THE COPPERBELT UNIVERSITY

SCHOOL OF MATHEMATICS AND NATURAL SCIENCES

SESSIONAL EXAMINATION 2014

TITLE OF PAPER: GENERAL CHEMISTRY

COURSE NUMBER: CH110/CH120/FO130

TIME: THREE (3) HOURS

INSTRUCTIONS:

There are seven questions in two sections, A and B. Each question is worth 20 marks. Answer a total of five (5) questions with at least two(2) questions from each section.

All calculated quantities must have units and be reported to the correct number of significant figures otherwise marks will be deducted.

A data sheet and a periodic table are provided

Non-programmable electronic calculators may be used.

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SECTION A: ANSWER AT LEAST TWO QUESTIONS FROM THIS SECTION

Question 1 (20 Marks)–(chemequil and acid base equil)

(a) **Explain the following key terms in the study of chemical equilibrium,**
    (i) Law of mass action **[1]**

A general description of the equilibrium condition; it defines the equilibrium constant expression

    (ii) Reaction quotient **[1]**

A quotient obtained by applying the law of mass action to initial concentrations rather than to equilibrium concentrations

(b) **Write the equilibrium expression for each of the following reactions**
    (i) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ **[1]**

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

    (ii) $S_8(g) \rightleftharpoons 8S(g)$ **[1]**

$$K_c = \frac{[S]^8}{[S_8]}$$

    (iii) $Cl_2O_7(g) + 8H_2(g) \rightleftharpoons 2HCl(g) + 7H_2O(g)$ **[1]**

$$K_c = \frac{[H_2O]^7[HCl]^2}{[Cl_2O_7][H_2]^8}$$

(c) **Consider the decomposition equilibrium reaction of methanol to carbon monoxide and hydrogen:$CH_3OH(g) \rightleftharpoons CO(g) + 2H_2(g)$.**
    (i) **Calculate K_p for the reaction given that $P_{CH_3OH} = 6.10 \times 10^{-4}$ atm, $P_{CO} = 0.387$ atm and $P_{H_2} = 2.68$ atm.** **[2]**

We use the appropriate mass action expression $K_p = \frac{P_{H_2}^2 P_{CO}}{P_{CH_3OH}}$[1].

Substituting the given values yields $K_p = \frac{(2.68\ atm)^2(0.387\ atm)}{(6.10\times10^{-4}\ atm)} = 4.56 \times 10^3 atm^2$ [1]

    (ii) **Derive the expression that relates K_c at K_p for this reaction** **[4]**

The ideal gas law,$PV = nRT$, gives the concentration expression $\frac{n}{V} = \frac{P}{RT}$[½]

The mass expression for the reaction is $K_c = \frac{[H_2]^2[CO]}{[CH_3OH]}$[½]

Substituting the LHS of ideal gas concentration expression into the RHS of K_c, we have $K_c = \frac{\left[\frac{n}{V} \text{ for } H_2\right]^2 \left[\frac{n}{V} \text{ for } CO\right]}{\left[\frac{n}{V} \text{ for } CH_3OH\right]}$[1]

2

Substituting $\frac{n}{V}$ with $\frac{P}{RT}$ in the above equation, we have $K_c = \frac{\left(\frac{P_{H_2}}{RT}\right)^2 \left(\frac{P_{CO}}{RT}\right)}{\left(\frac{P_{CH_3OH}}{RT}\right)}$ [1]

Factoring out pressures, we have $K_c = \frac{P_{H_2}^2 P_{CO}}{P_{CH_3OH}} \left[\frac{(RT)^{-2}(RT)^{-1}}{(RT)^{-1}}\right]$ [1]

Since $K_p = \frac{P_{H_2}^2 P_{CO}}{P_{CH_3OH}}$, therefore $K_c = K_p(RT)^{-2}$ [1]

(d)   **The average pH of normal arterial blood is 7.40. At normal body temperature (37 ºC) , K_w = 2.4 x 10^{-14}. Calculate [H⁺], [OH⁻] and pOH for blood at this temperature.**   **[5]**
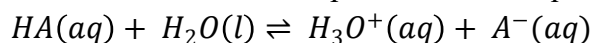
$[H_3O^+] = 10^{-pH} = 10^{-7.40} = 3.98\ X10^{-8} M$ [2]

$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{2.4 x 10^{-14}}{3.98 x 10^{-8}} = 6.0 x 10^{-7}\ M$ [2]

$pOH = -\ Log\ [OH^-] = 6.2$ [1]

(e)   **Lactic acid ($HC_3H_5O_3$) has one acidic hydrogen atom. A 0.10 M solution of lactic acid has a pH of 2.44. Calculate its K_a at 298 K.**   **[4]**

