- The ideal gas law is $PV = nRT$, where n represents the amount of gas and R represents the gas constant. The molar mass or relative molecular mass of a gas can be found by weighing a known volume of the gas.
- The concentration of a solution is expressed in grams per cubic decimetre ($g\,dm^{-3}$) or moles per cubic decimetre ($mol\,dm^{-3}$).
- Concentration $(mol\,dm^{-3}) = \dfrac{\text{amount (mol)}}{\text{volume of solution } (dm^3)}$
- Titrations can be used to determine the reacting volumes of solution and, from the volumes and concentrations, the equation for the reaction. Titration of a solution of unknown concentration against a standard solution, with the equation for the reaction, allows the unknown concentration to be calculated.
- A back titration allows the determination of the concentration of a reactant of unknown concentration by reacting it with an excess volume of another reactant of known concentration. The resulting mixture is then titrated, taking into account the excess of reagent which is present. Back titrations are used when the sample under analysis is insoluble in water.
- Many stoichiometry calculations require the following approach:
  - Translate the mass or volume (of gas) of a given reactant into an amount (mol).
  - Use the stoichiometric ratio from the balanced equation to deduce the amount (mol) of the required reactant or product.
  - Reconvert to a mass, volume (for a gas), concentration ($mol\,dm^{-3}$ or $g\,dm^{-3}$) or percentage purity.

■ Examination questions – a selection

Paper 1 IB questions and IB style questions

Q1 Avogadro's number, L, is 6.02×10^{23} and the relative atomic mass of calcium is 40. What is the mass of one mole of calcium atoms?
A $40/L$ grams   C $L/40$ grams
B $40L$ grams   D 40 grams

Q2 What is the mass in grams of a single molecule of propane, C_3H_8?
A $7.3 \times 10^{-25}\,g$ C $6.02 \times 10^{-23}\,g$
B $44\,g$ D $7.3 \times 10^{-23}\,g$

Q3 Which sample has the greatest mass?
A $1.0\,mol$ of N_2H_4 C $3.0\,mol$ of NH_3
B $2.0\,mol$ of N_2 D $25.0\,mol$ of H_2
Standard Level Paper 1, May 99, Q1

Q4 Which of the following samples contains the smallest number of atoms?
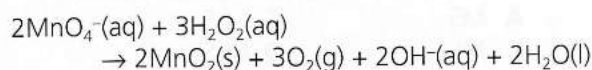A $1g\,H_2$ C $1g\,S_8$
B $1g\,O_2$ D $1g\,Br_2$

Q5 How many molecules are there in $180\,g$ of H_2O?
A 6.0×10^{22} C 6.0×10^{24}
B 6.0×10^{23} D 6.0×10^{25}
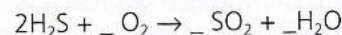**Standard** Level Paper 1, May 00, Q1

Q6 Hydrogen peroxide, H_2O_2, reacts with manganate(VII) ions, MnO_4^-, in basic solution according to the following equation:

$2MnO_4^-(aq) + 3H_2O_2(aq)$
$\rightarrow 2MnO_2(s) + 3O_2(g) + 2OH^-(aq) + 2H_2O(l)$

How many moles of hydrogen peroxide would be needed to produce eight moles of water?
A one   C three
B two   D twelve

Q7 Hydrogen sulfide, H_2S, reacts with oxygen to form sulfur dioxide and water as shown below.

$2H_2S + \_\,O_2 \rightarrow \_\,SO_2 + \_H_2O$

What is the whole number coefficient for oxygen when this equation is balanced?
A 1   C 3
B 2   D 6
Standard Level Paper 1, May 99, Q4

Q8 A certain compound has a molecular mass of $56\,g\,mol^{-1}$. Which of the following cannot be an empirical formula for this compound?
A BH_3 C MgN_2H_4
B C_3H_4O D HCl

STOECHIOMETRIE

②

Q9 Which of the following is an empirical formula?
  A N_2F_2 **C** C_2H_4O
  B $C_2H_4O_2$ **D** C_2N_2

Q10 $2Cl_2$ represents:
  A two chlorine molecules
  B two chlorine atoms
  C two chloride ions
  D four free chlorine atoms

Q11 Which one of the following equations correctly represents the combustion of calcium (a member of group 2) in oxygen?
  A $Ca(s) + O_2(g) \rightarrow CaO_2(s)$
  B $Ca(s) + 2O(g) \rightarrow CaO_2(s)$
  C $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$
  D $Ca(s) + O(g) \rightarrow CaO(s)$

Q12 Which one of the following equations is **not** correctly balanced?
  A $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
  B $Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$
  C $Fe^{2+}(aq) + Ag^+(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$
  D $Fe^{2+}(aq) + Cl_2(g) \rightarrow Fe^{3+}(aq) + 2Cl^-(aq)$

Q13 What is the total number of ions present in the formula, $Fe_2(SO_4)_3$?
  A 5   **C** 2
  B 3   **D** 6

Q14 32.0 grams of sulfur (atomic mass of 32.0) combine with a metal, M (atomic mass of 40.0) to give a product which weighs 52.0 g. What is the empirical formula of the sulfide formed?
  A MS   **C** M_2S
  B MS_2 **D** M_2S_5

Q15 An unknown element M combines with oxygen to form the compound MO_2. If 36.0 g of element M combines exactly with 16.0 g of oxygen, what is the atomic mass of M in grams?
  A 12.0   **C** 24.0
  B 16.0   **D** 72.0

Q16 When 16.00 grams of hydrogen gas reacts with 64.0 grams of oxygen gas in a reaction (atomic masses are H = 1.00, O = 16.00), what will be present in the resulting mixture?
  A H_2, H_2O, and O_2 **C** O_2, H_2O
  B H_2, H_2O **D** H_2, O_2

Q17 2.4 g of magnesium metal reacted vigorously when heated with excess iron(III) oxide, Fe_2O_3. What mass of metallic iron could be produced in this process?
  A 2.8 g   **C** 5.6 g
  B 3.7 g   **D** 8.4 g

Q18 One molecule of a small protein contains 63 atoms of carbon. The mass percentage of carbon in the protein is 55.74%. What is the molar mass of the protein?
  A 1357 g mol^{-1} **C** 821.3 g mol^{-1}
  B 421.7 g mol^{-1} **D** 756.6 g mol^{-1}

Q19 Nitrogen(II) oxide, NO, is made from the oxidation of NH_3:

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$

An 8.5 g sample of NH_3 gives 15.0 g of NO. What is the percentage yield of NO?
  A 40%   **C** 80%
  B 60%   **D** 100%

Q20 Which one of the following is an incorrect assumption of the kinetic theory of gases?
  A Atoms or molecules travel in straight lines between collisions but are in overall random motion.
  B Atoms or molecules of a gas are much smaller than the average distances between them.
  C Collisions between atoms or molecules of a gas and the containing vessel are perfectly elastic.
  D In a given gas, all particles have the same kinetic energy at a given temperature.

Q21 1000 cm³ of hydrogen gas (hydrogen molecules, H_2) contains Z molecules at room temperature and pressure. What will be the number of atoms in 500 cm³ of radon gas (radon atoms) at the same temperature and pressure? (Assume both gases behave ideally.)
  A Z   **C** Z/2
  B 2Z   **D** Z/4

Q22 Which graph shows how the average kinetic energy of the atoms or molecules varies with absolute temperature (in kelvin) for an ideal gas?
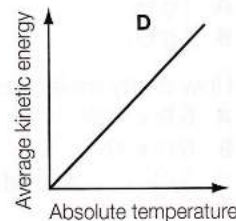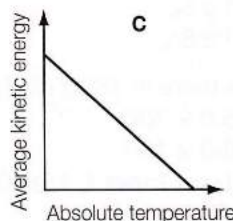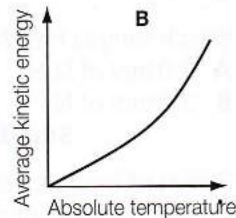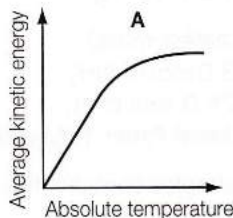
Q23 Under what conditions of temperature and pressure will a real gas behave most like an ideal gas?

	Temperature	Pressure
A	Low	Low
B	High	Low
C	High	High
D	Low	High

Q24 For an ideal gas, which variables are inversely proportional to each other (if all other factors remain constant)?

A P, V C V, T
B P, T D n, P

Q25 When compared at the same pressure and temperature, which one of the following physical properties has the same value for H_2, and for D_2? [D = $_1^2$H].

A average molecular speed
B relative molecular mass
C collision rate between molecules
D average kinetic energy of molecules

Q26 A 350 cm³ sample of helium gas is collected at 22.0 °C and 99.3 kPa. What volume would this gas occupy at stp?

A 318 cm³   C 477 cm³
B 450 cm³   D 220 cm³

Q27 A 27.0 g sample of an unknown carbon–hydrogen compound was burned in excess oxygen to form 88.0 g of CO_2 and 27.0 g H_2O. What is the possible molecular formula of the hydrocarbon?

A CH_4 C C_2H_6
B C_4H_6 D C_6H_6

Q28 1000 cm³ of ammonia gas combines with 1250 cm³ of oxygen to produce two gaseous compounds with a combined volume of 2500 cm³, all volumes being measured at 200 °C and 0.500 atm pressure. Which of the following equations fits these facts?

A $4NH_3 + 7O_2 \rightarrow 4NO_2 + 6H_2O$
B $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$
C $4NH_3 + 5O_2 \rightarrow 2N_2O_2 + 6H_2O$
D $4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$

Q29 In order to dilute 40.0 cm³ of 0.600 mol dm⁻³ HCl(aq) to 0.100 mol dm⁻³, what volume of water must be added?

A 60 cm³   C 200 cm³
B 160 cm³   D 240 cm³

Q30 What is the concentration of nitrate ions in 0.500 cm³ of 0.60 mol dm⁻³ $Fe(NO_3)_3$ solution?

A 0.60 mol dm⁻³   C 1.20 mol dm⁻³
B 0.90 mol dm⁻³   D 1.8 mol dm⁻³

Q31 If 35.50 cm³ of a NaOH solution are required for the neutralization of a 25.00 cm³ sample of 0.200 mol dm⁻³ H_2SO_4, what is the concentration of the NaOH?

A 0.143 mol dm⁻³   C 0.429 mol dm⁻³
B 0.282 mol dm⁻³   D 0.895 mol dm⁻³

Q32 How many grams of AgCl would be precipitated if an excess of $AgNO_3$ solution were added to 55.0 cm³ of 0.200 mol dm⁻³ KCl solution?

[Molar mass of silver chloride = 143.32 g mol⁻¹]

A 1.58 g   C 6.43 g
B 1.11 g   D 7.80 g

Q33 The temperature of an ideal gas sample is changed from 100 °C to 200 °C at constant pressure. What is the ratio of the final volume to the initial volume?

A 1:2   C 1.27:1
B 4:1   D 1:1.27

Q34 At stp (i.e. 0 °C and 1 atm pressure (101 kPa)), it was found that 1.15 dm³ of a gas weighed 3.96 g. What is its molar mass?

A 77 g mol⁻¹   C 47 g mol⁻¹
B 39 g mol⁻¹   D 4 g mol⁻¹

Q35 A sample of argon gas in a sealed container of fixed volume is heated from 50 to 250 °C. Which quantity will remain constant?

A average speed of the atoms
B pressure of the gas
C average kinetic energy of the atoms
D density of the argon

Paper 2 IB questions and IB style questions

Q1 10 cm³ of ethene, C_2H_4, is burned in 40 cm³ of oxygen, producing carbon dioxide and some liquid water. Some oxygen remained.

a Write the equation for the complete combustion of ethene. [2]
b Calculate the volume of carbon dioxide and the volume of oxygen remaining. [2]

Q2 a Write an equation for the formation of zinc iodide from zinc and iodine. [1]
b 100.0 g of zinc is allowed to react with 100.0 g of iodine producing zinc iodide. Calculate the amount (in moles) of zinc and iodine, and hence determine which reactant is in excess. [3]
c Calculate the mass of zinc iodide that will be produced. [1]

Higher Level Paper 2, May 04, Q3

4

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Q3 A balloon, which can hold a maximum of 1000 cm³ of nitrogen before bursting, contains 955 cm³ of nitrogen at 5 °C.

a Determine whether the balloon will burst if the temperature is increased to 30 °C. [3]

b Use the kinetic theory to explain what happens to the molecules of nitrogen inside the balloon as the temperature is increased to 30 °C. [2]

Q4 Hydrated sodium carbonate has the formula $Na_2CO_3.nH_2O$. An experiment was performed to determine n, the amount of water of crystallization. A sample of 50.00 grams of hydrated sodium carbonate was dissolved in 250 cm³ of water. 20.00 cm³ of this solution reacted completely with 13.95 cm³ of 2.00 mol dm⁻³ hydrochloric acid.

$$Na_2CO_3(aq) + 2HCl(aq)$$
$$\rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$$
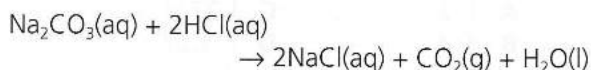
a Calculate the amount of hydrochloric acid reacted. [1]

b Calculate the amount of sodium carbonate in the 20 cm³ of the solution used in the reaction. [1]

c Calculate the concentration of sodium carbonate in the sample. [1]

d Calculate the molar mass of the hydrated sodium carbonate. [1]

e Calculate the value of n. [2]

Q5 a Aqueous XO_4^{3-} ions form a precipitate with aqueous silver ions. Write a balanced equation for the reaction, including state symbols. [1]

b When 41.18 cm³ of a solution of aqueous silver ions with a concentration of 0.2040 mol dm⁻³ is added to a solution of ions, 1.172 g of the precipitate is formed.

i Calculate the amount (in moles) of Ag^+ ions used in the reaction. [1]

ii Calculate the amount (in moles) of the precipitate formed. [1]

iii Calculate the molar mass of the precipitate. [2]

iv Determine the relative atomic mass of X and identify the element. [2]

Higher Level Paper 2, Nov 03, Q2

Examination questions – a selection

Paper 1 IB questions and IB style questions

Q1 Which statement is correct about the isotopes of an element?
A They have the same mass number.
B They have the same numbers of protons and neutrons in the nucleus.
C They have more protons than neutrons.
D They have the same electron arrangement or configuration.

Q2 A chemical element with the symbol X has the electron arrangement 2,8,6. Which chemical species is this chemical element most likely to form?
A the ion X^{3+} C the compound H_2X [$2H^+X^{2-}$]
B the ion X^{6+} D the compound XF_8 [$X^{8+}8F^-$]

Q3 Which of the following particles contains more electrons than neutrons?
I ^1_1H II $^{35}_{17}Cl^-$ III $^{39}_{19}K^+$
A I only   C I and II only
B II only   D II and III only
                Standard Level Paper 1, May 00, Q6

Q4 What information about the structure of a helium atom can be gained from its emission spectrum?
A Most of the mass of the atom is in its nucleus.
B A helium atom contains two electrons and two protons.
C The electrons in the helium atom are held near the nucleus.
D The electrons may exist in any of several energy levels.

Q5 An element has the electron arrangement 2,8,6. What is the element?
A C   C S   B P   D Ar

Q6 Which is an incorrect statement about the atomic emission spectrum of hydrogen?
A The frequency of each line depends on the difference in energy between the higher and lower energy levels.
B The spectrum consists of several series of lines.
C Electronic transitions to the level $n = 2$ give rise to lines in the visible region.
D It is a continuous spectrum.

Q7 What is the correct number of each particle in a fluoride ion, $^{19}F^-$?

