1atm = 101325 Pa = 760 mmHg = 760 torr, ,	1 dm^3 = 1 L	The molar volume of any gas at s.t.p. is 22.4 dm^3·mol^{-1}	$pV = nRT$
density of H$_2$O = 1.0 g·cm^{-3} at 25°C	R= 8.31 J mol^{-1} K^{-1}= 0.0821 L·atm·mol^{-1}·K^{-1}	Avogadro's constant = 6.02 × 10^{23} mol^{-1}.	$\Delta H = -mc\Delta T$

SECTION I: Multiple Choice Questions

Select the **SINGLE** best alternative in each of the following cases and mark the appropriate letter on the separate answer sheet by shading in the appropriate square with soft pencil. (1 mark each = **30** marks)

1) Which one of the following species in its ground state has three unpaired electrons in the 2p sub-shell?
   - A   a hydrogen atom.
   - B   an oxygen atom.
   - C   a boron atom.
   - D   a nitrogen atom.
   - E   a phosphorus atom.

2) A real gas most closely approaches the behavior of an ideal gas at
   - A   0.50 atm and 500K
   - B   1 atm and 273K
   - C   1 atm and 298K
   - D   15 atm and 500K
   - E   15 atm and 200K

3) Which is the correct ground-state electron configuration for iron?
   - A   [Ar] 3$d^2$4s^2
   - B   [Ar] 3$d^3$4s^2
   - C   [Ar] 3$d^4$4s^2
   - D   [Ar] 3$d^5$4s^2
   - E   [Ar] 3$d^6$4s^2

4) For the reaction below the $\Delta H° = +128.1$ kJ.
   $$CH_3OH (l) \rightarrow CO (g) + 2H_2 (g)$$
   How many kJ of heat are consumed when 0.200 mol of H$_2$ (g) is formed as shown in the equation?
   - A   12.8
   - B   25.6
   - C   128
   - D   256
   - E   6.4

5) Which of the following is an isoelectronic series?
   - A   B^{5-}, Si^{4-}
   - B   F$^-$, Cl$^-$
   - C   S, Cl
   - D   Si^{2-}, P^{2-}
   - E   O^{2-}, F$^-$

6) 2 mol of benzoic acid, C$_6$H$_5$COOH, contains
   - A   2 mol of oxygen atoms
   - B   6 mol of hydrogen atoms
   - C   10 mol of hydrogen atoms
   - D   12 mol of carbon atoms
   - E   14 mol of carbon atoms

7) How many grams of hydrogen are in 48 g of CH$_4$O?
   - A   2
   - B   4
   - C   6
   - D   8
   - E   10

8) Electronegativity increases in which of the following orders?
   - A   Cl < S < P
   - B   I < Br < Se
   - C   C < N < O
   - D   Cl < Br < I
   - E   F < S < Se

9) Which gas would effuse more rapidly under the same conditions of temperature and pressure?
   - A   CO$_2$ (RMM = 44)
   - B   SO$_2$ (RMM = 64)
   - C   C$_2$H$_2$ (RMM = 26)
   - D   UF$_6$ (RMM = 352)
   - E   Ne   (RMM = 10)

10) Which one of the following molecules has a dipole moment?
   - A   O$_2$
   - B   N$_2$
   - C   Cl$_2$
   - D   F$_2$
   - E   IF

11) 50 g of an ionic compound XY$_2$ (RFM=100) were dissolved in water and made up to 1.0 dm^3. What is the concentration of the solution with respect to Y$^-$ ions?
   - A   0.05 M
   - B   0.10 M
   - C   0.15 M
   - D   0.20 M
   - E   0.40 M

12) For the following elements, the correct order of increasing first ionization energy is
   - A   Be > Li
   - B   Li > Be
   - C   Na > Be
   - D   O > F
   - E   Na > Li

13) A sample of CH_2F_2 with a mass of 10.4 g contains _____ atoms of F.
   A 6.02×10^{23}
   B 6.02×10^{22}
   C 6.02×10^{21}
   D 1.204×10^{23}
   E 1.204×10^{22}

