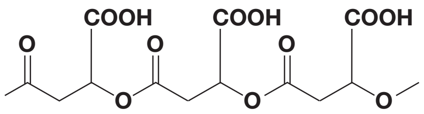
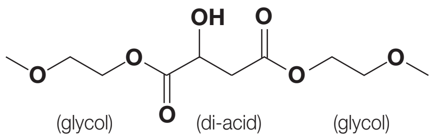


[1]

For the heteropolymer, either using the glycol compound and the di-acid the minimum is:



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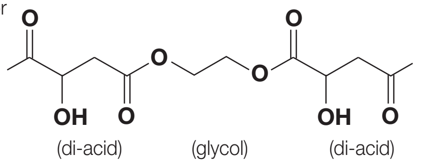
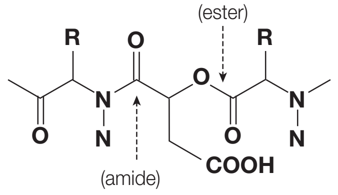


[1]

or using the amino acid and the di-acid, the minimum is:



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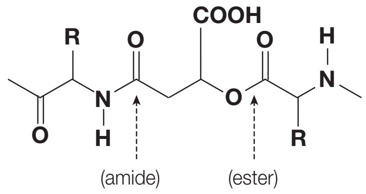

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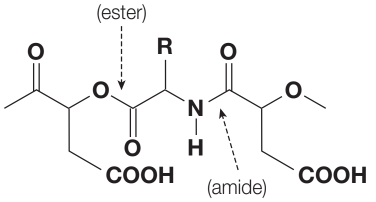

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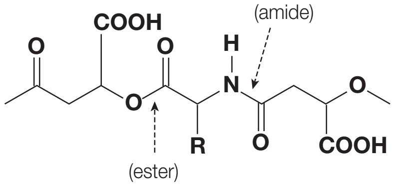


[1]

A heteropolymer incorporating all three monomers can also be drawn. [1]
Max. [3]

Q12 a **i** Sc: $1s^22s^22p^63s^23p^63d^14s^2$ [1]
Sc^{3+}: $1s^22s^22p^63s^23p^6$ [1]
ii $[(Sc^{3+})_3C_{82}^{9-}]$ [1]

b **i** sp^2 [1]
sp^3 [1]
and sp^2 [1]
ii Graphite: giant covalent/network solid/two-dimensional giant covalent. [1]
Fullerenes: simple molecular/molecular covalent. [1]
iii Graphite does not contain pentagonal rings; [1]
the sheets of carbon atoms are flat. [1]

c **i** 984 [1]
ii C_{82}^{+} [1]
iii Helium is an excellent coolant; [1]
and provides an inert and non-oxidizing environment for fullerene formation. [1]

d **i** The fullerenes are simple molecular substances. [1]
They are associated in the solid state due to the relatively weak van der Waals' forces. [1]
ii C_{82} has a larger number of atoms and hence a greater number of polarizable electrons available for van der Waals' forces. [1]
More energy is required to overcome the intermolecular forces before sublimation occurs. [1]

Q13 a Carried out at very high pressures (1000–3000 atmospheres); [1]
at a temperature of about 500 K; [1]
an initiator such as an organic peroxide or a trace of molecular oxygen is added. [1]

b **i** A regular polymer, the molecules of which can be described in terms of only one species of stereo-repeating unit; [1]
in a single sequential arrangement. [1]
ii Titanium(IV) chloride [1]
$TiCl_4$ [1]
with diethylaluminium chloride [1]
$Al(C_2H_5)_2Cl$ [1]
iii It accepts a dative covalent/coordinate bond from the monomer's π bond. [1]

c LDPE: free-radical polymerization involves long-chain primary (terminal) radicals that can abstract hydrogen atoms within the chain; [1]
subsequent polymerization from the new internal radical site generates randomly branched alkanes that do not pack together efficiently and give low density. [1]
HDPE: cationic polymerization with Ziegler–Natta catalysis yields linear polyethene chains that pack efficiently to give crystalline regions and high density. [1]

Q14 **a** **i** An alloy is a homogeneous mixture of metals or a mixture of a metal and non-metal. [1]
ii Any two from: alloys are harder/stronger/decreased malleability/increased corrosion resistance/increased electrical or thermal conductivity. [2]
b **i** The carbon gains oxygen/its oxidation number increases from 0 to +2. [1]
ii Brass [1]
One use from: locks, gears, bearings, doorknobs, ammunition, valves for plumbing; electrical applications; musical instrument, such as horns and bells; zippers; to make coins. [1]
c **i** Carbon [1]
ii One use from: cutlery, surgical instruments, watches, sinks. [1]

Q15 **a** A substance which, when present in relatively small amounts, increases the rate of a chemical reaction; [1]
but which is not consumed during the overall process. [1]
The function of a catalyst is to provide a new reaction pathway with a lower activation energy. [1]
b **i** +1 [1]
ii Homogeneous catalysis occurs when the catalyst and the reactants are in the same phase (or physical states). [1]
Heterogeneous catalysis occurs when the catalyst and the reactants are in different phases (or physical states). [1]
c
- Advantages of heterogeneous catalysts: the products are easily separated by physical means from the catalyst; or they tend to be less specific than homogeneous catalysts so can be used to catalyse a variety of reactions. [1]
- However, because it is only the surface area, or parts of the surface area, that is active they depend very much on particle size and the smaller particles may aggregate together during the reaction, so making the catalyst less effective. [1]
- Homogeneous catalysts tend to be more specific for particular reactions and are often harder to separate from the products. [1]
- However homogeneous catalysts can be more efficient as they will be in molecular, atomic or ionic form. [1]

d They form two or more stable oxidation states and can bind the reactants by dative covalent bonding. [2]
OR
They can form dative covalent bonds with ligands using vacant orbitals. [2]
e Hydrogen iodide is corrosive; [1]
and undergoes decomposition when heated. [1]
Rhodium is expensive; [1]
and hence the cost of the process is raised. [1]
f Any four from:
- selectivity – produce only the desired product and not catalyse side reactions
- efficiency
- ability to work under mild/severe conditions
- environmental impact – e.g. mining for metals
- problems caused by catalysts becoming poisoned by impurities which bind to the surface of heterogeneous catalysts [4]