	Protons	Neutrons	Electrons
A	9	10	8
B	9	10	9
C	9	10	10
D	9	19	10

Standard Level Paper 1, Nov 03, Q5

Q8 What fraction of a radioisotope will remain after three half-lives?
A 1/16   C 1/3
B 1/8   D 3/4

Q9 Why was the Bohr theory of the atom developed?
A To account for changes in gas volumes with temperature.
B To account for the ratios by mass of elements in compounds.
C To account for the emission or line spectrum of hydrogen atoms.
D To account for chemical formulas.

Q10 A particular element consists of two isotopes: 72% of mass number 85 and 28% of mass number 87. What is the expected range of the relative atomic mass?
A less than 85
B between 86 and 87
C between 85 and 86
D more than 88

Q11 How many valence electrons (electrons in the outermost shell) are present in the element of atomic number 14?
A 4   C 2   B 3   D 1

Q12 Which one of the following atoms will have the same number of neutrons as an atom of $^{88}_{38}Sr$?
A $^{91}_{39}Y$ C $^{89}_{38}Sr$ B $^{87}_{37}Rb$ D $^{84}_{36}Kr$

Q13 Which statement is correct for the emission spectrum of the hydrogen atom?
A The lines converge at lower energies.
B The lines are produced when electrons move from lower to higher energy levels.
C The lines in the visible region involve electron transitions into the energy level closest to the nucleus.
D The line corresponding to the greatest emission of energy is in the ultraviolet region.
                Standard Level Paper 1, Nov 03, Q6

Q14 Naturally occurring chlorine consists of the isotopes chlorine-35 and chlorine-37. The relative atomic mass of chlorine is 35.5. Which one of the following statements is true?
A The chlorine-35 and chlorine-37 atoms are present in equal amounts.
B The ratio of chlorine-37 atoms to chlorine-35 atoms is 2 : 1.
C The ratio of chlorine-37 to chlorine-35 atoms is 37/35.
D There are three times as many as chlorine-35 atoms as chlorine-37 atoms.

Q15 Which statement is correct about a line emission spectrum?
- **A** Electrons neither absorb nor release energy as they move from low to high energy levels.
- **B** Electrons absorb energy as they move from high to low energy levels.
- **C** Electrons release energy as they move from low to high energy levels.
- **D** Electrons release energy as they move from high to low energy levels.

Standard Level Paper 1, Nov 05, Q6

Q16 Which electronic transition within a hydrogen atom requires the greatest energy?
- **A** $n = 1 \rightarrow n = 2$
- **C** $n = 2 \rightarrow n = 3$
- **B** $n = 3 \rightarrow n = 5$
- **D** $n = 5 \rightarrow n = \infty$

Q17 Which of the following radioisotopes is used in nuclear medicine to image the thyroid gland?
- **A** iodine-131
- **C** fluorine-18
- **B** carbon-14
- **D** uranium-235

Q18 The atomic numbers and mass numbers for four different nuclei are given in the table below. Which two are isotopes?

	Atomic number	Mass number
I	101	258
II	102	258
III	102	260
IV	103	259

- **A** I and II
- **C** III and IV
- **B** II and III
- **D** I and IV

Standard Level Paper 1, Nov 98, Q6

Q19 All isotopes of uranium have the same:
- I number of protons
- II number of neutrons
- III mass number
- **A** I only
- **C** III only
- **B** II only
- **D** I and III only

Q20 Which is the correct sequence for some of the various stages that typically occur in the analysis of an element during mass spectrometry?
- **A** vaporization, electron bombardment, acceleration, deflection, detection
- **B** electron bombardment, vaporization, acceleration, deflection, detection
- **C** vaporization, electron bombardment, deflection, acceleration, detection
- **D** deflection, acceleration, electron bombardment, vaporization, detection

Paper 2 IB questions and IB style questions

Q1 The element bromine exists as the isotopes ^{79}Br and ^{81}Br, and has a relative atomic mass of 79.90.
- **a** Copy and complete the following table to show the numbers of sub-atomic particles in the species shown. [3]

	An atom of ^{79}Br	An ion of ^{81}Br$^-$
Protons		
Neutrons		
Electrons		

- **b** State and explain which of the two isotopes ^{79}Br and ^{81}Br is more common in the element bromine. [1]

Standard Level Paper 2, Nov 05, Q3

Q2 The element silver has two isotopes, $^{107}_{47}$Ag and $^{109}_{47}$Ag, and a relative atomic mass of 107.87.
- **a** Define the term *isotope*. [1]
- **b** State the number of protons, electrons and neutrons in $^{107}_{47}$Ag$^+$. [2]
- **c** State the name and the mass number of the isotope relative to which all atomic masses are measured. [1]

Q3 A sample of iridium is analysed in a mass spectrometer. The first and last processes in mass spectrometry are vaporization and detection.
- **a i** State the names of the second and third processes in the order in which they occur in a mass spectrometer. [2]
- **ii** Outline what occurs during the second process. [2]
- **iii** State and explain which one of the following ions will undergo the greatest deflection (under the same conditions in a mass spectrometer): ^{191}Ir$^+$ or ^{193}Ir$^+$ [2]
- **b** The sample of iridium is found to have the following composition of stable isotopes:

Isotope	Ir-191	Ir-193
Relative abundance/%	37.1	62.9

- **i** Define the term *relative atomic mass*. [2]
- **ii** Calculate the relative atomic mass of this sample of iridium, giving your answer to two decimal places. [2]
- **c** Iridium-192 is a short-lived radioisotope used to treat cancer. Define the term radioisotope and name another radioisotope used in nuclear medicine. [2]

Q4 Describe the emission or line spectrum of gaseous hydrogen atoms and explain how this is related to the energy levels in the atom. [3]

Examination questions – a selection

Paper 1 IB questions and IB style questions

Q1 Which element shows chemical behaviour similar to calcium?
A strontium   C sodium
B chlorine   D boron

Q2 The following are three statements concerning the periodic table.
I   The horizontal rows are called periods and the vertical columns are called groups.
II   Electronegativity decreases down any group and across a period from left to right.
III   Reactivity increases down all groups.

Which of the above is/are true?
A I, II and III   C II and III only
B I and II only   D I only

Q3 Which is the correct trend (left to right) across period 3 for the oxides?
A basic to acidic   C increasingly basic
B acidic to basic   D neutral to acidic

Q4 What happens when chlorine water is added to an aqueous solution of potassium iodide?
A No reaction occurs because chlorine is less reactive than iodine.
B Chlorine molecules are oxidized to chloride ions.
C Iodide ions are oxidized to iodine molecules.
D A purple precipitate of iodine is formed.

Q5 Which of the following determines the order in which the elements are arranged in the modern form of the periodic table?
A relative atomic mass
B mass number
C atomic number
D chemical reactivity

Q6 Which is a correct statement about the element with an atomic number of 20?
A It is in group 4.
B It is in group 2.
C It is a transition metal.
D It is in group 7 and is a halogen.

Q7 In general, atomic radii decrease:
A within a group from lower to higher atomic number
B within a period from lower to higher atomic number
C with an increase in the number of isotopes of an element
D with an increase in the shielding of the nuclear charge

Q8 When the elements are listed in order of increasing reactivity with air, the correct order is:
A Na, K, Cs
B Cs, K, Na
C Cs, Na, K
D K, Cs, Na

Q9 For which type of isoelectronic ions do ionic radii decrease with increasing nuclear charge?
A positive ions only
B negative ions only
C neither positive or negative ions
D both positive and negative ions

Q10 Which properties are typical of most non-metals in period 3 (Na to Ar)?
I   They form ions by gaining one or more electrons.
II   They are poor conductors of heat and electricity.
III   They have high melting points.
A I and II only   C II and III only
B I and III only   D I, II and III

Standard Level Paper 1, Nov 05, Q7

Q11 On the periodic table, groups of elements show similarities in their chemical properties. This can be best explained by the:
A differences in the number of protons in the nucleus of the atoms
B similarities in the results of emission spectrum analysis of gaseous samples of a group
C similarities in the electronic structures of the atoms
D differences in the number of neutrons in the nucleus of the atoms

Q12 Which atom has the smallest atomic radius?
A $_{31}Ga$ B $_{20}Ca$ C $_{35}Br$ D $_{37}Rb$

Q13 Which one of the following series represents the correct size order for the various iodine species?
A $I < I^- < I^+$ C $I^+ < I < I^-$
B $I < I^+ < I^-$ D $I^- < I < I^+$

Q14 Which one of the following will be observed as the atomic number of the elements in a single group of elements on the periodic table increases?
A an increase in atomic radius
B an increase in ionization energy and hence decrease in reactivity
C a decrease in ionic radius
D an increase in electronegativities

Q15 Which of the following properties of the halogens increase from F to I?

    I atomic radius
    II melting point
    III electronegativity

    A I only   **C** I and III only
    B I and II only   **D** I, II and III

        Standard Level Paper 1, Nov 03, Q7

Q16 In general, how do ionization energies vary as the periodic table is crossed from left to right?

    A They remain constant.
    B They increase.
    C They increase to a maximum and then decrease.
    D They decrease.

Q17 0.01 mol samples of the following oxides were added to separate 1 dm³ portions of water. Which will produce the most acidic solution?

    A $Al_2O_3(s)$ **C** $Na_2O(s)$
    B $SiO_2(s)$ **D** $SO_3(g)$

Q18 Which property increases with increasing atomic number for both the alkali metals and the halogens?

    A melting points
    B first ionization energies
    C electronegativities
    D atomic radii

Q19 Which one of the following elements has the lowest first ionization energy?

    A Li   **C** B   **B** Na   **D** Mg

Q20 Barium is an element in group 2 of the periodic table (below strontium with atomic number 56). Which of the following statements about barium is not correct?

    A Its first ionization energy is lower than that of strontium.
    B It has two electrons in its outermost energy level.
    C Its atomic radius is smaller than that of strontium.
    D It forms a chloride with the formula $BaCl_2$.

Paper 2 IB questions and IB style questions
(IB *Chemistry data booklet* required)

Q1 **a** **i** Define the term *ionization energy*. [2]
        ii Write an equation, including state symbols, for the process occurring when measuring the first ionization energy of aluminium. [1]
    b Explain why the first ionization energy of magnesium is greater than that of sodium. [3]
    c Lithium reacts with water. Write an equation for the reaction and state two observations that could be made during the reaction. [3]
        Standard Level Paper 2, Nov 05, Q4

Q2 **a** **i** Explain why the ionic radius of bromine is less than that of selenium. [2]
        ii Explain what is meant by the term *electronegativity* and explain why the electronegativity of fluorine is greater than that of chlorine. [3]
    b For each of the following reactions in aqueous solution, state one observation that would be made, and deduce the equation.
        i The reaction between chlorine and potassium iodide. [2]
        ii The reaction between silver ions and bromide ions. [2]
    c Deduce whether or not each of the reactions in **b** is a redox reaction, giving a reason in each case. [4]

Q3 **a** What factors determine the size of an atom or ion? [3]
    b **i** Explain why the ionic radius of sodium is much smaller than its atomic radius. [2]
        ii Explain why the cations of group 1 increase in size with increasing atomic number. [2]
    c Explain why the ionic radius of Mg^{2+} is less than that of Na^+. [2]
    d Arrange the following species in order of increasing size:
        i N, N^{3-} [1]
        ii Fe, Fe^{2+} and Fe^{3+} [1]

Q4 Describe and explain the variation in ionic radius of the elements across period 3 from sodium to chlorine. [6]

Liaison ①

■ A dipole–dipole force is the attraction between oppositely charged regions of polar molecules. Dipole–dipole forces involve permanent dipoles.

■ Van der Waals' forces of attraction arise due to the random movement of electrons which create temporary dipoles which induce dipoles in adjacent molecules. The result is a very weak force of attraction between now oppositely charged regions of the two molecules.

■ Van der Waals' forces depend on a molecule's size and shape. Van der Waals' forces are maximized in large molecules with linear shapes.

■ In metals, individual valence electrons separate from the individual atoms and become delocalized. Metallic bonding is the electrostatic attraction between the nuclei of metal ions and the delocalized valence electrons.

■ The strength of metallic bonding increases with the number of valence electrons and decreasing ionic radius.

■ The nature of metallic bonding accounts for the electrical conductivity of metals: the delocalized electrons flow when a voltage is applied.

■ The nature of metallic bonding accounts for the malleability of metals: the layers of ions slide over each other when a force is applied.

■ The physical properties of a compound or element are related to the type of bonds and/or intermolecular forces present. The type of bonding in a compound or element can be deduced from its physical properties.

■ Substances with a simple molecular structure have low melting points, high volatility and are electrical insulators. Some simple molecular substances hydrolyse with water.

■ Substances with a giant molecular structure have very high melting points, low volatility, and are electrical insulators under all conditions.

■ Ionic substances have high melting points, are brittle (when stressed), and only conduct electricity when molten or in aqueous solution (if soluble).

■ Graphite is an electrical conductor due to the presence of delocalized electrons in its layers.

■ Hydration is the process by which water molecules penetrate a solid lattice and attach themselves to the particles (atoms, ions or molecules).

■ Solvent molecules can be attached to particles by (i) van der Waals' forces, (ii) dipole–dipole interactions, (iii) hydrogen bonds, (iv) ion–dipole forces or (v) dative covalent bonds.

Examination questions – a selection

Paper 1 IB questions and IB style questions

Q1 Which compound contains ionic bonds?
  A magnesium chloride, $MgCl_2$
  B dichloroethane, CH_2Cl_2
  C ethanoic acid, CH_3COOH
  D silicon tetrabromide, $SiBr_4$

Q2 When CH_4, NH_3, H_2O are arranged in order of **decreasing bond angle**, what is the correct order?
  A CH_4, NH_3, H_2O
  B NH_3, H_2O, CH_4
  C NH_3, CH_4, H_2O
  D H_2O, NH_3, CH_4

Q3 In which of the following pairs does the second substance have the lower boiling point?
  A Cl_2, Br_2
  B H_2O, H_2S
  C C_3H_8, C_4H_{10}
  D CH_3OCH_3, CH_3CH_2OH

Q4 A group 1 element, X, bonds with a group 7 element, Y. What is the most likely formula and type of bonding in this compound?
  A XY_2 covalent
  B XY ionic
  C XY covalent
  D X_2Y ionic

Q5 In which of the following is there at least one double bond?
  I O_2
  II CO_2
  III C_2F_4
  A I only   **C** II and III only
  B III only   **D** I, II and III

Q6 According to VSEPR theory, which molecule would be expected to have the smallest bond angle?
  A H_2O **C** SiH_4
  B H_2CO **D** NH_3

Q7 In which of the following substances would hydrogen bonding would be expected to occur?
  I C_2H_6
  II CH_3CH_2COOH
  III CH_3OCH_3
  A II only   **C** II and III only
  B I and III only   **D** I, II and III

Q8 Why is the boiling point of ethane greater than that of neon?
  A The ethane molecule is polar.
  B Hydrogen bonds form between ethane molecules but are not present in liquid neon.
  C More electrons are present in ethane than in neon.
  D A molecule of ethane has a greater mass than a neon atom.

Q9 Which molecule has the greatest polarity?
  A fluorine
  B hydrogen fluoride
  C hydrogen iodide
  D tetrafluoromethane

Q10 Which is the best description of metallic bonding?
  A The attraction between positive and negative ions.
  B The attraction between protons and electrons.
  C The attraction between the nuclei of positive ions and delocalized valence electrons.
  D The attraction between nuclei and shared electron pairs.

Q11 Which compound is the most soluble in water?
  A ethane   **C** propan-1-ol
  B propane   **D** hexan-1-ol

Q12 Which statement is **not** true about metallic bonding?
  A It is present in mixtures of metals (alloys).
  B It results from the transfer of electrons from metal atoms to non-metal atoms.
  C It involves the delocalization of electrons.
  D It is electrostatic in nature.