Represent acid HA and then the proton transfer equilibrium is
$$HA(aq) +\ H_2O(l) \rightleftharpoons\ H_3O^+(aq) +\ A^-(aq)$$
$K_a = \frac{[H_3O^+][A^-]}{[HA]}$ [1]
$[H_3O^+] = 10^{-pH} = 10^{-2.44} = 3.63 x 10^{-3}$ M[1]
$[A^-] = [H_3O^+]$
Since $\frac{3.63 x 10^{-3}}{0.10} x 100\% = 3.6\% < 5\%$ we assume $[HA] \approx 0.10 M$, hence
$K_a = \frac{3.63 x 10^{-3}\ x\ 3.63 x 10^{-3}}{0.10} = 1.3 x 10^{-4}$ [2]
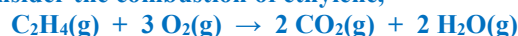
Question 2 (20marks)- (kinetics)

(a) **List four factors that affect reaction rates.**   **[2]**
Concentration, physical state of reactants, temperature and catalyst [½ each]

(b) **The rate law for a reaction is $rate = [A][B]^2$. What is the order with respect to [A], with respect to [B] and the overall order?**   **[3]**

Order with respect to [A] is 1, order with respect to [B} is 2 and overall order is 3

(c) **Consider the combustion of ethylene,**
     $C_2H_4(g) +\ 3\ O_2(g) \rightarrow\ 2\ CO_2(g) +\ 2\ H_2O(g)$
If the concentration of C_2H_4 is decreasing at the rate of 0.37 M/s, what are the rates of change in the concentration of CO_2 and O_2?   **[4]**

$$\frac{d[CO_2]}{dt} = 2 \times \frac{d[C_2H_4]}{dt} = 0.74 \text{ M/s} \qquad \text{[2]}$$

$$\frac{d[CO_2]}{dt} = 3 \times \frac{d[C_2H_4]}{dt} = 1.1 \text{ M/s} \text{ [2]}$$

(d) The first order rate constant for the decomposition N_2O_5,

$N_2O_5(g) \rightarrow 2\ NO_2(g) + O_2(g)$,

At 70 °C is 6.82 x 10^{-3} s^{-1}. Suppose we start with 0.0250 mol of $N_2O_5(g)$ in a volume of 2.0 l. (First order rate law $[A]_t = [A]_0 e^{-kt}$)

     (i)   How many moles of N_2O_5 will remain after 2.5 min?   **[2]**

Since $n \propto [J]$ we   can   write

$n_t = n_0 e^{-kt} = 0.0250\ mol\ x e^{-6.82 x 10^{-3} s^{-1} x 2.5 x 60\ s} = 9.0\ x 10^{-3} mol$**[2]**

     (ii)   How many minutes will it take for the quantity of N_2O_5 to drop to 0.010 mol?   **[2]**

Using $n_t = n_0 e^{-kt}$ and taking logarithms on both sides and solving for t gives

$$t = \frac{ln([A]_0/[A]_t)}{k} = \frac{ln(0.0250/0.010)}{6.82 x 10^{-3} s^{-1}} = \mathbf{134}\ \mathbf{s} \text{or } \mathbf{2.2\ min} \qquad \text{[2]}$$

     (iii)   What is the half-life of N_2O_5 at 70 °C?   **[2]**

$$t_{1/2} = \frac{ln\ 2}{k} = \frac{0.693}{6.82 x 10^{-3} s^{-1}} = \mathbf{102\ s\ 0r\ 1.70\ min}\text{[2]}$$

(e) A certain first order reaction has a rate constant of 2.75 x 10^{-2} s^{-1} at 20 °C. What is the value of k at 60 °C if the E_a = 75.5 kJ/mol. (Arrhenius equation $k = Ae^{-E_a/RT}$) **[5]**

From $k = Ae^{-E_a/RT}$ we get

$$Ln\left(\frac{k_2}{k_1}\right) = -\frac{E_a}{R}\left[\frac{1}{T_2} - \frac{1}{T_1}\right] \quad (2)$$

$$= -\frac{75.5 x 10^3 J mol^{-1}}{8.314 J\ K^{-1} mol^{-1}}\left[\frac{1}{333K} - \frac{1}{293\ K}\right] = 3.7229 (2)$$

$$\frac{k_2}{K_1} = e^{3.7229} = 41.38$$

$$k_2 = k_1 x 41.38 = 2.75 x 10^{-2} s^{-1} x 41.38 = 1.38\ x 10^1 s^{-1} (1)$$

Question 3 (20 marks) (thermochemistry)

(a)   Naphthalene, $C_{10}H_8$, is a solid aromatic compound often sold as mothballs. The complete combustion of this substance to $H_2O(l)$ and $CO_2(g)$ at 25 °C yields 5154 kJ/mol. Write a thermo-chemical equation for the combustion of $C_{10}H_8$.   **[2]**

$$\mathbf{C_{10}H_8(s) + O_2(g) \rightarrow CO_2(g) + H_2O(l)} \quad \mathbf{\Delta H = -5154\ kJ} \quad (2)$$
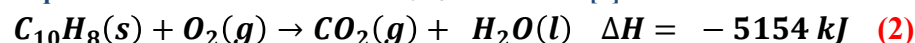
(b)   Consider the combustion of liquid methanol, $CH_3OH(l)$