14) The molarity (M) of an aqueous solution containing 52.5 g of sucrose ($C_{12}H_{22}O_{11}$) RMM = 342 g/mol in 3550 mL of solution is nearest to
   A 5.46×10^{-2}
   B 1.48×10^{-2}
   C 0.104×10^{-2}
   D 4.32×10^{-2}
   E 1.85×10^{-2}

15) The density (D = PM/RT) of a gas Z (RMM = 44) at s.t.p. in g L^{-1} is closest to
   A 0.50
   B 2.0
   C 1.0×10^{2}
   D 0.13
   E 2.5×10^{-4}

16) The enthalpy change for which reaction represents the first ionization energy of magnesium?
   A $Mg(s) \rightarrow Mg^+(g) + e^-$
   B $Mg(g) \rightarrow Mg^+(g) + e^-$
   C $Mg(g) \rightarrow Mg^+(s) + e^-$
   D $Mg(s) \rightarrow Mg^{2+}(g) + 2e^-$
   E $Mg(g) \rightarrow Mg^{2+}(g) + 2e^-$

17) The O-C-O bond angle in the CO_2 is approximately
   A 90°
   B 109.5°
   C 120°
   D 180°
   E 60°

18) Successive ionization energies of an element increase because
   A As each electron is removed, proton number increases.
   B As each electron is removed, proton number decreases.
   C As each electron is removed, mass number decreases.
   D As each electron is removed the charge on the resulting ion increases.
   E As each electron is removed there are a greater number of protons attracting fewer electrons.

19) Which element is expected to have the most negative electron affinity?
   A S
   B Cl
   C Se
   D Br
   E I

20) Calcium carbide (CaC_2) reacts with water to produce acetylene (C_2H_2):
   $CaC_2 (s) + 2H_2O (g) \rightarrow Ca(OH)_2 (s) + C_2H_2 (g)$
   Production of 13 g of C_2H_2 requires consumption of _____ g of H_2O.
   A 4.5
   B 9.0
   C 18
   D 4.8×10^2
   E 506

21) A sample of O_2 compressed to a smaller volume at constant temperature. Which statement(s) is correct regarding O_2 molecules?
   A The average kinetic energy of O_2 molecules is increased.
   B The average speed of O_2 molecules is decreased.
   C The average distance between molecules increases.
   D The number of collisions per unit time, of O_2 molecules with unit area of wall, increases.
   E The average mass of the molecules increases.

22) The shell of lowest energy that contains f orbitals is the shell with n =
   A 3
   B 2
   C 4
   D 1
   E 5

23) Electrons in the outer shell of which atom experiences the greatest effective nuclear charge?
   A Li
   B Be
   C B
   D C
   E N

24) A chemical reaction occurs in a calorimeter containing 100 g of dilute solution producing a rise in temperature of 3.2°C. The specific heat of the contents is 4.2 J·g^{-1}·K^{-1}. ΔH in kJ for the amount of reaction that occurs is nearest to:
   A +1.3
   B -1.3
   C 1300
   D -1300
   E 0.076

25) Which statement is not true for oxygen in its lowest energy state?
   A It is in group VI.
   B It is in period 2.
   C It is a d-block element.
   D The electrons of lowest energy are in a 1s-orbital.
   E The oxygen atom has 2 unpaired electrons.

26) Which one of the following is a unit of pressure?

    A $J \cdot K^{-1}$

    B $g \cdot cm^{-3}$

    C $N \cdot m^{-2}$

    D $g \cdot mol^{-1}$

    E $mol \cdot dm^{-3}$

27) In the nitrite ion (NO_2^-),

    A both bonds are single bonds

    B both bonds are double bonds

    C one bond is a double bond and the other is a single bond

    D both bonds are the same because of resonance

    E there is one fewer electron than the number of protons.

28) Ethane has a lower boiling point than water partly because

    A it is a covalent compound.

    B its molecules do not form hydrogen bonds.

    C it has a larger molecular mass.

    D it does not contain ionic bonds.

    E it contains carbon atoms.

29) Which atom/ion has eight valence electrons?