Q13 Element X is in group 3 and element Y is in group 6 of the periodic table. Which is the most likely formula of the compound formed when X and Y react together?
  A XY   **B** X_2Y_3 **C** X_3Y_2 **D** XY_2

Q14 Which molecule contains a multiple bond?
  A H_2 **B** H_2O **C** C_2F_4 **D** C_2F_6

Q15 Which is **not** present in $C_2H_5OC_2H_5$ in the liquid state?
  A covalent bonding
  B van der Waals' forces
  C dipole–dipole attractions
  D hydrogen bonding

Q16 Chlorine has a lower boiling point than bromine. Which property of the two elements is responsible for this observation?
  A ionization energies
  B bond enthalpies
  C bond polarities
  D number of electrons

Q17 Which of the following molecules is planar?
  A NCl_3 **B** C_2H_4 **C** C_3H_6 **D** SF_6

Q18 A solid has a high melting point, does not conduct electricity as a solid, but does when it is dissolved in water. What type of substance is the solid?
  A ionic   **C** giant molecular
  B simple molecular **D** metallic

Q19 When the Lewis structure for $HCOOCH_3$ is drawn, how many bonds and how many lone pairs of electrons are present?
  A 8 and 4   **C** 5 and 5
  B 7 and 5   **D** 7 and 4

Q20 The angle between the two carbon–carbon bonds in CH_3CHCF_2 is closest to:
  A 180°   **B** 120°   **C** 109°   **D** 90°

Q21 The compounds X, Y, Z, have approximately the same molar mass.
    X C_5H_{12}
    Y $CH_3CH_2CH_2CH_2OH$
    Z $CH_3OCH_2CH_2CH_3$
  When these compounds are arranged in order of increasing boiling point (lowest boiling point first), the correct order is:
  A X, Z, Y   **C** Y, Z, X
  B X, Y, Z   **D** Z, Y, X

Q22 What is the formula for the compound formed by strontium and nitrogen?
  A SrN   **B** Sr_2N **C** Sr_2N_3 **D** Sr_3N_2

Q23 The molar masses of C_2H_6, CH_3OH and CH_3F are similar. How do their boiling points compare?
- **A** $C_2H_6 < CH_3OH < CH_3F$
- **B** $CH_3F < CH_3OH < C_2H_6$
- **C** $CH_3OH < CH_3F < C_2H_6$
- **D** $C_2H_6 < CH_3F < CH_3OH$

Q24 Which intermolecular forces exist in dry ice, $CO_2(s)$?
- **A** dipole–dipole interactions
- **B** covalent bonds
- **C** van der Waals' forces
- **D** hydrogen bonds

Q25 Which one of the following molecules would be expected to be linear?
- **A** H_2O_2 **B** NO_2 **C** SO_3 **D** CO_2

Q26 Which of the compounds H_2O, H_2S, H_2Se and HCl has the highest boiling point?
- **A** H_2O **B** H_2S **C** H_2Se **D** HCl

Q27 Which is an incorrect statement about carbon-60 (C_{60})?
- **A** It is a giant molecular substance.
- **B** It is a soft powder.
- **C** The surface of its molecules is composed of rings of five and six carbon atoms.
- **D** Van der Waals' forces of attraction hold the molecules in a lattice.

Q28 In general, the strengths of the following intermolecular forces and bonds increase in the order:
- **A** covalent bonds, hydrogen bonds, van der Waals' forces
- **B** covalent bonds, van der Waals' forces, hydrogen bonds
- **C** hydrogen bonds, covalent bonds, van der Waals' forces
- **D** van der Waals' forces, hydrogen bonds, covalent bonds

Q29 Given the following electronegativity values:
H: 2.1   N: 3.0   O: 3.5   F: 4.0
which bond has the greatest polarity?
- **A** O–H in H_2O **C** N–O in NO_2
- **B** N–F in NF_3 **D** N–H in NH_3

Q30 Which one of the following species has a triangular pyramidal geometry?
- **A** BCl_3 **B** NCl_3 **C** H_2Se **D** C_2H_2

Paper 2 IB questions and IB style questions

Q1 Describe the variation in melting points and electrical conductivities of the elements sodium to argon, and explain these variations in terms of their structures and bonding. [6]

Q2 a Draw electron dot structures for N_2 and F_2 and explain why F_2 is much more reactive than N_2. [3]
b Compare the polarity of the bonds N–F and C–F. Are the molecules NF_3 and CF_4 polar or non-polar? In all your answers give your reasons. [5]
Standard Level Paper 2, May 99, Q5

Q3 Explain at the molecular level why ethanol (C_2H_5OH) is soluble in water, but cholesterol $(C_{27}H_{45}OH)$ and ethane (C_2H_6) are not. [4]
Standard Level Paper 2, May 01, Q6

Q4 The elements potassium and fluorine and the compound potassium fluoride can be used to show the connection between bonding, structure and physical properties.
a Describe the type of bonding in potassium metal and explain why potassium is a good conductor of electricity. [4]
b Draw a Lewis structure for fluorine. Name and describe the bonding within and between the molecules in liquid fluorine. [4]
c Write the electronic structures of both potassium and fluorine and describe how the atoms combine to form potassium fluoride. [4]
d Explain why potassium fluoride does not conduct electricity until it is heated above its melting point. [1]

Q5 a The letters A, B, C and D represent four consecutive elements in the periodic table. The number of electrons in the highest occupied energy levels are:
A: 3   B: 4   C: 5   D: 6
Write the formula for:
i an ionic compound formed from A and C, showing the charges [2]
ii a covalent compound containing B and D. [1]
b State the number of protons, neutrons and electrons in the ion $^{12}_{6}C_{60}^{6-}$. [2]
c State the type of bonding in the compound SiF_4. Draw the Lewis structure for this compound. [3]
d Outline the principles of the valence shell electron pair repulsion (VSEPR) theory. [3]
e i Use the VSEPR theory to predict and explain the shape and bond angle of each of the molecules SBr_2 and C_2Br_2. [6]
ii Deduce whether or not each molecule is polar, giving a reason for your answer. [3]

◼ Hess's law states that the enthalpy change in a reaction depends only on the enthalpies of the reactant and products, and is therefore independent of the route or path taken.

◼ Chemical equations can be manipulated like algebraic equations: equations can be added or subtracted to eliminate chemical species. If a chemical reaction is reversed the sign of the enthalpy change is reversed.

◼ Enthalpy changes can be directly determined by measuring the heat released or absorbed during a reaction, using a calorimeter. Practical methods of measuring enthalpy changes for liquids and solutions involve letting the reaction mixture heat itself. For combustion reactions the reaction is allowed to heat a known quantity of water.

◼ Enthalpy changes can be derived using the following formula: $q = mc\Delta T$, where m represents the mass of water or solution, c represents the specific heat capacity of water and ΔT the change in temperature of the water or solution.

◼ The specific heat capacity is the amount of heat required to raise the temperature of a unit mass of the substance by one degree Celsius or one kelvin.

◼ The average bond enthalpy is the enthalpy change when one mole of covalent bonds between atoms of X and Y are broken in the gas phase:

$$X-Y(g) \rightarrow X(g) + Y(g); \qquad \Delta H = E(X-Y)$$

◼ The enthalpy change for a reaction is equal to the sum of the bond enthalpies of the bonds broken minus the sum of the bond enthalpies of bonds formed: $\Delta H = \Sigma$ average bond enthalpies of the reactants $- \Sigma$ average bond enthalpies of the products.

◼ In an exothermic reaction the bond enthalpies of the products are greater than the bond enthalpies of the reactants. In an endothermic reaction the bond enthalpies of the products are less than the bond enthalpies of the reactants.

◼ Examination questions – a selection

Paper 1 IB questions and IB style questions

Q1 Which of the following must have a negative value for an exothermic reaction?
A voltage of a voltaic cell   C enthalpy change
B change in state   D equilibrium constant

Q2 Which of the following is observed when the change in enthalpy is positive for the dissolving of a salt in water in an insulated copper beaker?
A Heat is evolved to the surroundings and the beaker feels cold.
B Heat is evolved to the surroundings and the beaker feels warm.
C Heat is absorbed from the surroundings and the beaker feels warm.
D Heat is absorbed from the surroundings and the beaker feels cold.

Q3 What is the specific heat capacity of an alcohol in $Jg^{-1}K^{-1}$ if 560.0 J of heat are required to raise the temperature of a 64.0 g sample of ethanol from 295.0 K to 310.0 K?
A 0.583   B 0.194   C 8.75   D 0.292

Q4 The following equation shows the formation of calcium oxide from calcium metal.
$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$ $\Delta H^{\oplus} = -1270\,kJ$
Which of these statements is true?

A 1270 kJ of energy are released for every mol of calcium reacted.
B 635 kJ of energy are absorbed for every mol of calcium oxide formed.
C 635 kJ of energy are released for every mol of oxygen gas reacted.
D 1270 kJ of energy are released for every 2 mol of calcium oxide formed.

Q5 Which of the following processes is/are endothermic?
I $H_2O(s) \rightarrow H_2O(g)$
II $CO_2(g) \rightarrow CO_2(s)$
III $N_2(g) \rightarrow 2N(g)$
A I only   C I and II only
B III only   D I and III only

Q6 All the following processes are exothermic **except**:
A $2C_2H_5(g) \rightarrow C_4H_{10}(g)$
B $F_2(g) \rightarrow 2F(g)$
C $Cl(g) + e^- \rightarrow Cl^-(g)$
D $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

Q7 Which of the following reactions would you expect to provide the largest amount of heat?
A $C_2H_6(l) + 7O_2(l) \rightarrow 4CO_2(g) + 6H_2O(g)$
B $C_2H_6(l) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$
C $C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$
D $C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$

Q8 Why does the temperature of boiling water remain constant even though heat is supplied at a constant rate?
  A Heat is lost to the surroundings.
  B The heat is used to break the covalent bonds in the water molecules.
  C Heat is also taken in by the container.
  D The heat is used to overcome the intermolecular forces of attraction between water molecules.
              Standard Level Paper 1, Nov 05, Q14

**Q9** When 0.050 mol of nitric acid is reacted with 0.050 mol of potassium hydroxide in water, the temperature of the system increases by 13.7 °C. Calculate the enthalpy of reaction in kJ mol^{-1}.

$HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$
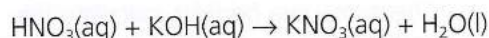
Assume that the heat capacity of the system was 209.2 J °C^{-1}.
  A +57.3 kJ mol^{-1} **C** −2.87 kJ mol^{-1}
  B +2.87 kJ mol^{-1} **D** −57.3 kJ mol^{-1}

Q10 What can be deduced about the relative stability of the reactants and products and the sign of ΔH, from the enthalpy level diagram below?
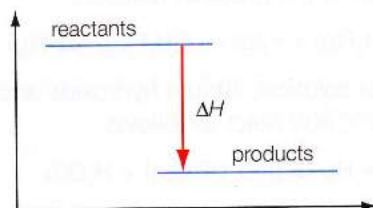
Relative stability	Sign of ΔH
A products more stable	−
B products more stable	+
C reactants more stable	−
D reactants more stable	+

              Standard Level Paper 1, May 99, Q16

Q11 The specific heat capacities of some metals are given below.

Metal	Specific heat capacity (J g^{-1} K^{-1})
copper	0.385
magnesium	1.020
mercury	0.138
platinum	0.130

If 100 kJ of heat is added to 10.0 g samples of each of the metals above, which are all at 25 °C, which metal will have the lowest temperature?
  A copper   **C** mercury
  B magnesium   **D** platinum

Q12 Consider the following bond energies:

C=C	615 kJ mol^{-1}
C−F	484 kJ mol^{-1}
C−C	348 kJ mol^{-1}
F−F	158 kJ mol^{-1}

Which one of the following gives the enthalpy change (in kJ mol^{-1}) for the addition reaction between fluorine and ethene (in the gaseous state)?
  A −615 − 158 + 348 + 2(484)
  B −615 − 158 − 348 −2(484)
  C 615 + 158 − 348 − 2(484)
  D 615 + 158 + 348 + 2(484)

Q13 Which one of the following statements is correct?
  A Breaking covalent bonds absorbs energy and making ionic bonds absorbs energy.
  B Bond breaking is endothermic and bond making is exothermic.
  C Bond breaking is exothermic and bond making is endothermic.
 **D** Breaking bonds releases energy and making bonds absorbs energy.

Q14 The bond energy for the H−F bond is equal to the enthalpy change for which process?
  A $H^+(g) + F^-(g) \rightarrow HF(g)$
  B $HF(g) \rightarrow H(g) + F(g)$
  C $\frac{1}{2}F_2(g) + \frac{1}{2}H_2(g) \rightarrow HF(g)$
  D $HF(g) \rightarrow \frac{1}{2}F_2(g) + \frac{1}{2}H_2(g)$

Q15 When a sample of a pure hydrocarbon (melting point 85 °C) cools, the temperature is observed to remain constant as it solidifies. Which statement accounts for this observation?
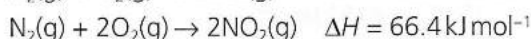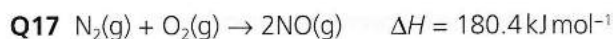 **A** The heat released in the change of state equals the heat loss to the surroundings.
  B The temperature of the system has fallen to room temperature.
  C The solid which forms insulates the system, preventing heat loss.
  D Heat is gained from the surroundings as the solid forms, maintaining a constant temperature.
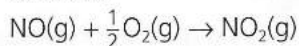
Q16 Consider the following equation:

$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$
$$\Delta H = 2824 \text{ kJ mol}^{-1}$$
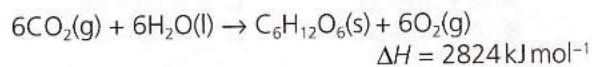
What is the enthalpy change associated with the production of 100.0 g of $C_6H_{12}O_6$?
  A 157 kJ   **C** 508 kJ
  B 282 kJ   **D** 1570 kJ
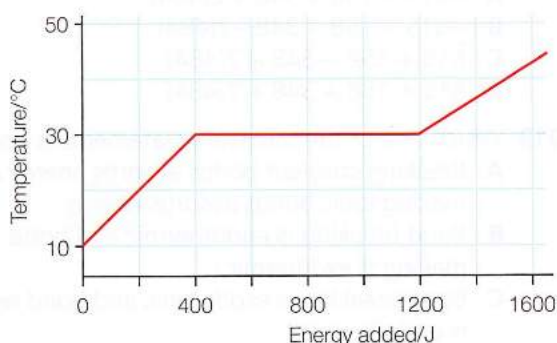
Thermo ③

Q17 $N_2(g) + O_2(g) \rightarrow 2NO(g)$ $\Delta H = 180.4 \text{ kJ mol}^{-1}$

$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$ $\Delta H = 66.4 \text{ kJ mol}^{-1}$

Use the enthalpy values to calculate ΔH for the reaction:

$NO(g) + \frac{1}{2}O_2(g) \rightarrow NO_2(g)$

A −57 kJ   C 57 kJ
B −114 kJ   D 114 kJ

Standard Level Paper 1, May 00, Q18

Q18

The heating curve for 10 g of a substance is given above. How much energy would be required to melt completely 40 g of the substance that is initially at 10 °C ?
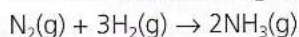
A 4800 J   C 1600 J   B 2400 J   D 800 J

Q19 The bond energies for H_2, I_2 and HI are 432, 149 and 295 kJ mol^{-1} respectively. From these data, what is the enthalpy change (in kJ) for the reaction below?