24 Drugs and medicines

Q1 **a** $Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$;
$Al(OH)_3 + 3HCl \rightarrow AlCl_3 + 3H_2O$ [1]
b Aluminium hydroxide, $Al(OH)_3$ [1]
c Corrosive to body/tissue/strong base/alkali [1]
d Heartburn occurs when gastric juice from the stomach rises into the esophagus. [1]
e Dimethicone is added as an anti-foaming agent to prevent flatulence (passing of wind) and hence reduce pain. [1]
f Production of carbon dioxide gas, CO_2. [1]

Q2 **a** An antibiotic is a synthetic or semi-synthetic substance that kills bacteria. [1]
b Prevents deactivation by stomach acid/more resistant to stomach acid; prevents deactivation by the enzyme penicillinase (produced by bacteria) increases resistance/tolerance to penicillinase. [2]
c Broad spectrum – effective against many types/strains of bacteria – and narrow spectrum – effective only for certain type of bacteria; [1]
d Penicillin interferes with the cell wall formation. [1]
Cells can expand/burst/disintegrate/bacteria die. [1]
e Makes penicillin less effective; destroys useful/beneficial bacteria; allows resistant population to build up. [1]
f Failure to complete a course of antibiotics favours the selection of antibiotic-resistant bacteria. [1]
g The strain within the four-membered ring structure increases the reactivity of the amide; [1]
the ring structure opens so that the penicillin becomes covalently bonded to the enzyme; [1]
that synthesizes the bacterial cell walls (blocking its action). [1]

Q3 **a** Intramuscular: into the muscles; intravenous: into veins; subcutaneous: into fat; anal: via the anus and inhalation. [1]
Intravenous has most rapid effect since the drug is circulated and transported quickly via the bloodstream (to various parts of the body). [1]
b A placebo. [1]
c A side-effect is an effect produced as well as the one intended and is an unwanted or undesired effect. [1]
d The synergistic effect occurs when the effects of two drugs taken together is greater than the individual effects of each drug. [1]

e The therapeutic window is an index for an estimation of Ganciclovir dosage that can treat disease effectively while remaining within the safe dosage range. [1]
The use of this index can help to avoid the potentially harmful effects (including side effects) of large doses of Ganciclovir or ineffective small doses. [1]

f Viruses undergo rapid mutation; viruses are not affected by antibiotics; it is difficult to destroy viruses without damaging cells; there is only a limited number of metabolic functions that can be targeted by antivirals. [1]

Q4 a A chiral auxiliary is an optically active species; [1]
that can be connected to a molecule (to make it optically active). [1]
When this species is reacted further (then removed) only one enantiomer results; [1]
which eliminates the need to carry out the separation of the desired enantiomer from a racemic mixture. [1]

b Both contain:
a six-membered ring;
a five-membered ring;
a (tertiary) amine group;
the nitrogen atom has a methyl group attached. Any [2]

c Short-term effects:
increased heart rate/blood pressure/restriction of blood vessels;
acts as an anti-diuretic/reduction in urine output.
Long-term effects:
increased risk of heart disease/coronary thrombosis;
risk of becoming addicted/physically dependent;
high cost;
(increased risk of) (lung, mouth, throat) cancer;
(increased risk of) bronchitis/emphysema;
reduction in capacity of blood to carry oxygen;
withdrawal symptoms/weight gain (on quitting). Any [6]

d i One from:
- interference with nerve impulse transmissions
- relief of anxiety/nervous tension/increase in feeling of calm
- reduction of mental activity
- reduction of inhibition. [1]

ii One from:
- sleep
- loss of consciousness/anaesthesia
- coma
- death
- decreases heart rate/breathing rate. [1]

e They relieve the symptoms of depression. [1]

f The need for larger doses to achieve the desired effect; increased risk of dependence/increased risk of negative effect/increases toxic effect. [2]

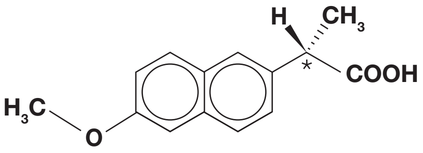
Q5 a Analgesic is a substance that reduces pain. [1]
Mild analgesics intercept pain at the source/interfere with production of substances that cause pain. [1]
Strong analgesics interact with receptor sites in the brain/prevent the transmission of pain impulses [1]
Antipyretic reduces fever. [1]

b Amide and ether functional groups. [2]

c Any one of the following beneficial effects*:* [1]
- used to treat mini-strokes
- prevents heart attacks/reduces risk of heart attack/thins the blood/anticoagulant
- relieves symptoms of rheumatological diseases/anti-inflammatory
- reduces fever.

Any one of the following side effects: [1]
- stomach bleeding
- allergic reaction
- Reye's syndrome
- hearing loss
- tinnitus (ringing in the ears)
- gastrointestinal irritation (e.g. heartburn, nausea).