    A Ti^{4+}

    B Kr

    C Cl^-

    D Na^+

    E all of the above

30) The most polar bond is found in this compound

    A HBr

    B HI

    C HCl

    D HF

    E H_2

30

SECTION B: STRUCTURED QUESTIONS

Answer each of the following questions in the spaces provided on the question paper. Full marks can only be given to calculations where working is included and the final answer is clearly indicated. Clear and concise expression is an essential part of a good answer. You may write your answers in pencil. There are a total of **65** marks for this section.

1) The following questions relate to qualitative analysis and net ionic equations.

   a) Three salt solutions were tested in an effort to identify the salts present. The results are presented in the table below. Use the results to identify the ions as completely as you can by filling in the spaces in the table with **ONE** possibility in each case.   (6)

TEST	Sample A	Sample B	Sample C
Add sodium hydroxide solution dropwise	White *ppt.* forms which is soluble in excess sodium hydroxide solution	No change observed	Green *ppt.* insol. in excess. *ppt* turns brown on standing/ exposure to air
Add ammonia solution	White *ppt.* forms which is soluble in excess ammonia solution	No change observed	Green *ppt.* insol. in excess. *Ppt* turns brown on standing.
Add barium chloride solution followed by dilute/ strong nitric acid	No change observed.	A dense white *ppt.* forms which is soluble in dil. strong acid to release a gas which turns limewater cloudy	A dense white *ppt.* forms which is insoluble in dil. strong acid.
Add silver nitrate solution	Cream *ppt.* Insoluble in dilute ammonia, soluble in conc. ammonia	Test not necessary.	Test not necessary.
Add MgCl₂ solution	Test not necessary.	Immediate white *ppt.*	Test not necessary.
Flame test	No flame colour	Lilac flame	No flame colour
CATION in sample?	Zn^{2+}	K^+	Fe^{2+}
ANION in sample?	Br^-	CO_3^{2-}	SO_4^{2-}

   b) Write a balanced net ionic equation, including states, to explain each of the following observations. If an equation is unbalanced it will attract zero marks: check carefully. Do not include spectator ions. Each answer line should be completed with an equation.
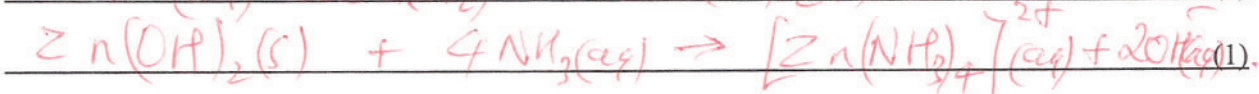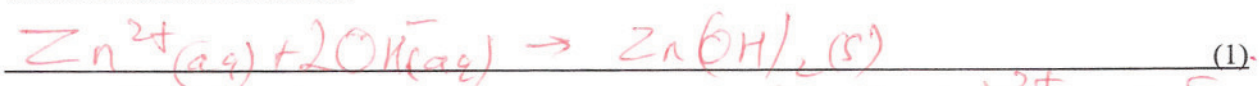
   i. Addition of dilute ammonia solution to a solution of magnesium nitrate results in a white precipitate.

   $$Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$$ (1)

   ii. The addition of dilute sodium hydroxide solution to a solution of aluminium sulfate results in a white precipitate which redissolves in excess to form a colourless solution.
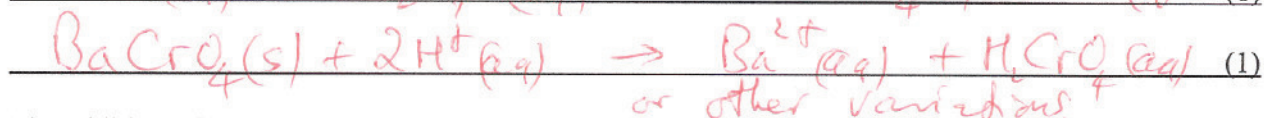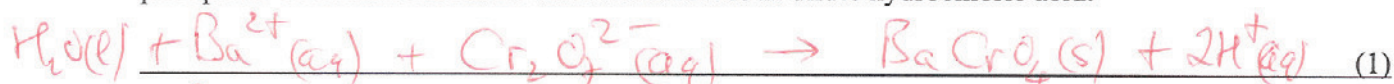
   $$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$$ (1)