$H_2(g) + I_2(g) \rightarrow 2HI (g)$
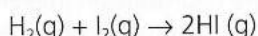
A +9   C −286
B +286   D −9

Q20 Consider the following reaction:

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

Bond enthalpies (in kJ mol^{-1}) involved in the reaction are:

N≡N a
H−H b
N−H c

Which expression could be used to calculate the enthalpy of reaction?

A $a + 3b − 6c$ C $a − 3b + 6c$
B $6c − a + 3b$ D $a + 3b − 2c$

Paper 2 IB questions and IB style questions

Q1 a i Explain what is meant by the term *standard enthalpy of reaction*. [3]

ii Describe an experiment to determine the enthalpy change of the reaction between dilute hydrochloric acid and aqueous sodium hydroxide. Show how the value

of ΔH would be calculated from the data obtained. [9]

iii Draw an enthalpy diagram for the neutralization reaction above. Indicate on your diagram the enthalpy change of the reaction and hence compare the relative stabilities of reactant and products. [4]

b Explain, giving one example, the usefulness of Hess's law in determining ΔH values. [4]

Standard Level Paper 2, May 00, Q5

Q2 a i Define the term *average bond enthalpy*. [3]

ii Explain why the fluorine molecule, F_2, is not suitable as an example to illustrate the term average bond enthalpy. [1]

b i Using values from Table 10 of the IB *Chemistry data booklet*, calculate the enthalpy change for the following reaction:
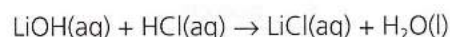
$CH_4(g) + F_2(g) \rightarrow CH_3F(g) + HF(g)$ [3]
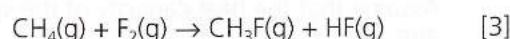
ii Sketch an enthalpy diagram for the reaction. [2]

iii Without carrying out a calculation, suggest, with a reason, how the enthalpy change for the following reaction compares with that of the previous reaction.

$CH_3F(g) + F_2(g) \rightarrow CH_2F_2(g) + HF(g)$ [2]

Q3 In aqueous solution, lithium hydroxide and hydrochloric acid react as follows.

$LiOH(aq) + HCl(aq) \rightarrow LiCl(aq) + H_2O(l)$

The data below is from an experiment to determine the standard enthalpy change of this reaction.

50.0 cm³ of a 0.500 mol dm⁻³ solution of LiOH was mixed rapidly in a glass beaker with 50.0 cm³ of a 0.500 mol dm⁻³ solution of HCl.
Initial temperature of each solution = 20.6 °C
Final temperature of the mixture = 24.1 °C

a State, with a reason, whether the reaction is exothermic or endothermic. [1]

b Explain why the solutions were mixed rapidly. [1]

c Calculate the enthalpy change of this reaction in kJ mol⁻¹. Assume that the specific heat capacity of the solution is the same as that of water. [4]

d Identify the major source of error in the experimental procedure described above. Explain how it could be minimized. [2]

e The experiment was repeated but with an HCl concentration of 0.520 mol dm⁻³ instead of 0.500 mol dm⁻³. State and explain what the temperature change would be. [2]

Cinétique ①.

- Rates are determined by measuring the change in concentration of a reactant or product over a period of time. Methods include titrations, pressure measurement (for gases), colour changes and mass changes.
- The rate of a reaction can be affected by: concentration of reactants in solution, surface area of solid reactants, temperature, light, pressure (of gaseous reactants) and presence of catalysts.
- The activation energy is the minimum kinetic energy that reactants need to form products. The activation energy is the energy barrier to a reaction (or elementary step) and controls the rate of the reaction.
- Collision theory accounts for the variation in rate with temperature, surface area, concentration of reactants and use of catalysts.
  - Increasing the temperature. This increases the collision rate but a more significant factor is the increase in the proportion of reactants having a combined kinetic energy greater to or equal to the activation energy.
  - Increasing concentrations. This increases the chances of the reactant molecules colliding and reacting.
  - Increasing the surface area (for reactions that involve solids). This increases the surface area over which the reaction can occur.
  - Using a catalyst. This increases the number of successful collisions.
- Requirements for a chemical reaction to occur: the reactants must collide together with sufficient kinetic energy; for many reactions the reactants must have the correct collision geometry (steric factor).
- The Maxwell–Boltzmann distribution is a graph of the proportion of a sample of molecules that has a specific value of kinetic energy. Graphs at different temperatures show that the proportion of molecules with high kinetic energy increases with temperature.
- A catalyst is a chemical substance that increases the rate of a reaction without itself undergoing any permanent chemical change. Catalysts provide a new reaction mechanism with a lower activation energy than the uncatalysed reaction.

Examination questions – a selection

Paper 1 IB questions and IB style questions

Q1 The reaction between *excess* magnesium carbonate and hydrochloric acid can be followed by measuring the volume of carbon dioxide produced with time. The results of one such reaction are shown below. How does the rate of this reaction change with time and what is the main reason for this change?
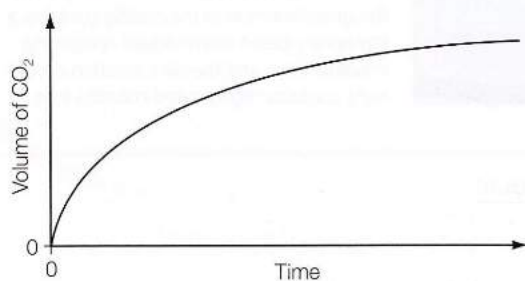
A The rate decreases with time because the acid becomes more dilute.

B The rate increases with time because the acid becomes more dilute.

C The rate decreases with time because the magnesium carbonate particles get smaller.

D The rate increases with time because the magnesium carbonate particles get smaller.

Q2 Zinc reacts with sulfuric acid as shown below.

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$
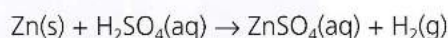
Two identical samples of zinc powder were reacted with separate samples of *excess* acid as follows:

Reaction 1: zinc added to $0.5 \, mol \, dm^{-3}$ sulfuric acid
Reaction 2: zinc added to $2.0 \, mol \, dm^{-3}$ sulfuric acid
What is the same for reactions 1 and 2?

A total mass of hydrogen formed

B total reaction time

C average rate of production of hydrogen gas

D initial reaction rate

Q3 The reaction between nitrogen and chlorine in the atmosphere under normal conditions is extremely slow.
Which statement best explains this?
A The concentration of nitrogen is too high, which slows the reaction.
B The molar mass of nitrogen molecules is less than that of chlorine molecules.
C Nitrogen and chlorine molecules are both non-polar molecules.
D Very few nitrogen and chlorine molecules have sufficient kinetic energy to react.

Q4 What is the action of a catalyst?
A provides a new mechanism or pathway for the reaction
B increases the enthalpy change, ΔH, for the reaction
C changes the sign of the enthalpy change, ΔH, for the reaction
D decreases the activation energy, E_a, for the forward reaction only

**Q5** Which statement is correct for a collision between reactant particles leading to a reaction?
A Colliding particles must have different energies.
B All reactant particles must have the same energy.
C Colliding particles must have a kinetic energy higher than the activation energy.
D Colliding particles must have the same velocity.
Standard Level Paper 1, Nov 05, Q19

Q6 Which change of condition will decrease the rate of the reaction between excess calcium granules and dilute hydrochloric acid?
A increasing the amount of calcium
B increasing the concentration of the hydrochloric acid
C decreasing the temperature of the acid
D converting the calcium granules into powder

Q7 *Excess* magnesium was added to a beaker of aqueous hydrochloric acid on a balance. A graph of the mass of the beaker and contents was plotted against time (line 1).
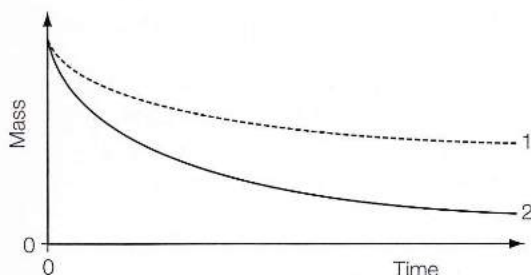
What change in the experiment could give line 2?
I the same mass of magnesium but in smaller pieces
II the same volume of a more concentrated solution of hydrochloric acid
III a lower temperature
A I only
B II only
C III only
D none of the above
Standard Level Paper 1, Nov 03, Q19

Q8 The rate of a reaction between two gases increases when the temperature is increased and a catalyst is added. Which statements are correct for the effect of these changes on the reaction?

	Increasing the temperature	Adding a catalyst
A	collision frequency increases	activation energy increases
B	activation energy increases	activation energy does not change
C	activation energy does not change	activation energy decreases
D	activation energy increases	collision frequency increases

Standard Level Paper 1, Nov 03, Q20

Q9 Hydrogen peroxide undergoes decomposition in the presence of a manganese(IV) oxide catalyst to produce oxygen and water. The graph below shows how the total volume of oxygen generated varies with time.
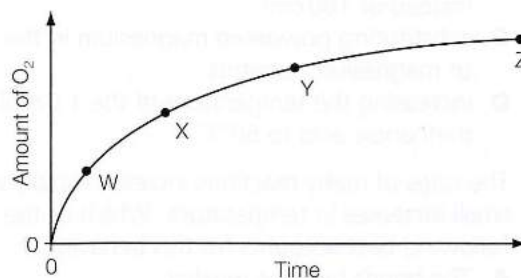
Which point on the graph represents the fastest rate of reaction?
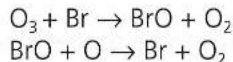A W   B X   C Y   D Z

Q10 For the reaction:
$$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$
which gives the fastest rate?
A 1.0 mol dm^{-3} HCl and CaCO$_3$ chips
B 2.0 mol dm^{-3} HCl and CaCO$_3$ chips
C 2.0 mol dm^{-3} HCl and CaCO$_3$ powder
D 4.0 mol dm^{-3} HCl and CaCO$_3$ powder

Cinétique ③.

Q11 In the reaction shown below, which species may be acting as a catalyst?

$O_3 + Br \rightarrow BrO + O_2$
$BrO + O \rightarrow Br + O_2$

A O_2 **B** Br   **C** O   **D** BrO

Q12 An increase in the rate of a reaction is generally brought about by increases in which of the following?

    I   reactant concentration
    II   particle size
    III   temperature

A I and II only   **C** I, II and III
B I and III only   **D** II and III only

Q13 All of the following statements are correct **except**:
    A The smaller the size of the reacting particles, the faster the reaction.
    B The rate of a chemical reaction can be decreased by decreasing the temperature.
    C Increasing the pressure often decreases the rates of reaction involving gases.
    D The rates of most chemical reactions decrease with time.

Q14 When 100 cm³ of 1.0 mol dm⁻³ methanoic acid, HCOOH, are added to 1 gram of magnesium turnings at 20 °C, hydrogen gas is slowly produced. All of the following will increase the initial rate of hydrogen production **except**:
    A substituting 1.0 mol dm⁻³ hydrochloric acid for 1.0 mol dm⁻³ methanoic acid
    B using 300 cm³ of 1.0 mol dm⁻³ methanoic acid instead of 100 cm³
    C substituting powdered magnesium in the place of magnesium turnings
    D increasing the temperature of the 1.0 mol dm⁻³ methanoic acid to 60 °C

Q15 The rates of many reactions increase rapidly with small increases in temperature. Which of the following best accounts for this behaviour?
    A The bonds become weaker.
    B The activation energy decreases.
    C The number of molecular collisions increases.
    D The number of molecules with the necessary kinetic energy increases.

Q16 A catalyst:
    A creates another reaction pathway
    B is consumed in a chemical reaction
    C is always a transition element or compound
    D is always a solid

Q17 All of the following substances are observed to react with aqueous ethanoic acid at room temperature. Which one probably reacts fastest?
    A a strip of zinc metal
    B magnesium oxide powder
    C marble chips (calcium carbonate)
    D an aqueous solution of sodium hydroxide

Q18 What are the usual units for rate?
    A $mol^{-1} (dm^3)^2 s^{-1}$ **C** $mol\, dm^{-3} s^{-1}$
    B $mol\, dm^{-3}$ **D** $mol\, s^{-1}$

Q19 For which one of the following reactions would pressure measurements be the least sensitive for measuring the initial rate of reaction?
    A $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$
    B $N_2O_4(g) \rightarrow 2NO_2(g)$
    C $H_2(g) + I_2(g) \rightarrow 2HI(g)$
    D $2NCl_3(l) \rightarrow N_2(g) + 3Cl_2(g)$

Q20 Which statement explains why the speed of some chemical reactions is increased when the surface area of the reactant is increased?
    A This change increases the density of the reactant particles.
    B This change increases the concentration of the reactant.
    C This change exposes more reactant particles to a possible collision.
    D This change alters the electrical conductivity of the reactant particles.

Paper 2 IB questions and IB style questions

Q1 **a** Define the term *rate of reaction*.   [1]
    b The reaction between gases **C** and **D** is slow at room temperature.
       i Suggest **two** reasons why the reaction is slow at room temperature.   [2]
       ii A relatively small increase in temperature causes a relatively large increase in the rate of this reaction. State **two** reasons for this.   [2]
       iii Suggest **two** ways of increasing the rate of reaction between **C** and **D** other than increasing temperature.   [2]
   **Standard** Level Paper 2, Nov 05, Q2

Q2 The following results were obtained when 2.40 grams of magnesium ribbon were added to a large excess of dilute hydrochloric acid while the temperature was kept constant (by means of a water bath). The hydrogen gas was collected in a large gas syringe.

Time/seconds	Volume of hydrogen gas evolved/cm^3
0	0
20	900
40	1400
60	1720
80	1950
100	2100
120	2240
140	2240

a What was the volume of gas produced in
  i the first 20 second interval (from 0 seconds to 20 seconds) [1]
  ii the second 20 second interval (from 20 seconds to 40 seconds) [1]
  iii the third 20 second interval (from 40 seconds to 60 seconds)? [1]
b Explain why the volume of hydrogen gas changes in each of these 20 second intervals. [2]
c Why is the volume the same after 140 seconds and 120 seconds? [1]
d How would the initial rate of production of gas change if
  i the temperature were increased [1]
  ii the volume of acid were diluted with an equal volume of water [1]
  iii the same mass of magnesium powder were used instead of magnesium ribbon? [1]
e How would the final volume of gas change if
  i a greater mass of magnesium were used with the same volume of acid [1]
  ii a larger volume of acid were used with the same mass of magnesium? [1]

f An alternative method of monitoring the change of rate of this reaction is to perform the reaction in an open flask on an electronic balance. Suggest why this method is less preferable. [2]
g State and explain what the relationship is between the rate of consumption of hydrochloric acid and the rate of formation of hydrogen gas. [2]
h Deduce the ionic equation and explain why the reaction is very unlikely to proceed directly via this reaction in one step. [2]

Q3 The graph below was obtained when calcium carbonate reacted with dilute hydrochloric acid, under two different conditions, X and Y.
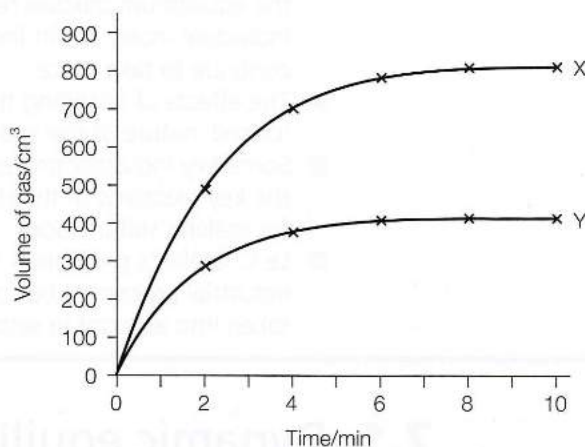
a i Name the gas produced in the reaction. [1]
  ii Write a balanced equation for the reaction occurring. [2]
b Identify the volume of gas produced and the time taken for the reaction under condition X to be complete.
c i Explain the shape of curve Y in terms of collision theory. [2]
  ii Suggest two possible reasons for the differences between curves X and Y. [2]

Examination questions – a selection

Paper 1 IB questions and IB style questions

Q1 Which statements are correct for a reaction at equilibrium?
- I   The forward and reverse reactions both continue.
- II   The rates of the forward and reverse reactions are equal.
- III   The concentrations of reactants and products are equal.