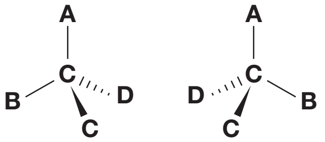
d i [1]



ii An equimolar/50:50 mixture of two enantiomers of the same molecule. [1]

iii Different enantiomers may have different biological effects, some of which may be harmful; [1]
genetic defects/deformities/other suitable example. [1]

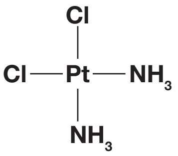
Q6 a [2]



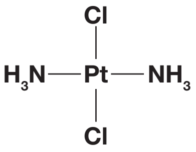
b Stereoisomerism: optical isomerism. [1]

c i Existence of compounds with the same molecular formula but with atoms/groups or Cl/NH$_3$ arranged differently in space [1]

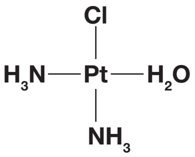
ii



cisplatin



transplatin



active form of cisplatin

[3]

iii Covalent bonding/shared pairs of electrons; dative/coordinate/hydrogen/polar bonding [1]

Q7 a Oxidizing agent/accepts electrons; [1]
orange to green. [1]

b Gas–liquid chromatography [1]
Infrared spectroscopy. [1]

c Stomach bleeding [1]

d Both substances: cause changes in visual and sound perception/hallucinations. [1]
LSD may cause a permanent effect of 'flashbacks'/effects of LSD may be experienced a year or more after the last use of the drug/psychological dependence [1]
OR Mescaline may cause nausea/trembling/liver damage/reduce appetite. [1]

e [2]

Mescaline	**LSD**
primary amine	secondary and tertiary amine
ether groups	amide
one ring	alkene(s)
	indole ring
	four cyclic rings
	heterocyclic

f A compound library is a large collection of related compounds synthesized individually and evaluated for biological properties. [1]

Computers are used to design three-dimensional models of drugs and simulate their interaction with the appropriate enzyme or protein. [1]

g Combinatorial chemistry involves the synthesis of libraries of compounds and the screening of those libraries for compounds with desirable properties. [1]
Combinatorial chemistry can rapidly generate large libraries of related compounds. [1]

25 Environmental chemistry

Q1 a Troposphere, stratosphere, mesosphere, thermosphere [2]

b

Gas	**Percentage abundance**
Nitrogen	78.09
Oxygen	20.94
Noble gases (mostly argon)	0.93
Carbon dioxide	0.035
Water vapour	0– 4

[5]

c Allow two marks for part **i** and **ii** for each gas.

Gas	**i Introduced to troposphere**	**ii Removed from troposphere**
O_3	Mechanism involving NO_2 $\bullet NO_2$ (+ light 430 nm) $\rightarrow NO\bullet + O\bullet$ $O\bullet + O_2 + M \rightarrow O_3$	Photolysis by UV radiation: O_3 + UV radiation $\rightarrow O_2 + O\bullet$ or, mechanism involving nitrogen monoxide: $NO\bullet + O_3 \rightarrow \bullet NO_2 + O_2$
CO	Decomposition of chlorophyll in green plants Oxidation of methane Forest fires Volcanic activity Incomplete combustion of fossil fuels	Reaction of carbon monoxide with hydroxyl radicals from the photolysis of ozone O_3 + UV radiation $\rightarrow O_2 + O\bullet$ $O\bullet + H_2O \rightarrow 2OH\bullet$ $CO + HO\bullet \rightarrow CO_2 + H\bullet$
NO	Combination of nitrogen and oxygen in vehicle engines and coal-fired power stations: $N_2 + O_2 \rightarrow 2NO$ Also microorganisms in soil, volcanic activity and lightning release NO	Further oxidation of nitrogen monoxide. Complex free-radical mechanism summarized as: $2NO\bullet + O_2 \rightarrow 2\bullet NO_2$
SO_2	Volcanic activity Decay of biological material Smelting of sulfide ores in the manufacture of zinc, lead and copper Combustion of fossil fuels, especially coal and heavy fuel oil: $S + O_2 \rightarrow SO_2$	Dissolves in rainwater $SO_2 + H_2O \rightarrow H_2SO_3$ Also, oxidizes further to sulfur trioxide. Complex mechanism summarized as: $2SO_2 + O_2 \rightarrow 2SO_3$

[16]

d The approximate amount of time it would take for the level of a pollutant to return to its natural level assuming emissions were stopped. This lifetime depends on the relative rates of the processes introducing the pollutant to the atmosphere, and those removing it. [2]

Q2 a Any three from:

- metal particles: manufacture of metal products
- metal oxide particles: oxidation of small amounts of metals contained as impurities in coal
- fly ash: during combustion of fossil fuels
- asbestos dust: mining and manufacture of asbestos
- aerosol mist: combination of sulfur dioxide and nitrogen dioxide with moisture [3]

b Larger particles are captured by the cilia and swallowed. Carcinogenic compounds are introduced to the digestive tract and may be carried to the organs. Smaller particles enter the lungs where they contribute to respiratory diseases. [2]

c Atmospheric particles offer a catalytic surface on which other atmospheric pollutants can be adsorbed, leading to increased rates of potentially harmful reactions. For example, sulfur dioxide may be converted to sulfur trioxide on the surface of soot or fine metallic particles. [2]

d Any four from:

- flue gases are passed through a wire mesh carrying an electric charge
- particles carried by gases become charged
- gases then pass between charged metal plates
- particles stick to the plates
- plates are shaken at intervals to remove accumulated particles. [4]

Q3 a Any three from:

- affects the pH of lakes and rivers, which impacts the organisms living in them
- affects the availability of metal ions in soil, which goes on to affect nearby plant life or surface water
- directly affects plants, damaging chlorophyll and turning leaves brown
- affects buildings and other materials, reacting with calcium carbonate
- acidic gases directly affect human health, by attacking the respiratory system, leading to asthma or bronchitis.