   $$Al(OH)_3(s) + OH^-(aq) \rightarrow [Al(OH)_4]^-(aq)$$ (1)
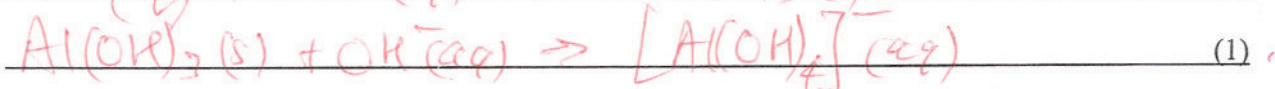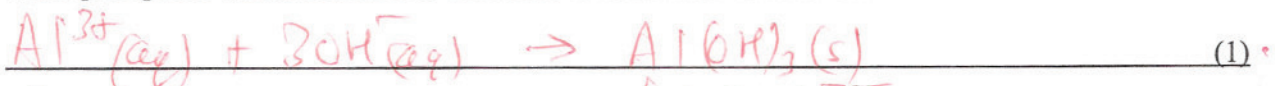
   iii. The addition of dilute ammonia solution to a solution of zinc chloride results in a white precipitate which redissolves in excess.
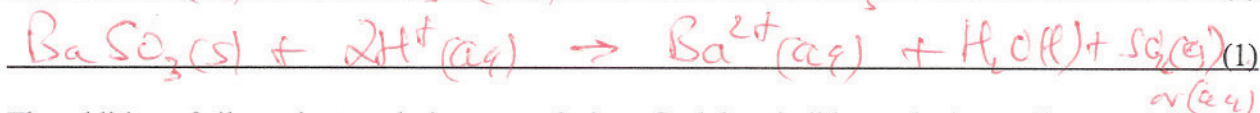
   $$Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn(OH)_2(s)$$ (1)

   $$Zn(OH)_2(s) + 4NH_3(aq) \rightarrow [Zn(NH_3)_4]^{2+}(aq) + 2OH^-(aq)$$ (1)

   iv. The addition of dilute sodium hydroxide to a solution of silver nitrate results in a brown precipitate.

   $$2Ag^+(aq) + 2OH^-(aq) \rightarrow Ag_2O(s) + 2H_2O(l)$$ (1)

   _____   (1)

v. The addition of barium nitrate solution to a solution of sodium dichromate results in a yellow precipitate of barium chromate which redissolves in dilute hydrochloric acid.

$$H_2O(l) + Ba^{2+}(aq) + Cr_2O_7^{2-}(aq) \rightarrow BaCrO_4(s) + 2H^+(aq) \quad (1)$$

$$BaCrO_4(s) + 2H^+(aq) \rightarrow Ba^{2+}(aq) + H_2CrO_4(aq) \quad (1)$$

or other variations

vi. The addition of barium nitrate solution to a solution of sodium sulfite results in a white precipitate which redissolves on the addition of dilute nitric acid to produce a solution which smells of sulfur dioxide.

$$Ba^{2+}(aq) + SO_3^{2-}(aq) \rightarrow BaSO_3(s) \quad (1)$$

$$BaSO_3(s) + 2H^+(aq) \rightarrow Ba^{2+}(aq) + H_2O(l) + SO_2(g) \quad (1)$$

or (aq)

vii. The addition of silver nitrate solution to a solution of calcium iodide results in a yellow precipitate which is insoluble in concentrated ammonia solution.

$$Ag^+(aq) + I^-(aq) \rightarrow AgI(s) \quad (1)$$
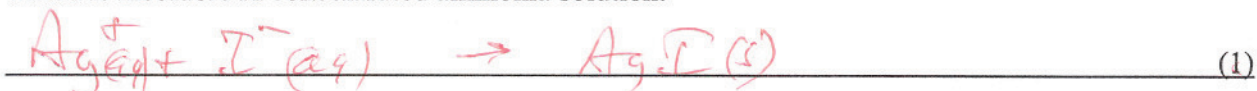
2) The following questions relate to thermochemistry.

a) Define the standard enthalpy of formation, ΔH_f° (2)

This is the enthalpy change involved in the formation of 1 mol of a substance from its constituent elements in their standard states at 1 atm.