A I and II only   **C** II and III only
B I and III only   **D** I, II and III

Standard Level Paper 1, May 05, Q21

Q2 Which statement is always true for a chemical reaction that has reached equilibrium?
- **A** The yield of product(s) is less than 50%.
- **B** The rate of the reverse reaction is greater than the rate of the forward reaction.
- **C** The amounts of reactants and products do not change.
- **D** Both forward and reverse reactions have stopped.

Q3 Which statement(s) is/are true for a mixture of ice and water at equilibrium?
- I   The rates of melting and freezing are equal.
- II   The amounts of ice and water are equal.
- III   The same position of equilibrium can be reached by cooling water or heating ice.

**A** I only   **C** I and III only
B II only   **D** III only

Q4 Which of the following is a correct statement for the effect of a catalyst on a reversible chemical reaction?
- **A** It increases the amount of product.
- **B** It increases the activation energy of the reaction.
- **C** It allows the chemical reaction to reach equilibrium more quickly.
- **D** It only increases the rate of the forward reaction.

**Q5** Which statement is true about chemical reactions at equilibrium?
- **A** The forward and reverse reactions proceed at equal rates.
- **B** The forward and reverse reactions have stopped.
- **C** The concentrations of the reactants and products are equal.
- **D** The forward reaction is exothermic.

Q6 What changes occur when the temperature is increased in the following reaction at equilibrium?

$$Br_2(g) + Cl_2(g) \rightleftharpoons 2BrCl(g) \qquad \Delta H^\ominus = +14\,kJ\,mol^{-1}$$
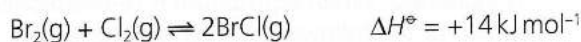
	Position of equilibrium	Value of equilibrium constant
A	Shifts towards the products	Decreases
B	Shifts towards the reactants	Decreases
C	Shifts towards the reactants	Increases
D	Shifts towards the products	Increases

Q7 Which changes will shift the position of equilibrium to the right in the following reaction?

$$2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$$
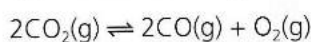
- I   adding a catalyst
- II   decreasing the oxygen concentration
- III   increasing the volume of the container

A I, II and III   **C** II and III only
B I and II only   **D** I and III only

Q8 The key reaction in the manufacture of sulfuric acid can be represented by the equation below.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H^\ominus = -197\,kJ\,mol^{-1}$$
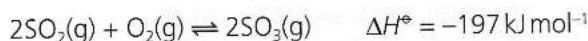
What happens when a catalyst is added to an equilibrium mixture from this reaction?
- **A** The rate of the forward reaction increases and that of the reverse reaction decreases.
- **B** The rates of both forward and reverse reactions increase.
- **C** The value of $\Delta H^\ominus$ increases.
- **D** The yield of sulfur trioxide increases.

Q9 $I_2(g) + 3Cl_2(g) \rightleftharpoons 2ICl_3(g)$

What is the equilibrium constant expression for the reaction above?

A $K_c = \dfrac{[I_2][Cl_2]^3}{[ICl_3]^2}$ **B** $K_c = \dfrac{2[ICl_3]}{[I_2]+3[Cl_2]}$

C $K_c = \dfrac{2[ICl_3]}{3[I_2][Cl_2]}$ **D** $K_c = \dfrac{[ICl_3]^2}{[I_2][Cl_2]^3}$

Q10 $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H^\ominus = -197\,kJ\,mol^{-1}$

According to the above information, what temperature and pressure conditions produce the greatest amount of SO_3?

	Temperature	Pressure
A	Low	Low
B	Low	High
C	High	High
D	High	Low

Q11 The volume of the reaction vessel containing the following equilibrium mixture

$$SO_2Cl_2(g) \rightleftharpoons SO_2(g) + Cl_2(g)$$

is increased. When equilibrium is re-established, which of the following will have occurred?
A The amount of $SO_2Cl_2(g)$ will have increased.
B The amount of $SO_2Cl_2(g)$ will have decreased.
C The amount of $Cl_2(g)$ will have remained unchanged.
D The amount of $Cl_2(g)$ will have decreased.

Q12 In which of the following reactions does the position of equilibrium remain unaffected by change in pressure?
A $2O_3(g) \rightleftharpoons 3O_2(g)$
B $2NO_2(g) \rightleftharpoons N_2O_4(g)$
C $2NO(g) + Cl_2(g) \rightleftharpoons 2NOCl(g)$
D $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$

Q13 For a gaseous reaction, the equilibrium constant expression is:

$$K_c = \frac{[O_2]^5[NH_3]^4}{[NO]^4[H_2O]^6}$$

Which equation corresponds to this equilibrium expression?
A $4NH_3 + 5O_2 \rightleftharpoons 4NO + 6H_2O$
B $4NO + 6H_2O \rightleftharpoons 4NH_3 + 5O_2$
C $8NH_3 + 10O_2 \rightleftharpoons 8NO + 12H_2O$
D $2NO + 3H_2O \rightleftharpoons 2NH_3 + \frac{5}{2}O_2$

Standard Level Paper 1, May 02, Q21

Q14 The following equilibrium can be set up in a closed container.

$$N_2O_4(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H^\oplus = +57\,kJ\,mol^{-1}$$
colourless   brown
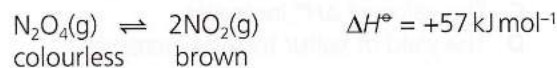
Which of the following changes would produce a darkening of the colour of the gaseous mixture in the container?
A an increase in temperature
B adding a catalyst
C an increase in pressure
D a decrease in temperature

Q15 Which of the following changes will shift the position of equilibrium of this reaction in the forward direction?

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g) \qquad \Delta H^\oplus = +181\,kJ\,mol^{-1}$$
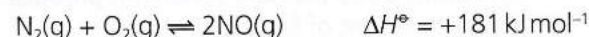
I   increasing the pressure
II   adding a catalyst
III increasing the temperature
A I only   C III only
B II only   D I and III

Standard Level Paper 1, Nov 01, Q21

Q16 $2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$

Methanol is made in industry by means of the reaction above. The equilibrium expression for this reaction is:

A $K_c = \dfrac{[CH_3OH]}{2[H_2][CO]}$ B $K_c = \dfrac{[CH_3OH]}{[H_2]^2[CO]}$

C $K_c = \dfrac{2[H_2][CO]}{[CH_3OH]}$ D $K_c = \dfrac{[H_2]^2[CO]}{[CH_3OH]}$

Q17 $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H^\oplus = -91.8\,kJ\,mol^{-1}$

The industrial synthesis of ammonia is based on the reaction above. Which factor(s) will increase the equilibrium concentration of ammonia?
I   increase in pressure
II increase in temperature
A I only   C neither I nor II
B both I and II   D II only

Q18 N_2O_4 and NO_2 produce an equilibrium mixture according to the equation below:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H^\oplus = +57\,kJ\,mol^{-1}$$

An increase in the equilibrium concentration of NO_2 can be produced by increasing which of the factors below?
I   pressure
II temperature
A neither I nor II   C I only
B both I and II   D II only

Q19 The hydration of ethene to ethanol occurs according to the following equation:

$$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$$
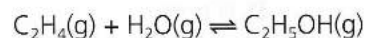
If this reaction is exothermic, which of the following sets of conditions would give the best equilibrium yield of ethanol?

|	Temperature/°C	Pressure/atm
A	1000	3
B	750	2
C	250	10
D	500	4

Q20 The smaller an equilibrium constant, K_c:
A the slower the reaction rate
B the lower the concentration of products at equilibrium
C the more endothermic the reaction
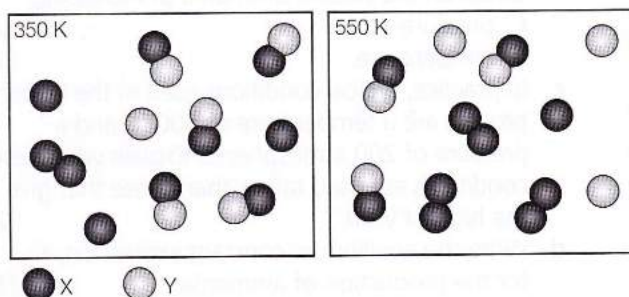D the faster the reactants are converted to products

Paper 2 IB questions and IB style questions

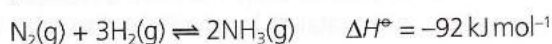**Q1 a** The diagrams below represent equilibrium mixtures for the reaction:

$$Y + X_2 \rightleftharpoons XY + X$$
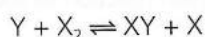
at 350 K and 550 K, respectively. Deduce and explain whether the reaction is exothermic or endothermic. [2]

● X ○ Y

b The equation for the main reaction in the Haber process is:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H^{\ominus} = -92\,kJ\,mol^{-1}$$

  i State **two** characteristics of a reversible reaction at equilibrium. [2]

  ii This reaction is described as *homogeneous*. State what is meant by the term *homogeneous*. [1]

  iii Write the equilibrium constant expression for the reaction. [2]

c When nitrogen and hydrogen are mixed together at room temperature and atmospheric pressure the reaction is very slow. In industry, typical values of pressure and temperature used can be 250 atmospheres and 450 °C.

  i State the effects on both the rate of reaction and the value of the equilibrium constant of increasing the temperature. [2]

  ii State the effects on both the rate of reaction and the value of the equilibrium constant of increasing the pressure. [2]

  iii Suggest why a pressure of 1000 atmospheres is not used. [1]

  iv State and explain what will happen to the position of the equilibrium if some of the ammonia is removed. [2]

d Name the catalyst used in the Haber process. State and explain its effect on the value of the equilibrium constant. [3]

e Use the collision theory to explain the effect of increasing the temperature on the rate of reaction between nitrogen and hydrogen. [3]

Q2 For the reversible reaction:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad \Delta H > 0$$

the equilibrium constant $K_c = 60$ at a particular temperature.

a Give the equilibrium expression. [1]

b For this reaction, what information does the value of K_c provide about the **relative** concentrations of the product and reactants at equilibrium? [1]

c What effect, if any, will an increase in pressure have on the **equilibrium position**? [1]

d Explain why an increase in temperature increases the value of the **equilibrium constant** for the above reaction. [1]

**Q3** The equation for one reversible reaction involving oxides of nitrogen is shown below:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g) \qquad \Delta H^{\ominus} = +57\,kJ\,mol^{-1}$$

Experimental data for this reaction can be represented on the following graph:
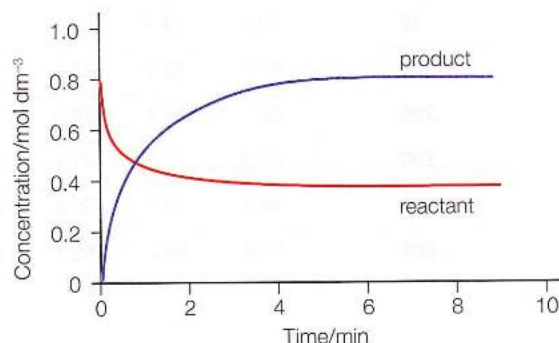
a Write an expression for the equilibrium constant, K_c, for the reaction. Explain the significance of the horizontal parts of the lines on the graph. State what can be deduced about the magnitude of K_c for the reaction, giving a reason. [4]

b Use Le Châtelier's principle to predict and explain the effect of increasing the temperature on the position of equilibrium. [2]

c Use Le Châtelier's principle to predict and explain the effect of increasing the pressure on the position of equilibrium. [2]

d State and explain the effects of a catalyst on the forward and reverse reactions, on the position of equilibrium and on the value of K_c. [6]

Standard Level Paper 2, Nov 05, Q6(a)

Equilibre ④.

Q4 a The following equilibrium is established at 1700°C.

$$CO_2(g) + H_2(g) \rightleftharpoons H_2O(g) + CO(g)$$
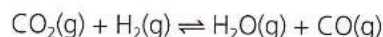
If only carbon dioxide gas and hydrogen gas are present initially, sketch on a graph a line representing rate against time for **i** the forward reaction *and* **ii** the reverse reaction until shortly after equilibrium is established. Explain the shape of each line. [7]

b K_c for the equilibrium reaction is determined at two different temperatures. At 850°C, $K_c = 1.1$ whereas at 1700°C, $K_c = 4.9$. On the basis of these K_c values explain whether the reaction is exothermic or endothermic. [3]

Q5 The table below gives information about the percentage yield of ammonia obtained in the Haber process under different conditions.

Pressure/atm	Temperature/°C			
	200	300	400	500
10	50.7	14.7	3.9	1.2
100	81.7	52.5	25.2	10.6
200	89.1	66.7	38.8	18.3
300	89.9	71.1	47.1	24.4
400	94.6	79.7	55.4	31.9
600	95.4	84.2	65.2	42.3

a From the table, identify which combination of temperature and pressure gives the highest yield of ammonia. [1]

b The equation for the main reaction in the Haber process is:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H \text{ is negative}$$
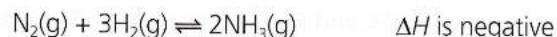
Use this information to state and explain the effect on the yield of ammonia of increasing

i pressure [2]
ii temperature. [2]

c In practice, typical conditions used in the Haber process are a temperature of 500°C and a pressure of 200 atmospheres. Explain why these conditions are used rather than those that give the highest yield. [2]

**d** Write the equilibrium constant expression, K_c, for the production of ammonia. [1]

e i Suggest why this reaction is important for humanity. [1]

ii A chemist claims to have developed a new catalyst for the Haber process, which increases the yield of ammonia. State the catalyst normally used for the Haber process, and comment on the claim made by this chemist. [2]
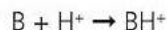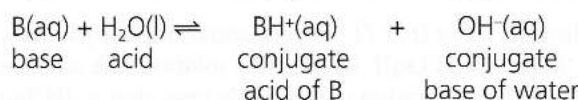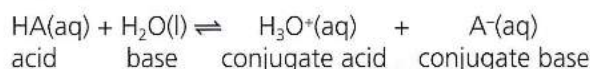
SUMMARY OF KNOWLEDGE

▦ Under the Brønsted–Lowry theory for defining acidity and basicity, an acid is defined as a proton donor, while a base is a proton acceptor. The reaction between an acid and a base involves the transfer of a proton (H^+).

▦ Under the Lewis theory for defining acidity and basicity, an acid is defined as an electron pair acceptor, while a base is an electron pair donor.

▦ Acids have polarized covalent bonds to hydrogen atoms which carry a partial positive charge. These bonds break in aqueous solution to release protons.

$$HA \rightarrow H^+ + A^-$$
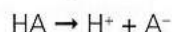
▦ Bases have a lone pair or lone pairs of electrons. These electrons receive the donated proton and a dative bond is formed.

$$B + H^+ \rightarrow BH^+$$

▦ Equations showing generalized acid and base behaviour:

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$
acid   base   conjugate acid   conjugate base

$$B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq)$$
base   acid   conjugate   conjugate
               acid of B   base of water

▦ A pair of species differing by a single proton is called a conjugate pair.

▦ A strong acid is almost completely ionized in aqueous solution. Strong acids include sulfuric, nitric and hydrochoric acids. Weak acids are only slightly ionized in solution. Organic acids, such as ethanoic acid, are all weak acids.