[3]

b $S(s) + O_2(g) \rightarrow SO(g)$ [1]
$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$ [1]
$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$ [1]

c A mesh of platinum or rhodium acts as a heterogeneous catalyst, adsorbing pollutants onto the surface and allowing oxidation to take place. The three main processes are: reduction of nitrogen monoxide to nitrogen; oxidation of unburned hydrocarbons to carbon dioxide and water vapour; oxidation of carbon monoxide to carbon dioxide. [3]

Equations:
$2NO + 2CO \rightarrow N_2 + 2CO_2$
$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$ (using octane as an example)
$2CO + O_2 \rightarrow 2CO_2$ [3]

Q4 a Visible light, (near) infrared, (near) ultraviolet [3]

b Mostly visible radiation reaches the Earth's surface. It is absorbed and re-radiated as infrared radiation, which warms the atmosphere from below. [2]

c Carbon dioxide absorbs and emits infrared radiation. [1]
Carbon dioxide absorbs infrared radiation radiated from the Earth's surface, lowering the percentage of radiation that escapes into space. [1]
Also, carbon dioxide re-radiates infrared downwards towards the surface. [1]

d GWP describes the ability of the gas to absorb infrared radiation and its lifetime in the atmosphere. [2]
It is measured relative to carbon dioxide, which is assigned a GWP of 1. [1]

GWP must specify the timescale over which it is measured. [1]

e Methane has a much higher GWP than CO_2 (72, measured over 20 years). [1]
Therefore, despite its lower concentration, it makes a larger contribution to global warming. [1]

f Any four from: rising sea-levels; glacier retreat; changing weather patterns; drought; changing crop yields; changing soil characteristics; changing availability of land for agriculture; changing patterns of disease and pests. [4]

Q5 **a** **i** Ozone formation:
Photolysis of dioxygen: O_2 (+ UV 242 nm) $\rightarrow O\bullet + O\bullet$
Combination of dioxygen with an oxygen radical:
$O_2 + O\bullet + M \rightarrow O_3 + M$ (third body *M* required) [2]

ii Ozone depletion:
Photolysis of ozone: O_3 (+ UV 290–320 nm) $\rightarrow O_2 + O\bullet$
Combination of an ozone molecule with an oxygen radical:
$O_3 + O\bullet \rightarrow 2O_2$ [2]

b Carbon–halogen bonds in chlorofluorocarbons are very strong. [1]
The frequencies of ultraviolet radiation that can break these bonds are absorbed higher in the stratosphere, so are unable to dissociate CFCs in the troposphere. [1]

c In the stratosphere, ultraviolet radiation of the correct frequencies is available to break the C–Cl bond in CFCs.
$CFCl_3 + UV \rightarrow \bullet CFCl_2 + Cl\bullet$ (photodissociation of CFC-11)
$CF_2Cl_2 + UV \rightarrow \bullet CF_2Cl + Cl\bullet$ (photodissociation of CFC-12) [1]

d The C–H bond in hydrochlorofluorocarbons can be broken by collision with a hydroxyl radical:
$CHClF_2 + HO\bullet \rightarrow CClF_2\bullet + H_2O$ [1]

e **i** [1]

ii [1]

$O{=}O{\rightarrow}O$

The oxygen molecule has an O=O double bond, whilst the bonds in the ozone molecule are intermediate in strength between O—O and O=O. Photolysis of oxygen therefore requires a higher energy photon (shorter wavelength) than the photolysis of ozone. [1]

f Initiation steps: photolysis of CFC-11
$CFCl_3 + UV \rightarrow \bullet CFCl_2 + Cl\bullet$ (photodissociation of CFC-11) [1]

Chain propagation steps: breakdown of ozone by chlorine atom (radical), followed by regeneration of chlorine atom (radical) by collision with an oxygen atom (radical).
$Cl\bullet + O_3 \rightarrow ClO\bullet + O_2$ [1]
$ClO\bullet + O\bullet \rightarrow O_2 + Cl\bullet$ [1]
Termination steps: formation of chlorine 'sink' compounds – chlorine nitrate and hydrogen chloride.
$ClO\bullet + \bullet NO_2 \rightarrow ClONO_2$ (chlorine nitrate) [1]
$Cl\bullet + CH_4 \rightarrow HCl + \bullet CH_3$ [1]
$Cl\bullet + H_2 \rightarrow HCl + H\bullet$ [1]

g Chlorine nitrate; hydrogen chloride [2]

h Little mixing of air at poles leads to very low temperatures (polar vortex). [1]
Water and nitric acid condense to form polar stratospheric clouds (PSCs). [1]
The surface of a PSC catalyses decomposition of chlorine sink compounds such as chlorine nitrate. [1]
Concentration of $ClO\bullet$ radicals increases. [1]
When sunlight returns in spring ozone depletion by $ClO\bullet$ is rapid. [1]
During summer $ClO\bullet$ radicals are used up, reforming sink compounds; [1]
and PSCs disperse because of increased temperatures. [1]
Therefore the rate of ozone depletion falls and the hole repairs itself through natural ozone formation processes. [1]

Q6 **a** Biochemical oxygen demand describes the quantity of oxygen used when the organic material in the water is decayed by microorganisms. [1]
A small BOD corresponds to a small amount of organic matter present, i.e. the water is quite pure. [1]
A high BOD corresponds to impure water, as much of the oxygen present will be used up in decaying the organic material, with less available for the fish. [1]

b $(CH_2O)_n(aq) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$ [2]

c Process of eutrophication:
Fertilizer run-off contains nitrogen and phosphorus compounds. [1]
These are nutrients which lead to increased growth of algae. [1]
Algae dies and is consumed by aerobic bacteria. [1]
Oxygen in water is used up. [1]
Fish and other higher organisms have insufficient oxygen and populations decline. [1]
Environmental effects: (any three of following):