b) Write the equation for the complete combustion of **1 mol**, of ethane gas, C_2H_6, (to produce carbon dioxide and liquid water). (1)

$$C_2H_6(g) + 3\tfrac{1}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

The standard enthalpy of combustion of ethane gas, C_2H_6 is -1560 kJmol^{-1}.
Given that $\Delta H_f^\circ(CO_2, (g)) = -394$ kJ·mol^{-1} and $\Delta H_f^\circ(H_2O, (l)) = -286$ kJ·mol^{-1}

c) Calculate ΔH_f° (C_2H_6 (g)) (4)

$$\Delta H_R^\circ = \sum \Delta H_f^\circ (products) - \sum \Delta H_f^\circ (reactants)$$

$$\therefore -1560 = 2 \times (-394) + 3 \times (-286) - \Delta H_f^\circ(C_2H_6) - 0$$

$$\therefore \Delta H_f^\circ(C_2H_6) = 2 \times (-394) + 3 \times (-286) + 1560 = -86 \, kJmol^{-1}$$

d) Use the average bond energies below to estimate a value for the standard enthalpy change, ΔH_f° for the hydrogenation of 1 mol of ethane. Mean bond energies are as follows: C-C = 348 kJ·mol^{-1}, C=C = 615 kJ·mol^{-1}, H-H = 436 kJ·mol^{-1}, C-H = 413 kJ·mol^{-1} (4)
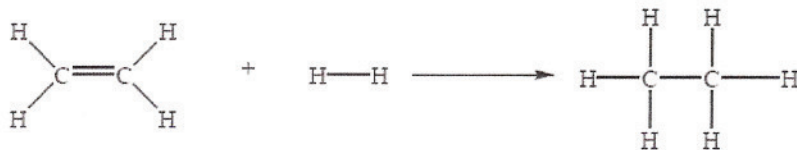
Bonds broken

$4 \times C-H = 4 \times 413$
$1 \times C=C = 1 \times 615$
$1 \times H-H = 436$
2703 kJ

Bonds made.

$6 \times C-H = 6 \times (-413)$
$1 \times C-C = -348$
Total = -2826 kJ

⑯

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$$\therefore \Delta H_R = 2703 - 2826 = -123 \, kJ$$

3) The following questions relate to stoichiometry:
   An organic ester, X, contains 54.55% C, 9.09% H, and 36.36% O by mass.
   a) Determine the empirical formula of X. (RAM of H = 1.00, C = 12.0, O = 16.0) (3)

$$
\begin{array}{cccc}
 & C & H & O \\
 & 54.55\% & 9.09\% & 36.36\% \\
RAM & 12 & 1 & 16 \\
Moles & 4.5483 & 9.09 & 2.2725 \\
\end{array}
$$

Divide by smallest : 2.00037 4.00000 1

∴ Empirical formula is $\underline{C_2H_4O}$

b) 0.15 g of X, when vaporized in a suitable apparatus, occupied a volume of 57.0 cm³ at 127°C and 1.00×10⁵ Nm⁻². Using the expression PV = nRT (or otherwise) calculate the relative molecular mass of X. (3)

$$n = \frac{PV}{RT} = \frac{1.00\times10^5 \times 57\times10^{-6}}{8.31 \times (273+127)} = 1.71480\times10^{-3} \text{ mol}.$$

∴ Molar mass $= \dfrac{0.15g}{1.71480\times10^{-3} mol} = 8.74737\times10^{1}$

∴ RMM $\simeq \underline{87 \ gmol^{-1}} \ to \ 2d.p.$

c) What is the molecular formula of X? (2)

RFM of $C_2H_4O = 2\times12 + 4\times1 + 16 = 44$

Since $87 \simeq 2\times44$

Molecular formula must be $\underline{C_4H_8O_2}$

4) The following questions relate to bonding and electronic configuration.
   a) State:
   i) the Aufbau (building-up) Principle, (2)

This states that in building up the electronic structure of an atom, electrons are added to the ~~nuclear energy~~ levels orbitals in order of increasing energy, lowest first.

ii) the Pauli Exclusion Principle (2)