▦ A strong base is almost completely ionized in aqueous solution. Strong bases include potassium, sodium and barium hydroxides. Weak bases are only slightly ionized in solution. Aqueous ammonia solution and organic bases, such as the amines, are all weak bases.

▦ Strong acids have a higher electrical conductivity, lower pH and faster reactions with bases than weak acids (of the same concentration and temperature). Strong bases have a higher electrical conductivity and a higher pH than weak bases (of the same concentration and temperature).

▦ A strong acid is a good proton donor, hence its conjugate base is a poor proton acceptor. So strong acids will have weak conjugate bases, and conversely, weak acids will have strong conjugate bases.

▦ Acids react with reactive metals (those above hydrogen in the reactivity series) to release hydrogen and form a salt. Acids react with metal carbonates to form water, a salt and carbon dioxide. An acid reacts with a metal oxide to form a salt and water only. This reaction is termed neutralization.

▮ *Examination questions – a selection*

Paper 1 IB questions and IB style questions

Q1 Which one of the following descriptions defines a strong acid?
     A It is concentrated.
     B It does dissociate in water.
     C It absorbs water from the air.
     D It almost completely dissociated in water.

Q2 Which of the following represents the reaction between zinc powder and a dilute aqueous solution of sulfuric acid?
     A $Zn + 2H_2SO_4 \rightarrow 2ZnS + 2H_2O + 3O_2$
     B $4Zn + H_2SO_4 \rightarrow 4ZnO + H_2S$
     C $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
     D $Zn + H_2SO_4 \rightarrow ZnH_2 + SO_2 + O_2$

Acids ②

Q3 When the following 1.0 mol dm⁻³ solutions are
listed in increasing order of pH (lowest first), what
is the correct order?
- **A** $HNO_3 < H_2CO_3 < NH_3 < Ba(OH)_2$
- **B** $NH_3 < Ba(OH)_2 < H_2CO_3 < HNO_3$
- **C** $Ba(OH)_2 < H_2CO_3 < NH_3 < HNO_3$
- **D** $HNO_3 < H_2CO_3 < Ba(OH)_2 < NH_3$

Standard Level Paper 1, Nov 05, Q24

Q4 What is the pH of pure distilled water?
- **A** 0   **B** 4   **C** 7   **D** 6

Q5 Which one of the following represents the reaction
between calcium hydroxide and dilute hydrochloric
acid?
- **A** $Ca(OH)_2 + HCl \rightarrow CaOCl + H_2O$
- **B** $CaOH + HCl \rightarrow CaCl + H_2O$
- **C** $CaOH + 2HCl \rightarrow Cl_2 + CaOH_2$
- **D** $Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

Q6 The amino acid alanine has the structure:
$H_2N-CH(CH_3)-COOH$
Which of the following species represents its
conjugate acid?
- **A** $^+NH_3CH(CH_3)COOH$
- **B** $^+NH_3CH(CH_3)COOH_2^+$
- **C** $^+NH_3CH(CH_3)COO^-$
- **D** $NH_2CH(CH_3)COO^-$

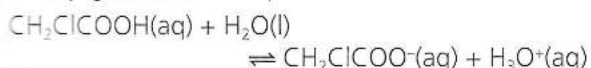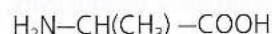**Q7** Four flasks labelled A, B, C and D contain
equal volumes of hydrochloric acid at different
concentrations. When equal volumes of 1 mol dm⁻³
sodium hydroxide are added to each flask the pH
values below are produced.
Which flask has the most concentrated acid?

Flask	A	B	C	D
pH	2	6	8	12

- **A** flask A   **B** flask B   **C** flask C   **D** flask D

Q8 In the equilibrium below, which species represents
a conjugate acid–base pair?
$$CH_2ClCOOH(aq) + H_2O(l)$$
$$\rightleftharpoons CH_2ClCOO^-(aq) + H_3O^+(aq)$$
- **A** $CH_2ClCOOH/H_2O$ **C** H_2O/CH_2ClCOO^-
- **B** CH_2ClCOO^-/H_3O^+ **D** H_2O/H_3O^+

Q9 Hydrogen chloride dissolved in water reacts with
magnesium. Hydrogen chloride dissolved in
ethanol does not react with magnesium. Which
statement accounts for this observation?
- **A** Ethanol accepts hydrogen ions (protons) and
water does not.
- **B** Water is a hydrogen ion/proton acceptor.
- **C** Magnesium is very soluble in ethanol but
insoluble in water.
- **D** Hydrogen chloride does not form ions in water.

Q10 Which of the following 1.00 mol dm⁻³ aqueous
solutions would have the highest pH value?
- **A** ammonia   **C** sulfuric acid
- **B** ethanoic acid   **D** sodium hydroxide

Q11 A dilute aqueous solution of benzenecarboxylic
acid (an organic acid) is a poor conductor of
electricity. Which of the following statements
accounts for this observation?
- **A** Benzenecarboxylic acid solution has a high
concentration of ions.
- **B** Benzenecarboxylic acid is only slightly
dissociated in water.
- **C** Benzenecarboxylic acid is completely
dissociated in water.
- **D** It is a strong acid.

Q12 Which is the correct description for an aqueous
solution with a pH of 9.5?
- **A** alkaline   **C** neutral
- **B** acidic   **D** amphoteric

Q13 Methanoic acid, HCOOH, is a stronger acid than
propanoic acid, CH_3CH_2COOH. Which one of the
statements about these acids is correct?
- **A** Propanoic acid is more dissociated in water
than methanoic acid.
- **B** Magnesium will react with methanoic acid but
not with propanoic acid.
- **C** A 1.0 mol dm⁻³ solution of methanoic acid
will turn blue litmus red, but a 1.0 mol dm⁻³
solution of propanoic acid will turn red litmus
blue.
- **D** The pH of a solution of 1 mol dm⁻³ propanoic
acid is higher than that of 1 mol dm⁻³
methanoic acid.

Q14 A decrease in the pH of an aqueous solution
corresponds to:
- **A** a decrease in the H⁺ concentration and an
increase in the OH⁻ concentration
- **B** an increase in the H⁺ concentration and a
decrease in the OH⁻ concentration
- **C** a decrease in the H⁺ concentration with no
change in the OH⁻ concentration
- **D** an increase in the OH⁻ concentration and an
increase in the H⁺ concentration

Q15 Calcium oxide is added to a lake to neutralize
the effects of acid rain. The pH value of the lake
water rises from 4 to 6. What is the change in
concentration of [H⁺(aq)] in the lake water?
- **A** an increase by a factor of 2
- **B** an increase by a factor of 100
- **C** a decrease by a factor of 2
- **D** a decrease by a factor of 100

Q3 When the following $1.0\,mol\,dm^{-3}$ solutions are listed in increasing order of pH (lowest first), what is the correct order?
  A $HNO_3 < H_2CO_3 < NH_3 < Ba(OH)_2$
  B $NH_3 < Ba(OH)_2 < H_2CO_3 < HNO_3$
  C $Ba(OH)_2 < H_2CO_3 < NH_3 < HNO_3$
  D $HNO_3 < H_2CO_3 < Ba(OH)_2 < NH_3$
      Standard Level Paper 1, Nov 05, Q24

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  A 0   **B** 4   **C** 7   **D** 6

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  A $Ca(OH)_2 + HCl \rightarrow CaOCl + H_2O$
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  D $Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

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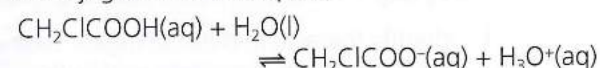 **C** $^+NH_3CH(CH_3)COO^-$
  D $NH_2CH(CH_3)COO^-$

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    Which flask has the most concentrated acid?

Flask	A	B	C	D
pH	2	6	8	12

  A flask A   **B** flask B   **C** flask C   **D** flask D

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$CH_2ClCOOH(aq) + H_2O(l)$
$\rightleftharpoons CH_2ClCOO^-(aq) + H_3O^+(aq)$

  A $CH_2ClCOOH/H_2O$ **C** H_2O/CH_2ClCOO^-
  B CH_2ClCOO^-/H_3O^+ **D** H_2O/H_3O^+

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  A Ethanol accepts hydrogen ions (protons) and water does not.
  B Water is a hydrogen ion/proton acceptor.
  C Magnesium is very soluble in ethanol but insoluble in water.
  D Hydrogen chloride does not form ions in water.

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  A ammonia   **C** sulfuric acid
  B ethanoic acid   **D** sodium hydroxide

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  A Benzenecarboxylic acid solution has a high concentration of ions.
  B Benzenecarboxylic acid is only slightly dissociated in water.
  C Benzenecarboxylic acid is completely dissociated in water.
  D It is a strong acid.

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  A alkaline   **C** neutral
  B acidic   **D** amphoteric

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  A Propanoic acid is more dissociated in water than methanoic acid.
  B Magnesium will react with methanoic acid but not with propanoic acid.
  C A $1.0\,mol\,dm^{-3}$ solution of methanoic acid will turn blue litmus red, but a $1.0\,mol\,dm^{-3}$ solution of propanoic acid will turn red litmus blue.
  D The pH of a solution of $1\,mol\,dm^{-3}$ propanoic acid is higher than that of $1\,mol\,dm^{-3}$ methanoic acid.

Q14 A decrease in the pH of an aqueous solution corresponds to:
  A a decrease in the H^+ concentration and an increase in the OH^- concentration
  B an increase in the H^+ concentration and a decrease in the OH^- concentration
  C a decrease in the H^+ concentration with no change in the OH^- concentration
  D an increase in the OH^- concentration and an increase in the H^+ concentration

Q15 Calcium oxide is added to a lake to neutralize the effects of acid rain. The pH value of the lake water rises from 4 to 6. What is the change in concentration of $[H^+(aq)]$ in the lake water?
  A an increase by a factor of 2
  B an increase by a factor of 100
  C a decrease by a factor of 2
  D a decrease by a factor of 100

Q16 Which chemical can behave as a Brønsted–Lowry base and as a Brønsted–Lowry acid?
A CO_3^{2-} **B** HSO_4^- **C** NO_3^-
D Such a species does not exist.

Q17 The pH of solution X is 1 and that of Y is 2. Which statement is correct about the hydrogen ion concentrations in the two solutions?
A $[H^+]$ in X is half that in Y.
B $[H^+]$ in X is twice that in Y.
C $[H^+]$ in X is one tenth of that in Y.
D $[H^+]$ in X is ten times that in Y.
Standard Level Paper 1, May 05, Q23

Q18 In which one of the following reactions does the nitric acid molecule act as a base?
A $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$
B $HNO_3 \rightarrow H^+ + NO_3^-$
C $HNO_3 + CH_3COOH \rightarrow CH_3COOH_2^+ + NO_3^-$
D $HNO_3 + 2H_2SO_4 \rightarrow NO_2^+ + 2HSO_4^- + H_3O^+$

Q19 Which one of the following species is not amphoteric (amphiprotic), that is, not capable of showing acidic and basic properties in aqueous solution?
A H_2O **C** NH_4^+
B $H_2PO_4^-$ **D** NH_2CHCH_3COOH

Q20 Which of the following would exactly neutralize $100\,cm^3$ of $1\,mol\,dm^{-3}$ sulfuric acid?
A $0.1\,mol$ of $Ba(OH)_2$ **C** $0.2\,mol$ of Na_2CO_3
B $0.1\,mol$ of KOH **D** $0.1\,mol$ of NH_3

Q21 Which of the following compounds containing hydrogen acts as an acid in aqueous solution?
A hydrogen chloride **C** methane
B ammonia **D** ethene

Q22 Which is not a strong acid?
A nitric acid **C** carbonic acid
B sulfuric acid **D** hydrochloric acid
Standard Level Paper 1, Nov 06, Q23

Q23 Which equation correctly describes phosphoric(v) acid behaving as a monoprotic acid in aqueous solution?
A $H_3PO_4(aq) \rightarrow H^+(aq) + H_2PO_4^-(aq)$
B $H_2PO_4^-(aq) \rightarrow H^+(aq) + HPO_4^{2-}(aq)$
C $H_3PO_4(aq) \rightarrow 3H^+(aq) + PO_4^{3-}(aq)$
D $H_3PO_4(aq) \rightarrow 2H^+(aq) + HPO_4^{2-}(aq)$

Q24 According to the Lewis theory, a base:
A is a proton acceptor
B is a proton donor
C makes available a share in a pair of electrons
D accepts a share in a pair of electrons

Q25 Ammonia molecules in aqueous solution can be considered as:
A a Lewis acid (only)
B a Lewis base (only)
C a Brønsted–Lowry base (only)
D both a Brønsted–Lowry base and a Lewis base
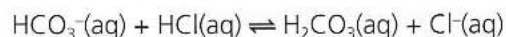
Paper 2 IB questions and IB style questions

Q1 **a** The pH values of solutions of three organic acids of the same concentration were measured.
acid X   pH = 5
acid Y   pH = 2
acid Z   pH = 3

i Identify which solution is the least acidic. [1]
ii Deduce how the $[H^+]$ values compare in solutions of acids Y and Z. [1]
iii Arrange the solutions of the three acids in decreasing order of electrical conductivity, starting with the greatest conductivity, giving a reason for your choice. [2]

Q2 Carbonic acid (H_2CO_3) is described as a weak acid and hydrochloric acid (HCl) is described as a strong acid.
a Explain, with the help of equations, what is meant by strong and weak acid using the above acids as examples. [4]
b Outline **two** ways, other than using pH, in which you could distinguish between carbonic acid and hydrochloric acid of the same concentration. [2]
c A solution of hydrochloric acid, HCl(aq), has a pH of 1 and a solution of carbonic acid, H_2CO_3(aq), has a pH of 5. Determine the ratio of the hydrogen ion concentrations in these solutions. [2]
d The relative strengths of the two acids can be illustrated by the following equation:

$$HCO_3^-(aq) + HCl(aq) \rightleftharpoons H_2CO_3(aq) + Cl^-(aq)$$

i Identify the acid and its conjugate base and the base and its conjugate acid in the above equation. [2]
ii Name the theory that is illustrated in **d i**. [1]
Standard Level Paper 2, Nov 02, Q5

Rodox ①

■ Examination questions – a selection

Paper 1 IB questions and IB style questions

Q1 When $MnO_4^-(aq)$ reacts in an acidic solution it produces:
A Mn^{2+} **B** Mn^{3+} **C** MnO_4^{2-} **D** MnO_2

Q2 In the reaction

$$MnO_2(s) + 4HCl(aq) \rightarrow Cl_2(g) + MnCl_2(aq) + 2H_2O(l)$$

A HCl is the oxidizing agent.
B Cl_2 is the oxidation product.
C MnO_2 is the reducing agent.
D H_2O is the reduction product.

Q3 The following reaction occurs in acid solution:
$$\_H^+(aq) + \_NO_3^-(aq) + I^-(aq)$$
$$\rightarrow IO_3^-(aq) + \_NO_2(g) + \_H_2O(l)$$

The equation is not balanced. What is the coefficient of NO_3^- in the balanced equation?
A 4   **B** 2   **C** 5   **D** 6

Q4 Which of the following reactions involves neither oxidation nor reduction?
A $Ag^+ + Br^- \rightarrow AgBr$
B $2H_2S + SO_2 \rightarrow 3S + 2H_2O$
C $2[Ag(NH_3)_2]^+ + Cu \rightarrow Cu[(NH_3)_4]^{2+} + 2Ag$
D $2Al + 2OH^- + 6H_2O \rightarrow 2Al(OH)_4^- + 3H_2$

Q5 Which one of the following equations represents the half-equation (or half-reaction) that occurs at the anode during the electrolysis of molten potassium iodide?
A $K^+ + e^- \rightarrow K$
B $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$
C $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
D $2I^- \rightarrow I_2 + 2e^-$

Q6 All of the following equations represent oxidation–reduction reactions except:
A $2C_3H_7OH + 9O_2 \rightarrow 6CO_2 + 8H_2O$
B $Ni + 4CO \rightarrow Ni(CO)_4$
C $2Na + 2H_2O \rightarrow 2NaOH + H_2$
D $Cl_2 + 2NaBr \rightarrow Br_2 + 2NaCl$

Q7 The following information is given about reactions involving the metals X, Y and Z and solutions of their sulfates.