- water becomes turbid (cloudy) with a green tint
- water has unpleasant taste and smell
- the presence of algae makes water difficult to treat for drinking
- leisure activities (fishing, boating) are affected
- decline in fish populations
- increase in organisms such as sludge worms, maggots and eels
- formation of foul smelling substances such as thioalcohols, ammonia and hydrogen sulfide. [3]

d Higher temperature decreases the solubility of oxygen. [1]
Microorganisms respire more quickly, speeding up decay processes. [1]
Enzymes in microorganisms may denature leading to inability to digest food molecules. [1]
The change in temperature may effect spawning cycle of fish. [1]

e Point source pollution: the source of pollution is clearly identifiable at one point. [1]
Example: a chemical factory releasing toxic substances into a river. [1]
Non-point source pollution: water collects pollutants over a larger area, and thus cannot be attributed to single source. [1]
Examples: acid rain, which pollutes lakes; fertilizer run-off, where a river accumulates large quantities of nitrogen and phosphorus compounds as it flows through farmland. [1]

f Any four of following: heavy metals; pesticides (insecticides, herbicides); chemical wastes (dioxins, PCBs); organic wastes (sewage); fertilizers. [4]

Q7 **a** Aerobic bacteria are used to oxidize the organic matter in waste water. [1]
A large surface area (e.g. a bed of stones) or constant mixing are used to maximize aeration. [1]
This is necessary for efficient oxidation. [1]

b **i** $Ca^{2+}(resin) + 2Hg^{+}(aq) \rightarrow Ca^{2+}(aq) + 2Hg^{+}(resin)$ [2]
ii $Fe^{3+}(aq) + 3OH^{-}(aq) \rightarrow Fe(OH)_3(s)$ [2]
iii $K_{sp} = [Fe^{3+}(aq)]\,[OH^{-}(aq)]^3$ [2]
iv $[Fe^{3+}(aq)]\,[OH^{-}(aq)]^3 = 4 \times 10^{-38}$
Let $[Fe^{3+}] = x$
$[OH^{-}]$ is therefore $3x$

$x \times (3x)^3 = 4 \times 10^{-38}$
$27x^4 = 4 \times 10^{-38}$
$x = 1.96 \times 10^{-10}$
so $[Fe(OH)_3] = 1.96 \times 10^{-10}$ mo dm^{-3}
M_r $(Fe(OH)_3) = 107$
Solubility $= 1.96 \times 10^{-10} \times 107 = 2.1 \times 10^{-8}$ g dm^{-3} [3]

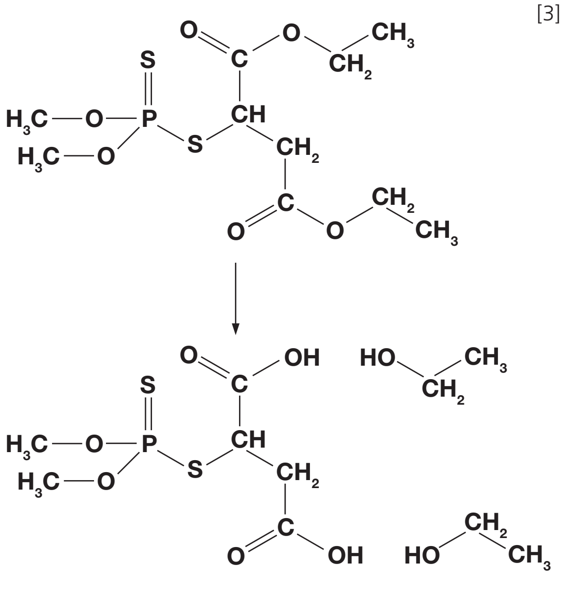
v $[Fe^{3+}(aq)]\,[OH^-(aq)]^3 = 4 \times 10^{-38}$
Let $[Fe^{3+}] = y$
$[OH^-]$ is 0.1 (from the sodium hydroxide)
$y \times (0.1)^3 = 4 \times 10^{-38}$
$y = 4 \times 10^{-35}$
so $[Fe(OH)_3] = 4 \times 10^{-35}$ mol dm^{-3}
M $(Fe(OH)_3) = 107$ g mol^{-1}
Solubility $= 4 \times 10^{-35} \times 107 = 4.3 \times 10^{-33}$ g dm^{-3} [3]

vi Solubility of iron(III) hydroxide in sodium hydroxide is much lower. [1]
Therefore the iron(III) ions will precipitate out of solution at much lower concentration. [1]
Almost all iron(III) can therefore be removed from the waste water by addition of sodium hydroxide. [1]

Q8 a i DDT:
No –OH groups/benzene rings [1]
Not water-soluble [1]
Not removed in urine [1]
Fat-soluble/accumulates in fatty tissues [1]

ii 1,4-Dioxin:
Symmetrical molecule/non-polar molecule [1]
Not water-soluble [1]
Not removed in urine [1]
Fat-soluble/accumulates in fatty tissues [1]

b [3]



Q9 a i Hydrocarbons and oxides of nitrogen. [2]
ii Ultraviolet radiation. [1]
iii It is an oxidizing smog due to the presence of ozone, a powerful oxidizing agent. [2]

b i Two from: ketones, aldehydes, peroxyacylnitrates (PANs) and nitric acid. [2]
ii Eye irritation and irritation to the respiratory system. [1]
iii Promote the use of public transport. [1]

c i $•NO_2 \rightarrow •NO + •O$
$•O + H_2O \rightarrow 2OH•$
$•OH + •NO_2 \rightarrow HNO_3$
$•OH + RH \rightarrow R• + H_2O$
$R• + O_2 \rightarrow ROO•$
$RCOOO• + •NO_2 \rightarrow RCOOONO_2$ [3]

ii Nitrogen oxides are also responsible for removal of ozone.
$NO• + O \rightarrow NO• + O_2$ [1]
Therefore the rate of ozone removal is slower when nitrogen oxide concentration falls, so the amount of ozone in smog increases. [1]