This states that no two electrons in the same atom can have the same set of 4 quantum nos. ∴ $n, \ell, M_\ell \ \& \ M_s$.
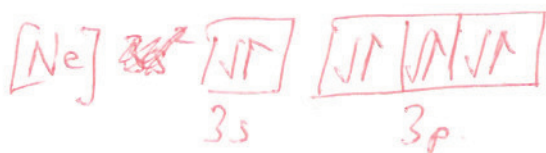
iii) Hund's Rule (2)

This states that when filling degenerate orbitals, each is filled singly (with spins parallel) before any pairing occurs.

b) Draw orbital diagrams for the ground state electron configurations of chloride ($_{17}Cl^-$), chromium ($_{24}Cr$) and nitrogen, ($_7N$). For example, the orbital diagram of nickel, $_{28}Ni$ in the 4th period would be

[Ar] **4s** **3d** (2)

$_{17}Cl^-$:

[Ne] 3s 3p

$_{24}Cr$:

[Ar] 4s 3d
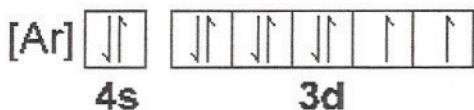
i) How many unpaired electrons are there in an atom of sodium, $_{11}Na$? (1)

One.

ii) Show that the maximum number of electrons that can be accommodated in all the orbitals with principal quantum number n = 4 in a given atom is 32. (2)

If n = 4, ℓ may be 0, 1, 2, 3. For $\ell = 0$, $M_\ell = 0$: $2e^-$. For $\ell = 1$ $M_\ell = -1, 0, +1$: $6e^-$. For $\ell = 2$ $M_\ell = -2, -1, 0, +1, +2$: $10e^-$. For $\ell = 3$, $M_\ell = -3, -2, -1, 0, +1, +2, +3$: $14e^-$. Total $e^- = 2 + 6 + 10 + 14 = 32$.

5) Draw Lewis electron dot diagrams to represent the bonding in the molecules: PCl_5, BF_3, and NH_3.

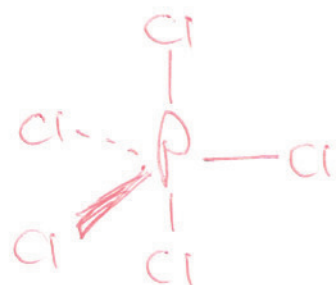a) Show all valence electrons. (3)

i. PCl_5 ii. BF_3 iii. NH_3

b) Draw diagrams to indicate the shapes of the above molecules according to VSEPR theory and write a descriptive term for each shape. (For example, for the methane molecule you might write "tetrahedral").(6)

trigonal bipyramidal

trigonal planar

trigonal pyramidal

6) The following question relates to gases.
   a) Define *partial pressure* (2)

   The partial pressure of a gas is the pressure it would
   _have if it alone filled the container at the same temp._

   b) State Dalton's law of partial pressure (2)

   The sum of the partial pressures of the components
   of a mixture of gases is the pressure of the mixture

   c) A flask of volume 600 cm^3 contains a gas, A, at 120 mmHg that is connected via a valve to a second flask of volume 1200 cm^3 that contains a gas, B, at 255 mmHg. The stopcock is opened and the gases are allowed to mix completely. Find the final pressure of the mixture, assuming no change in temperature. (Neglect the volume of the tube connecting the bulbs.) (5)

   Gas A Valve Gas B
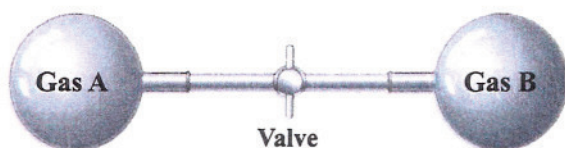
   Boyles Law $P_1 V_1 = P_2 V_2$

   For A: $P_2 = \dfrac{P_1 V_1}{V_2} = \dfrac{120 \times 600}{(1200 + 600)} = \dfrac{72000}{1800} = 40\,mmHg$

   For B: $P_2 = \dfrac{255 \times 1200}{1800} = 255 \times \dfrac{2}{3} = 85 \times 2$

   $= 170\,mmHg$

   ∴ Final pressure $= 40\,mmHg + 170\,mmHg$

   $= 210\,mmHg$

⑨

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