$X(s) + YSO_4(aq) \rightarrow$ no reaction
$Z(s) + YSO_4(aq) \rightarrow Y(s) + ZSO_4(aq)$

When the metals are listed in decreasing order of reactivity (most reactive first), what is the correct order?
A $Z > Y > X$ **C** $Y > X > Z$
B $X > Y > Z$ **D** $Y > Z > X$
Standard Level Paper 1, Nov 05, Q27

Q8 In the reaction

$$2MnO_2 + 4KOH + O_2 + Cl_2 \rightarrow 2KMnO_4 + 2KCl + 2H_2O$$

the oxidizing agent(s) is/are:
A $KMnO_4$ only   **C** MnO_2 and O_2
B MnO_2 only   **D** O_2 and Cl_2

Q9 Which one of the following represents an oxidation–reduction reaction?
A $I_2(s) + 2OH^-(aq) \rightarrow I^-(aq) + OI^-(aq) + H_2O(l)$
B $PO_4^{3-}(aq) + H_2O(l) \rightarrow HPO_4^{2-}(aq) + OH^-(aq)$
C $SO_3(g) + 2H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$
D $Cu^{2+}(aq) + H_2S(aq) \rightarrow CuS(s) + 2H^+(aq)$

Q10 Which of the following reactions involves neither oxidation nor reduction?
A $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$
B $2CrO_4^{2-}(aq) + 2H^+(aq) \rightarrow Cr_2O_7^{2-}(aq) + H_2O(l)$
C $C_3H_6(g) + H_2(g) \rightarrow C_3H_8(g)$
D $NH_4NO_2(s) \rightarrow N_2(g) + 2H_2O(g)$

Q11 All of the following would be expected to function as both oxidizing and reducing agents except:
A NO_2 **B** Cl^- **C** ClO^- **D** S

Q12 Magnesium is a more reactive metal than copper. Which is the strongest oxidizing agent?
A Mg   **B** Mg^{2+} **C** Cu   **D** Cu^{2+}
Standard Level Paper 1, Nov 03, Q26

Q13 Bromide ions are oxidized to bromine by all of the following except:
A $K_2Cr_2O_7/H^+$ **C** I_2
B Cl_2 **D** $KMnO_4$

Q14 When a direct current of electricity is conducted by an aqueous solution of an electrolyte, which one of the following statements is false?
A The movements of ions accounts for the current flow through the solution.
B During electrolysis, the solution remains electrically neutral.
C Electrons flow from the current source toward the solution at one electrode, and an equal number of electrons flows away from the solution at the other electrode.
D The number of positive ions moving toward one electrode is always equal to the number of negative ions moving toward the other electrode.

Q15 In acid solution, manganate(vii) ions, $MnO_4^-(aq)$, undergo reduction to manganese(ii) ions, $Mn^{2+}(aq)$. What amount of $MnO_4^-(aq)$ is required to convert 5.36×10^{-3} moles of the ion $Y^{2+}(aq)$ to $YO_3^-(aq)$?
A 1.07×10^{-3} mol   **C** 5.36×10^{-3} mol
B 3.22×10^{-3} mol   **D** 8.93×10^{-3} mol

Q16 What is the oxidation number of chromium in $Cr_2O_7^{2-}$?
A +7   **B** +6   **C** −6   **D** −2

Q17 Which statement is correct for the electrolysis of molten sodium chloride?
A Sodium ions move toward the positive electrode.
B A gas is produced at the negative electrode.
C Only electrons move in the electrolyte.
D Both sodium and chloride ions move toward electrodes.

Q18 In which of the following does the metal undergo a change in oxidation state?
  I $2MnO_4^{2-} + F_2 \rightarrow 2MnO_4^- + 2F^-$
  II $2CrO_4^{2-} + 2H^+ \rightarrow Cr_2O_7^{2-} + H_2O$
  III $[Fe(H_2O)_6]^{2+} + 6CN^- \rightarrow [Fe(CN)_6]^{3-} + 6H_2O$
A I only   **C** I and II only
B II only   **D** I and III only

Q19 Which one of the following does not represent a redox reaction?
A $Cu(NO_3)_2(aq) + Na_2S(aq) \rightarrow CuS(s) + 2NaNO_3(aq)$
B $2Na(s) + I_2(s) \rightarrow 2NaI(s)$
C $KH(s) + H_2O(l) \rightarrow KOH(aq) + H_2(g)$
D $H_2SO_4(aq) + 2HBr(g) \rightarrow SO_2(g) + 2H_2O(l) + Br_2(l)$

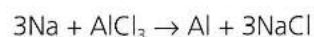**Q20** In which one of the following species does chlorine exhibit the highest oxidation number?
A Cl_2O **B** Cl_2 **C** $HClO_3$ **D** PCl_5

Paper 2 IB questions and IB style questions

Q1 a Use these equations, which refer to aqueous solutions, to answer the questions that follow:

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

$$Cu(s) + 2Au^+(aq) \rightarrow Cu^{2+}(aq) + 2Au(s)$$

$$Mg(s) + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)$$

(Au represents gold, which is below silver in the reactivity series.)
  i List the metals above in order of **decreasing** reactivity. [1]
  ii Define oxidation, in electronic terms, using **one** example from above. [2]
  iii Define reduction, in terms of oxidation number, using **one** example from above. [2]
  iv State and explain which is the **strongest reducing agent** in the examples above. [2]
  v State and explain which is the **strongest oxidizing agent** in the examples above. [2]
  vi Deduce whether a gold coin will react with aqueous magnesium nitrate. [2]
b Sketch a diagram of a cell used to electrolyse a molten salt. Label the essential components. [4]

c Describe how electrode reactions occur in an electrolytic cell and state the products at each electrode when molten copper(II) iodide is electrolysed. [4]

Q2 a Electrolysis can be used to obtain fluorine from molten potassium fluoride. Write an equation for the reaction occurring at each electrode and describe the two different ways in which electricity is conducted when the cell is in operation. [4]
b In one experiment involving the electrolysis of molten potassium fluoride, 0.1 mol of fluorine was formed. Deduce, giving a reason, the amount of potassium formed at the same time. [2]
c Sodium will displace aluminium from its chloride on heating:

$$3Na + AlCl_3 \rightarrow Al + 3NaCl$$

  i Explain, by reference to electrons, why the reaction is referred to as a redox reaction. [2]
  ii Deduce the oxidation numbers of sodium and aluminium in the reactants and products. [2]
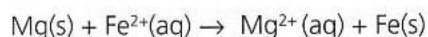
Q3 A voltaic cell is set up with a silver reference electrode and a series of other metals immersed in an electrolyte. The cell voltages were recorded in the table below.

Metal	Cell voltage/V
Aluminium	2.47
Zinc	1.55
Iron	1.19
Copper	0.46
Silver	0.00

a What is the relationship between the voltage of the cell and the position of the metal in the reactivity series? [1]
b Is the metal acting as the negative or positive electrode?
Explain your answer. [2]
c Construct the half-cell equations for a voltaic cell in which the metal is zinc and the electrolyte is silver nitrate. [2]

Examination questions – a selection

Paper 1 IB questions and IB style questions

Q1 How many structural isomers are possible with the molecular formula C_6H_{14}?
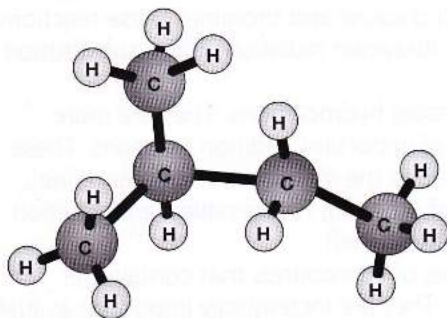**A** 4   **B** 6   **C** 5   **D** 7

Q2 Which compound is a member of the aldehyde homologous series?
A $CH_3CH_2COCH_3$ **C** $CH_3CH_2CH_2OH$
B CH_3CH_2CHO **D** CH_3CH_2COOH

Q3 Which compound is a member of the same homologous series as 1-chloropropane?
A 1-chloropropene   **C** 1-bromopropane
B 1,2-dichloropropane   **D** 1-chlorobutane

Q4 The following is a three-dimensional representation of an organic molecule.

Which statement is correct?
A The correct IUPAC name of the molecule is 2-methylpentane.
B All the bond angles will be approximately 90°.
C One isomer of this molecule is pentane.
D The boiling point of this compound would be higher than that of pentane.
      Standard Level Paper 1, Specimen 09, Q26

Q5 What is the organic product of the reaction between ethanol and ethanoic acid in the presence of concentrated sulfuric acid?
A CH_3CHO **C** $CH_3CH_2COOCH_3$
B $CH_3COOCH_2CH_3$ **D** CH_3COOCH_3

Q6 Which formulas represent butane or its isomer?
  I $CH_3(CH_2)_2CH_3$
 II $CH_3CH(CH_3)CH_3$
III $(CH_3)_3CH$
A I, II and III   **C** I and II only
B I and III only   **D** II and III only

Q7 A *gaseous* hydrocarbon, **X**, decolorizes aqueous bromine. Which one of the following molecular formulas could be **X**?
A C_2H_4 **B** C_6H_{14} **C** $C_{13}H_{28}$ **D** $C_{10}H_{20}$

Q8 What is the function of sunlight during the reaction between chlorine and methane?
A to dissociate the chlorine molecules into atoms
B to dissociate the chlorine molecules into ions
C to increase the temperature of the mixture
D to break C–H bonds in the methane molecules

Q9 Which one of the following is the best method of distinguishing between an alkane and an alkene?
A test with universal indicator paper
B burn the gases in excess oxygen
C test their solubility in water
D add bromine water

Q10 Which equation represents the combustion of methane in excess oxygen?
A $CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(l)$
B $CH_4(g) + O_2(g) \rightarrow CO_2(g) + 2H_2(g)$
C $2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(l)$
D $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

Q11 Which of the following compounds is not formed by the reaction between excess chlorine and methane (in the presence of sunlight)?
A C_2H_4 **B** C_2H_6 **C** $CHCl_3$ **D** CCl_4

Q12 What type of alcohol is 2-methylpropan-2-ol?
A an unsaturated alcohol   **C** a primary alcohol
B a tertiary alcohol   **D** a secondary alcohol

Q13 What is the IUPAC name for $(CH_3)_2C(OH)CH_2CH_3$?
A 2-methylbutan-2-ol   **C** 3-methylbutan-3-ol
B 4-methylbutan-3-ol   **D** pentan-3-ol

Q14 How do the bond angles and bond lengths in ethane and ethene compare?

	H–C–H bond angle in ethane	C–C bond length in ethane
A	larger	longer
B	smaller	longer
C	smaller	shorter
D	larger	shorter

Q15 Which compound is an ester?
A CH_3CH_2COOH **C** C_3H_7CHO
B $C_2H_5OC_2H_5$ **D** $HCOOCH_3$

Q16 The oxidation of propan-2-ol, $CH_3CH(OH)CH_3$, by sodium dichromate(VI) leads to the formation of:
A propanone (CH_3COCH_3)
B propan-1-ol $(CH_3CH_2CH_2OH)$
C propanal (CH_3CH_2CHO)
D propanoic acid $(CH_3CH_2CO_2H)$

:17 When the compounds below are listed in order of **decreasing** boiling point (highest to lowest) what is the correct order?

1. ethane   2. choroethane
3. ethanol   4. ethanoic acid

A 2, 1, 3, 4   **C** 3, 4, 1, 2
B 4, 3, 2, 1   **D** 4, 3, 1, 2

Q18 What is the name of the compound whose condensed structural formula is $CH_3CH_2CH_2COOCH_3$?

A butyl methanoate   **C** methyl propanoate
B methyl butanoate   **D** pentanone

Q19 Which of the following descriptions can be correctly applied to the homologous series of alkanes?

  I Members of the series have the general formula C_nH_{2n+2}.
  II Members of the series have similar chemical properties.
  III Members of the series are isomers of each other.

A I only   **C** I and II only
B II only   **D** I and III only

Q20 The compound which is expected to have the lowest boiling point at a pressure of one atmosphere is:

A $CH_3CH_2CH_2F$ **C** CH_3CH_2COOH
B $CH_3CH_2CH_2OH$ **D** $CH_3CH_2CH_2NH_2$

## Paper 2 IB questions and IB style questions

Q1 An alkane has the percentage composition 84.5% carbon and 15.5% hydrogen by mass.
  a Calculate the empirical formula of the alkane. [2]
  b The molecular mass of the alkane was found to be 142 using a mass spectrometer. What is the molecular formula? [2]
  c **i** The hydrocarbon can be used as a fuel. Write the balanced equation for the complete combustion of this alkane in oxygen. [2]
    ii Write a balanced equation for the incomplete combustion of this alkane in a limited supply of oxygen. [2]
  d When a hydrocarbon is cracked, it is broken into smaller molecules. Complete the following cracking reactions:
    i $C_8H_{18} \rightarrow C_4H_8 +$ _____ [1]
    ii $C_{13}H_{28} \rightarrow C_4H_{10} + C_4H_8 +$ _____ [1]

Q2 Ethene, propene and but-2-ene are members of the alkene homologous series.
  a Describe **three** features of members of a homologous series. [3]
  b State and explain which compound has the highest boiling point. [3]

  c Draw the structural formula and give the name of an alkene containing five carbon atoms. [2]
  d Write an equation for the reaction between but-2-ene and hydrogen bromide, showing the structure of the organic product. State the type of reaction occurring. [3]
  e Propene can be converted to propanoic acid in three steps:

    step 1   step 2   step 3
  propene → propan-1-ol → propanal → propanoic acid
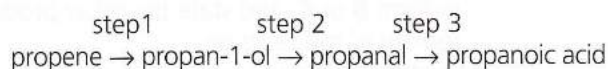
    State the type of reaction occurring in steps 2 and 3 and the reagents needed. Describe how the conditions of the reaction can be altered to obtain the maximum amount of propanal, and in a separate experiment, to obtain the maximum amount of propanoic acid. [5]
  f Identify the strongest type of intermolecular force present in each of the compounds propan-1-ol, propanal and propanoic acid. List these compounds in decreasing order of boiling point. [4]

Standard Level Paper 2, Nov 05, Q7

Q3 **a** An organic compound, **A**, containing only the elements carbon, hydrogen and oxygen was analysed.
    i **A** was found to contain 54.5% C and 9.1% H by mass, the remainder being oxygen. Determine the empirical formula of the compound. [3]
    ii The molecular mass of **A** is 88. What is the molecular formula of **A**? [2]
  b An organic compound **X** contains 40.00% carbon, 6.72% hydrogen and 53.28% oxygen by mass.
    i Determine the empirical formula of compound **X**. [2]
    ii Compound **X** has a relative molecular mass of 60.0. Deduce its molecular formula. [2]