Q10 a Clay contains negatively charged sites which can reversibly bind metal ions (cations). [1]
$Clay^-(s) + M^+(aq) \rightleftharpoons Clay–M(s)$ [1]

b Hydrogen ions will displace the metal ions and release them into aqueous solution. [1]
$Clay–M(s) + H^+(aq) \rightleftharpoons Clay–H(s) + M^+(aq)$ [1]

c $H^+(aq) + OH^-(aq) \rightarrow H_2O$ (l)
Amount of NaOH $= 0.023$ dm$^3 \times 0.100$ mol dm^{-3}
$= 2.3 \times 10^{-3}$ mol
Amount of $H^+(aq)$ in 25.0 cm^3 sample $= 2.3 \times 10^{-3}$ mol
Amount of $H^+(aq)$ in 250 cm^3 solution (20 g of soil)
$= 2.3 \times 10^{-2}$ mol
Amount of $H^+(aq)$ in 100 g of soil $= 2.3 \times 10^{-2} \times 5$
$= 0.115$ mol
Mass of $H^+(aq) = 0.115$ mol $\times 1.01$ g mol^{-1}
$= 0.11615$ g $= 116$ mg [3]

d i Aluminium ions release hydrogen ions due to the action of hydrolysis:
$[Al(H_2O)_6]^{3+}(aq) \rightleftharpoons [Al(H_2O)_5OH]^{2+}(aq) + H^+(aq)$
The addition of hydrogen ions will displace the equilibrium to the left, thereby removing the excess hydrogen ions. [2]

ii Humus (soil organic matter) can act as a buffer. [1]
Any two from:
- Humus can absorb harmful organic and inorganic chemicals present in the soil.
- Humus provides energy for bacteria and fungi and acts as a source of nutrients.
- It helps to bind soil particles and absorbs heat. [2]

e i Pale green iron(II) ions are oxidized to brown iron(III) ions. [2]
ii Silicon and oxygen [2]

26 Food chemistry

Q1 a i $2H^+ + \text{chlorophyll-}Mg^{2+} \rightarrow \text{chlorophyll-}2H^+ + Mg^{2+}$ [1]
ii $\text{chlorophyll-}2H^+ + Zn^{2+} \rightarrow 2H^+ + \text{chlorophyll-}Zn^{2+}$ [1]

b This can occur because zinc and magnesium ions have identical charges and similar ionic radii. [1]
(The porphyrin ring presumably can tolerate these changes.)

c Red and blue light are being absorbed; [1]
green light is reflected [1]

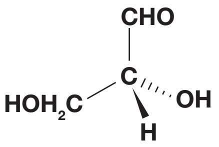
d Transitions involving the π electrons of the conjugated porphyrin ring. [1]
(The metal plays little role in determining the colour of the chlorophyll.)

Q2 a Any two from: toffee, caramel, fudge, milk chocolate. [2]
b Any named amino acid – e.g. glycine, cysteine or lysine [1]
a named reducing sugar – e.g. glucose or lactose. [1]

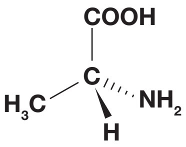
c Any two from:
- temperature
- the nature of the amino acid
- lysine browns the most; cysteine browns the least
- the rate of reaction is also affected by the water content of the reaction mixture. [2]

d Production of desirable brown colours; [1]
and pleasant smells and flavours [1]

Q3 a



D-glyceraldhyde [1]



D-alanine [1]

b An equimolar mixture of the enantiomers of the same compound. [1]

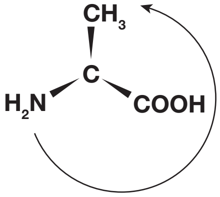
c D/L relates to the difference in spatial configuration of the enantiomers. [1]
+/– or *d*/*l* relates to the direction of rotation of plane-polarized light. [1]

d L form; [1]
tasteless [1]

e Odour; [1]
toxicity [1]

f Rotate the plane of plane-polarized light in opposite directions. [1]

g *S* [1]
[1]



h Glycine [1]
It lacks a chiral centre. [1]

Q4 a Carotenoids/carotenes [1]

b 11 [1]

c In humans: to act as a precursor for vitamin A synthesis. [1]
In plants: to act as an accessory pigment and prevent chlorophyll from damage. [1]

d i Transitions involving π electrons: $\pi \rightarrow \pi^*$ [1]

ii Repeated addition occurs. [1]
This causes loss of the double bonds, and hence the conjugated π system responsible for the colour. [1]

e Lycopene is a non-polar molecule due to its hydrocarbon structure. [1]
It does not possess any polar functional groups that would allow hydrogen bonding with water. [1]

f By reacting with an excited form of molecular oxygen and converting it back to the ground state. [1]

Q5 a A kinetically stable mixture of one phase in another, largely immiscible, phase. [1]

b i A mixture of a solid in a liquid; [1]
starch in warm water [1]

ii A mixture of two immiscible liquids, in which one liquid is distributed as droplets in the other liquid; [1]
cream [1]

iii mixture of a liquid and a gas which is dispersed through the liquid; [1]
shaving foam [1]

c i It contains hydrophobic and hydrophilic groups or soluble in fats/oils and water. [1]
Acts as an interface between the fat/oil and water and reduces the surface tension. [1]

ii Lecithin [1]
Phospholipid [1]

Q6 a Hemoglobin is involved in oxygen transport in blood; [1]
myoglobin is involved in oxygen storage in muscle. [1]

b The lone pairs on the oxygen atom are donated and shared with the iron atom. [1]
Dative (coordinate) covalent bonds are formed. [1]

c 2+ [1]

d Purple due to the presence of a heme group containing an iron(III) ion. [1]

e The oxygen converts the myoglobin; [1]
into bright red oxymyoglobin. [1]

f i The period of time that maintains the expected quality desired by the consumer. [1]

ii Two from: flavour, odour, texture, colour, mass [2]

iii It coats the meat with antimicrobial compounds that prevent bacterial growth. [1]

Q7 a i Cyanidin chloride [1]

ii They all lack covalently bonded sugar residues. [1]

b i Conjugation is increased. [1]

ii π bonding orbital to π antibonding orbital (π^*) [1]

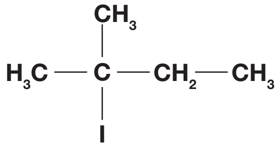
iii It decreases. [1]

c Phenol group [1]

d Temperature [1]
The presence of metal ions and the formation of complexes. [1]

27 Further organic chemistry

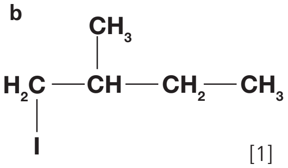
Q1 a [1]



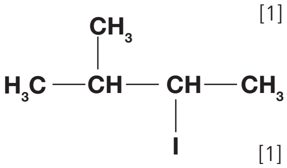
$H_3C-C(CH_3)(I)-CH_2-CH_3$

The organic intermediate is a tertiary carbocation. [1]
More thermodynamically stable. [1]
The three alkyl groups are electron donating/releasing – positive inductive effect. [1]

b [1]



$H_2C(I)-CH(CH_3)-CH_2-CH_3$ [1]

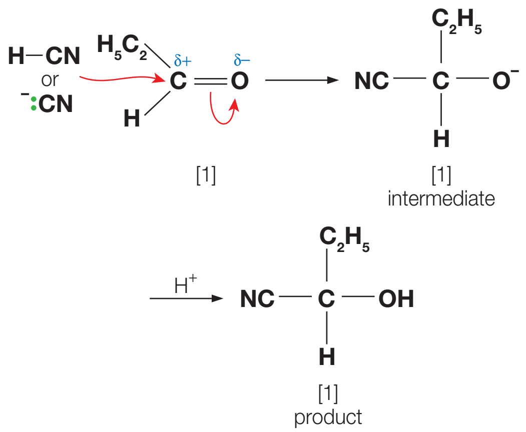


$H_3C-CH(CH_3)-CH(I)-CH_3$ [1]

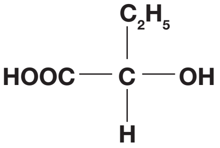
c i Similarities – any two from: both double bonds are made up of one σ bond and one π bond; the electrons are at 120° to the two other bonds attached to the carbon atoms; the carbon atom is sp^2 hybridized in both bonds. [2]
Differences – any two from: the shared pair of electrons are closer to the oxygen atom in butanal; the bond in butanal is polar; the C=O bond is shorter and stronger than the C=C bond. [2]

ii Butene: electrophilic addition [1]
Butanal: nucleophilic addition [1]

Q2 **a** Nucleophilic addition [2]



b [1]



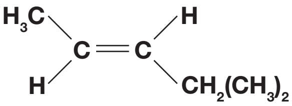
c A racemic (50 : 50/equimolar) mixture of enantiomers is formed. [1]

Q3 **a** The alkyl group (propyl) is electron releasing/causes a positive inductive effect. [1]
The electron density on the nitrogen atom is greater and the lone pair is more available for dative bond formation. [1]

b The —NO_2 groups are electron withdrawing/cause negative inductive effect. [1]
The negative charge is delocalized around the benzene ring. [1]

c CH_3I undergoes nucleophilic substitution far more readily than C_6H_5I with hydroxide ions. [1]
In CH_3I the incoming nucleophile can approach the carbon atom and displace the iodine/the π electrons of the benzene ring repel hydroxide ions and prevent attack on the carbon atom bonded to the iodine. [1]
The carbon–iodine bond in iodobenzene has partial double bond character/a lone pair of electrons from the iodine atom is partially delocalized into the benzene ring. [1]

Q4 **a** **i** [1]



ii Any two from:

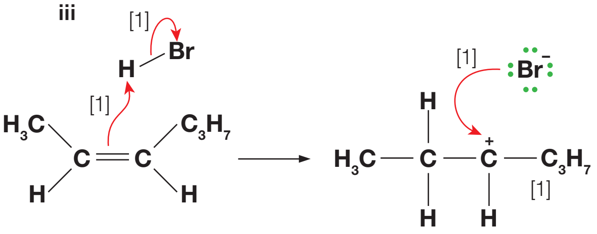
- Restricted rotation around a carbon–carbon double bond
- The functional groups cannot rotate since it would involve breaking a π bond
- The groups at both ends of the molecule are different so rotation would produce a different molecule [2]

iii *cis* [1]; hex-2-ene [1] [2]

b **i** Electrophilic addition [1]

ii $CH_3C^*HBrCHCH_2CH_2CH_3$ [1]
It contains an asymmetric or chiral carbon atom/a carbon atom bonded to four different atoms or functional groups. [1]

iii



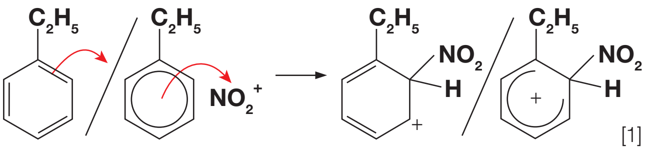
[4]

iv The rule states that the hydrogen adds to which ever carbon atom in the double bond already has the most hydrogen. [1]
In this example both carbon atoms have one hydrogen atom/the same number of hydrogen atoms/both are formed via secondary carbocations [1]