Q4 **a** Give the structural formulas for the isomers of molecular formula C_4H_{10} and name each isomer. [4]
  b Several compounds have the molecular formula $C_3H_6O_2$. Three of them, **A**, **B** and **C**, have the following properties:
   **A** is soluble in water and is acidic.
    B and **C** are neutral and do not react with bromine or organic acids.
    Give a structural formula for each of these compounds and name them. [6]
  c **i** Explain the solubility and acidity of **A** in water. [2]
    ii Write an equation for the reaction of **A** with sodium hydroxide solution. [1]

iii Explain why **B** and **C** do not react with bromine. [1]

d State and explain which one of **A**, **B** or **C** has the highest boiling point. [2]

e i Name the class of compounds to which **B** and **C** belong and state a use of this class of compounds. [2]

ii Name the **two** classes of compounds used to form **B** or **C**, and state the other product formed in this reaction. [3]

f Suggest the structural formula of an isomer of $C_3H_6O_2$ which does react rapidly with bromine. Name this type of reaction, and describe an observation that can be made during the reaction. [3]

Q5 a i List **three** characteristics of a homologous series. [3]

ii Draw the **four** different structural isomers with the formula C_4H_9OH that are alcohols. [4]

b i Ethanoic acid reacts with ethanol in the presence of concentrated sulfuric acid and heat. Identify the type of reaction that takes place. Write an equation for the reaction, name the organic product formed and draw its structure. [4]

ii State and explain the role of sulfuric acid in this reaction. [2]

iii State **one** major commercial use of the organic product from this type of reaction. [1]

c Two compounds are shown below.

$$HCOOCH_2CH_3 \quad \text{and} \quad HCOOCHCH_2$$
$$\textbf{I} \qquad\qquad\qquad \textbf{II}$$

i State and explain which of these two compounds can react readily with bromine. [2]

ii Compound **II** can form polymers. State the type of polymerization compound **II** can take part in, and draw the structure of the repeating unit of the polymer. [2]

Q6 a The following is a computer-generated representation of the molecule methyl 2-hydroxyl benzoate, better known as oil of wintergreen.
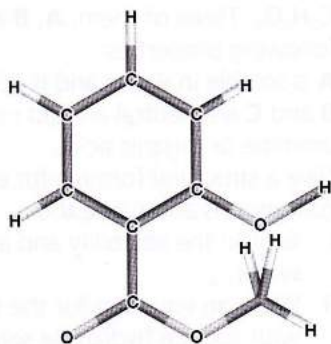
i Deduce the empirical formula of methyl 2-hydroxybenzoate and draw the full structural formula, including any multiple bonds that may be present. The computer-generated representation shown does not distinguish between single and multiple bonds. [2]

ii In this representation, two of the carbon–oxygen bond lengths shown are 0.1424 nm and 0.1373 nm, respectively. Explain why these are different and predict the carbon–oxygen bond length in carbon dioxide. [2]

iii Name two of the functional groups present in the molecule. [2]

b i State and explain the trend in the boiling points of the first six straight-chain alkanes. [2]

ii Write an equation for the reaction between methane and chlorine to form chloromethane. Explain this reaction in terms of a free-radical mechanism. [5]

c i Identify the formulas of the organic products, **A–E**, formed in the reactions **I–IV**:

I $CH_3(CH_2)_8OH + K_2Cr_2O_7 \xrightarrow{H^+} A \xrightarrow{H^+} B$

II $(CH_3)_3CBr + NaOH \rightarrow C$

III $(CH_3)_2CHOH + K_2Cr_2O_7 \xrightarrow{H^+} D$

IV $H_2C=CH_2 + Br_2 \rightarrow E$ [5]

ii $H_2C=CH_2$ can react to form a polymer. Name this **type** of polymer and draw the structural formula of a section of this polymer consisting of three repeating units. [2]

Standard Level Specimen Paper 2, 2009, Q8

Q7 a Name the following alcohols:

i $CH_3CH(OH)CH_3$

ii $CH_3CH_2CH_2OH$

iii $CH_3CH_2C(OH)(CH_3)CH_3$

iv $CH_2(OH)CH_2(OH)$ [4]

b For the four alcohols listed in **a**, state whether they are primary, secondary or tertiary alcohols. [4]

c If the alcohols in **a** are oxidized using acidified sodium dichromate(vi) under reflux, give the name and condensed structural formula of the organic product. [7]

- Exact numbers can be treated as if they have an infinite number of significant figures.
- When doing more than one calculation, do not round numbers until the end.
- Graphs can be used to display or interpret raw or processed data.
- A straight-line graph can be represented by $y = mx + c$, where m represents the gradient and c is the y-intercept (the value where the graph cuts the y-axis).
- Line graphs are appropriate for continuous data; histograms are appropriate for discontinuous data.
- Pie charts can be used to show the relative importance of differing components.
- Graphs can be used to obtain information from the gradient; a value between points on the graph can be measured by interpolation and a value outside the measured range by extrapolation.

Examination questions – a selection

Paper 1 IB questions and IB style questions

Q1 Which one of the following numbers is given to four significant figures?

A 0.00040 C 4.000
B 0.0040 D 4000

Q2 How many significant figures are there in this measured quantity?

0.040930

A 7 B 5 C 4 D 3

Q3 How many significant figures are there in this measured quantity?

5.010×10^3

A 4 B 2 C 5 D 6 E 3

Q4 Perform the indicated operation and give the answer to the appropriate accuracy.

48.2 m + 3.87 m + 48.4394 m
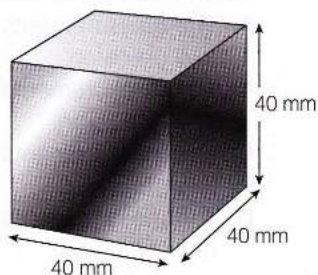
A 100.5094 m C 100.51 m
B 100.5 m D 101 m

Q5 Perform the indicated operation and give the answer to the appropriate accuracy.

451 g – 15.46 g

A 436 g C 435.5 g
B 435.54 g D 4.4×10^2 g

Q6 The dimensions of a cube are measured. The measured length of each side is 40 mm ± 0.1 mm. What is the approximate uncertainty in the value of its volume?

40 mm
40 mm
40 mm

A 1/64% B 1/4% C 4/10% D 0.75%

Q7 What is the percentage random uncertainty in a mass of pure water measured to the nearest microgram (µg) in a kilogram (kg)? ($1000000 \,µg = 1\,g$)

A 10^{-4} B 10^{-6} C 10^{-7} D 10^{-12}

Q8 An object of mass 2.000 kg is placed on four different balances (**A**, **B**, **C** and **D**) and for each balance the reading is recorded five times. The table shows the values obtained with the averages.

Which balance has the smallest systematic error *but is imprecise*?

Balance	1	2	3	4	5	Average/kg
A	2.000	2.000	2.002	2.001	2.002	2.001
B	2.011	1.999	2.001	1.989	1.995	1.999
C	2.012	2.013	2.012	2.014	2.014	2.013
D	1.993	1.987	2.002	2.000	1.983	1.993

Q9 An IB chemistry student records a series of precise measurements from which the student calculates the enthalpy of combustion of a hydrocarbon as 327.66 kJ mol^{-1}. The student estimates that the result is accurate to ±3%.

Which of the following gives the student's result expressed to the appropriate number of significant figures?

A 300 kJ mol^{-1} C 330 kJ mol^{-1}
B 328 kJ mol^{-1} D 327.7 kJ mol^{-1}

Q10 Which experimental technique reduces the systematic error in the investigation?

A Adjusting an electronic balance to remove its zero error before weighing a chemical.
B Repeating a titration a number of times and calculating an average titre.
C Using larger amounts of an indicator during a titration.
D Using a magnifying glass to enlarge the meniscus in a burette.

Q11 A titration is carried out by a large number of students in class and the number N of measurements giving a titre volume x is plotted against x. The true value of the titre volume is x_0.

Which graph best represents precise measurements with poor accuracy?
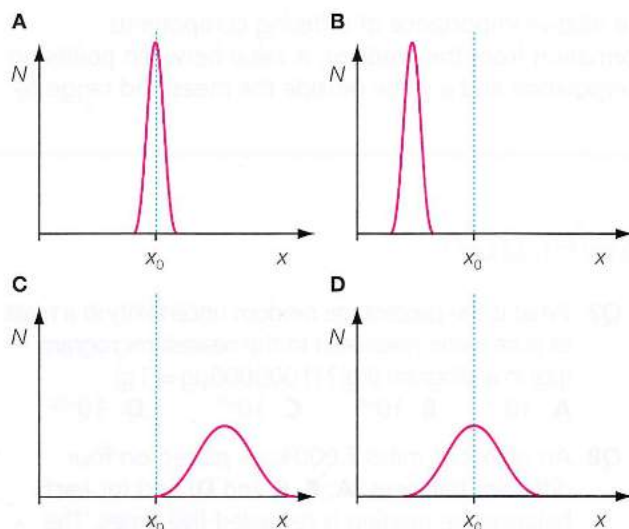
Q12 Which of the following four recorded measurements has the smallest percentage error?
- **A** $9.99\,cm \pm 0.005\,cm$
- **B** $4.44\,cm \pm 0.005\,cm$
- **C** $1.11\,cm \pm 0.005\,cm$
- **D** $5.55\,cm \pm 0.005\,cm$

Q13 The first mass reading is a weighing bottle and sodium hydroxide. The second mass reading is the empty weighing bottle.

What is the mass of the sodium hydroxide, and what is the random uncertainty in the value?
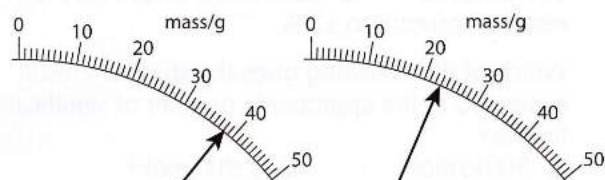
- **A** $15g \pm 2g$
- **B** $15g \pm 0.5g$
- **C** $18g \pm 2g$
- **D** $18g \pm 0.5g$

Q14 What are the greatest and smallest possible values for the following length:

34.3 cm?
- **A** 34.35 cm and 34.25 cm
- **B** 34.2 cm and 34.4 cm
- **C** 34.6 cm and 34.0 cm
- **D** 35.3 cm and 33.3 cm

Q15 What is the following measurement with the possible random uncertainty shown as a percentage instead of an absolute error:

$6.25\,cm^3 \pm 0.005\,cm^3$?
- **A** $6.25\,cm^3 \pm 0.4\%$
- **B** $6.25\,cm^3 \pm 0.8\%$
- **C** $6.25\,cm^3 \pm 0.04\%$
- **D** $6.25\,cm^3 \pm 0.08\%$

Q16 When comparing systematic errors and random uncertainties during an investigation, the following pairs of properties of errors in an experimental measurement may be considered:
- **I** error can possibly be removed
- **II** error cannot possibly be removed
- **III** error is of constant sign and size
- **IV** error is of varying sign and size
- **V** error will be reduced by averaging repeated measurements
- **VI** error will not be reduced by averaging repeated measurements

Which properties apply to random uncertainties?
- **A** I, II, III
- **B** I, IV, VI
- **C** II, IV, V
- **D** II, III, V

Q17 The diagram shows a set of experimental data points, X, determined when one experimental measurement was repeated four times. The centre of the diagram represents the ideal value calculated from theory. What statement is correct about these measurements?
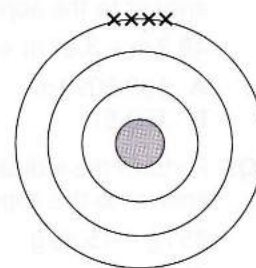
- **A** The measurements involve low accuracy and low precision.
- **B** The measurements involve low accuracy and high precision.
- **C** The measurements involve high accuracy and low precision.
- **D** The measurements involve high accuracy and high precision.

Q18 In a school laboratory, which of the pieces of apparatus listed below has the greatest random uncertainty in a measurement?

  A A $50\,cm^3$ burette when used to measure $25\,cm^3$ of ethanol.

  B A $25\,cm^3$ pipette when used to measure $25\,cm^3$ of ethanol.

  C A $50\,cm^3$ measuring cylinder when used to measure $25\,cm^3$ of ethanol.

  D An analytical balance (4 decimal places) when used to weigh $25\,cm^3$ of ethanol.

Q19 Perform the following density calculation to the correct number of significant figures:

$$\frac{1.00\,g}{3.00\,cm^3}$$

  A $0.333\,g\,cm^{-3}$ **C** $0.3\,g\,cm^{-3}$

  B $0.3333\,g\,cm^{-3}$ **D** $0.33\,g\,cm^{-3}$

Q20 An experiment to determine the molar mass of solid hydrated iron(II) sulfate, $FeSO_4.5H_2O$ gave a result of $258\,g\,mol^{-1}$.

  What is the experimental error?

  A 0.07% **C** 0.7%

  B 7% **D** 77%

Paper 2 IB questions and IB style questions

Q1 One method a chemist can use to investigate acid-base reactions is a titration. A pH titration is performed by adding small, accurate amounts of sodium hydroxide solution to hydrochloric acid of unknown concentration. The pH is recorded and is plotted versus the volume of base added to the acid solution.

  State how the following would affect the calculated concentration of acid:

  a i The burette is dirty and drops of sodium hydroxide cling to the side walls of the burette as it is drained. [2]

  ii The burette is not rinsed with sodium hydroxide prior to filling. [2]

  iii The burette tip is not filled at the start of the titration. [1]

  iv The sodium hydroxide solution is added too rapidly in the region of rapid pH change. [1]

  b i It is suspected that the pH meter consistently gives a reading 0.5 units above the actual value. What type of error is this? [1]

  ii How would you verify this error? [1]

Q2 The graph below shows how the mass of copper deposited during electrolysis varies with time.

Current: 4 amperes
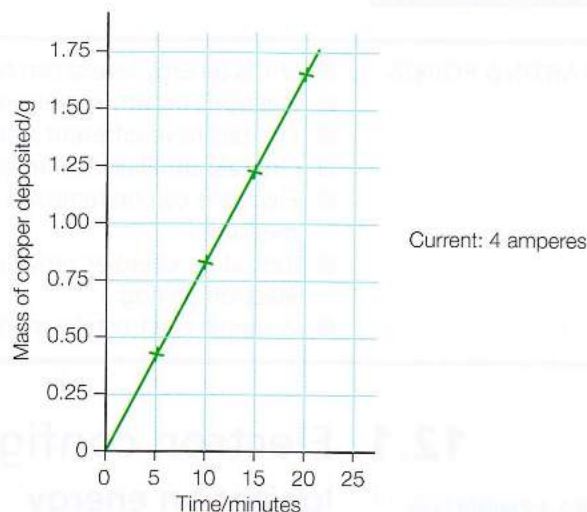
  a Identify the dependent, independent and controlled variables. [3]

  b State the relationship between the independent and dependent variables. [1]

  c Calculate the rate (in $g\,min^{-1}$) of copper deposition (to two decimal places). [2]

  d State **two** useful mathematical operations that could be performed on the graph. [2]

Q3 The length of a piece of paper was measured as $298\,mm \pm 1\,mm$. Its width was measured as $210\,mm \pm 1\,mm$.

  a Calculate the percentage random uncertainty in its length. [1]

  b Calculate the percentage random uncertainty in its width. [1]

  c Calculate the area of the piece of paper and its random uncertainty. [2]

Q4 A rectangular block with a density of $2.50\,g\,cm^{-3}$ has the following dimensions and random uncertainties. Calculate the uncertainty in the density. [2]

  mass = $25.0\,g \pm 0.1\,g$

  length = $5.00\,cm \pm 0.01\,cm$

  width = $2.00\,cm \pm 0.01\,cm$

  height = $1.00\,cm \pm 0.01\,cm$

Q5 A weighing bottle plus a sample of a pure liquid has a combined mass of $120.2\,g$. The weighing bottle has a mass of $119.0\,g$. The density of the liquid is $2.05\,g\,cm^{-3}$. Calculate the volume of the liquid to the correct number of significant digits. [2]