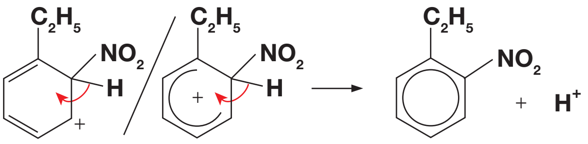
Q5 **a** Concentrated sulfuric acid and concentrated nitric acid [2]

b $HNO_3 + H_2SO_4 \longrightarrow H_2NO_3^+ + HSO_4^-$
$H_2NO_3^+ \longrightarrow H_2O + NO_2^+$;
or $HNO_3 + H_2SO_4 \longrightarrow H_2O + NO_2^+ + HSO_4^-$
or $HNO_3 + 2H_2SO_4 \longrightarrow H_3O^+ + NO_2^+ + 2HSO_4^-$ [1]

c

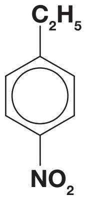


[1]

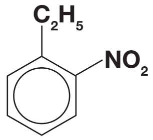


[1]

d **i**

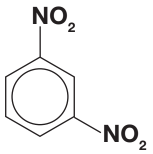


[1]



[1]

ii [1]



e The ethyl functional group, $-C_2H_5$, is electron-releasing/has a positive inductive effect. [1]
It increases attraction between the π electrons in the ring and the active electrophile, NO_2^+. [1]

Q6 **a** The electrons are spread over, and bond together, more than two atoms. [1]

b Physical: all C—C bonds the same length/same strength [1] intermediate between values for double and single bonds. [1]
Chemical: benzene undergoes electrophilic substitution not addition. [1]
There is only one isomer of 1,2-dibromobenzene not two. [1]

c Reaction of OH^- with (chloromethyl)benzene is faster; [1] similar to aliphatic halogenoalkanes. [1]
Chlorobenzene does not react because a π orbital on Cl interacts with the delocalized π orbital of benzene. [1]

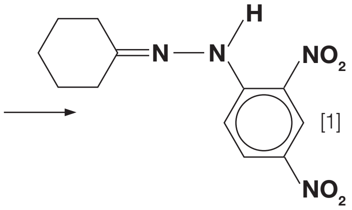
d 3-chloronitrobenzene [1]
Chlorination of nitrobenzene is slower than for benzene. [1]
The nitro group deactivates the ring; [1]
because it is an electron-withdrawing group. [1]
It has three highly electronegative atoms and no non-bonding pair of electrons on the nitrogen/it directs the substitution to the 3 and 5 positions. [1]

Q7 a $CH_3CH{=}CHCH_3$ and $CH_3CH_2CH{=}CH_2$ [2]
b React a halogenoalkane with fine turnings of magnesium set up for reflux/or equation:
$CH_3Br + Mg \longrightarrow CH_3MgBr$ [1]
in a solvent of dry ether – apparatus should be dry and protected by a drying agent. [1]
c $CH_3CH_2C(CH_3)(OH)CH_3$ [1]
Ketone is $CH_3CH_2COCH_3$ [1]
Grignard reagent, e.g. CH_3MgBr [1]

Q8 a Any two from:
- All C—C bonds in structure B are the same length /0.139 (nm) (long).
- Structure A would have different C—C bond lengths/0.154 and 0.134 (nm).
- Benzene does not have different C—C bond lengths. [2]

b $\Delta H^{\ominus}_2 = (6 \times \text{C—H}) + (3 \times \text{C{=}C}) + (3 \times \text{C—C}) = -5352\,\text{kJ mol}^{-1}$ [1]
$\Delta H^{\ominus}_4 = -31\,\text{kJ mol}^{-1}$
$\Delta H^{\ominus}_3 = (6 \times 717) + (3 \times 436) = +5610\,\text{kJ mol}^{-1}$ [1]
$\Delta H^{\ominus}_1 = 5610 + (-5352) + (-31) = +227\,\text{kJ mol}^{-1} = \Delta H^{\ominus}_f$ [1]
c Difference is $+177\,\text{kJ mol}^{-1}$; the actual structure of benzene is not structure **A**. [1]
It is more stable/stabilized by delocalization of the π electron cloud. [1]

Q9 a Addition–elimination/condensation [1]
b



[1] + H_2O [1]

c The (crystalline) solid has a characteristic melting point. [1]

Q10 a $CH_3CH_2CH_2CH_2OH \longrightarrow CH_3CH_2CH{=}CH_2 + H_2O$ [1]
phosphoric acid, 180 °C [1]
$CH_3CH_2CH{=}CH_2 + HBr \longrightarrow CH_3CH_2CHBrCH_3$ [1]
inert solvent [1]
b $CH_3CH_2COCH_3 + HCN \longrightarrow CH_3CH_2C(OH)(CN)CH_3$ [1]
nucleophilic addition [1]
$CH_3CH_2C(OH)(CN)CH_3 \longrightarrow CH_3CH_2C(OH)(COOH)CH_3$ [1]
acid hydrolysis [1]
Product is 2-hydroxy-2-methylbutanoic acid [1]
c i Step 1: alkylation of benzene [1]
using C_2H_5Cl/anhydrous $AlCl_3$ [1]
Step 2: nitration of ethylbenzene [1]
under reflux with nitrating mixture of concentrated nitric and sulfuric acids [1]
ii Step 1: nitration of benzene [1]
under reflux with nitrating mixture of concentrated nitric and sulfuric acids [1]
Step 2: alkylation of nitrobenzene [1]
using C_2H_5Cl/anhydrous $AlCl_3$ [1]
iii Step 1: alkylation of benzene [1]
using C_2H_5Cl/anhydrous $AlCl_3$ [1]
Step 2: acylation of ethylbenzene [1]
using C_2H_5COCl/anhydrous $AlCl_3$ [1]
d Step 1: acylation of methylbenzene [1]
using CH_3COCl/anhydrous $AlCl_3$ [1]
Step 2: reaction of product with 2,4-dinitrophenylhydrazine [1]
to produce the 2,4-dinitrophenylhydrazone derivative [1]