          $CH_3OH(l) + 3/2 O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$ $\Delta H = -726.5$ kJ

     (i)   What is the enthalpy change for the reverse reaction?   **[1]**

     (ii)   Balance the forward reaction with whole-number coefficients. What is ΔH for the reaction represented by this equation?   **[2]**

     (iii)   Which is more likely to be thermodynamically favoured, the forward reaction or the reverse reaction? Explain.   **[2]**

(iv)   If the reaction were written to produce $H_2O(g)$ instead of $H_2O(l)$, would you expect the magnitude of ΔH to increase, decrease, or stay the same? Explain.   [2]

(i)   ΔH = + 726.5 kJ   (1)
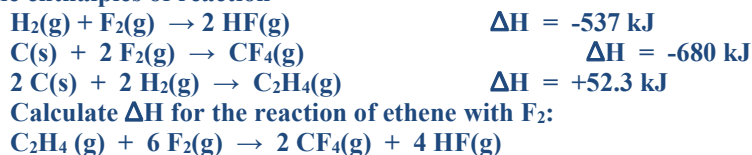
(ii)   2 $CH_3OH(l)$ + 3$O_2(g)$ → 2 $CO_2(g)$ + 4 $H_2O(l)$ ΔH = -1453 kJ (2)
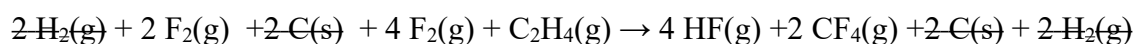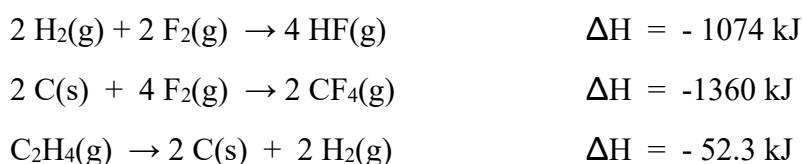
(iii) $\Delta G = \Delta H - T\Delta S$
ΔS for this reaction is negative but small (3 reactant gas moles gives 2 product gas moles). ΔH is negative and very large, hence ΔG is negative at low and moderate temperatures and forward reaction is favoured. (2)

(iv) ΔH would decrease as some of the energy would be used to vaporize the water. (2)

(c)   From the enthalpies of reaction
$H_2(g) + F_2(g)$ → 2 HF(g)   ΔH = -537 kJ
C(s) + 2 $F_2(g)$ → $CF_4(g)$ ΔH = -680 kJ
2 C(s) + 2 $H_2(g)$ → $C_2H_4(g)$ ΔH = +52.3 kJ
Calculate ΔH for the reaction of ethene with F_2:
C_2H_4 (g) + 6 $F_2(g)$ → 2 $CF_4(g)$ + 4 HF(g)   [5]

Reverse 3rd reaction, multiply 1st and 2nd reaction by 2 add the resulting three reactions to get required reaction: (3)

2 $H_2(g)$ + 2 $F_2(g)$ → 4 HF(g)   ΔH = - 1074 kJ

2 C(s) + 4 $F_2(g)$ → 2 $CF_4(g)$ ΔH = -1360 kJ

$C_2H_4(g)$ → 2 C(s) + 2 $H_2(g)$ ΔH = - 52.3 kJ

~~2 H₂(g)~~ + 2 $F_2(g)$ + ~~2 C(s)~~ + 4 $F_2(g)$ + $C_2H_4(g)$ → 4 HF(g) +2 $CF_4(g)$ + ~~2 C(s)~~ + ~~2 H₂(g)~~

6 $F_2(g)$ + $C_2H_4(g)$ → 4 HF(g) +2 $CF_4(g)$ which is the required reaction. Hence

ΔH = - 1074 kJ - 1360 kJ - 52.3 kJ = -2486 kJ   (2)

(d)   Given that standard enthalpy of formation, ΔH_f^θ, of HCN(g) is 135 kJ/mol and the standard entropy values in the table below:

Compound	$H_{2(g)}$	$N_{2(g)}$	$C_{(s)}$	$HCN_{(g)}$
S° J K^{-1} mol^{-1}	218	413	716	202

(i)   Write the thermo-chemical equation corresponding ΔH_f^θ of HCN(g).[1]

(ii)   Calculate the value of the Gibbs energy of formation, ΔG_f^θ, in kJ mol^{-1} at 25ºC of HCN[3]

(iii)   Determine the temperature <u>above</u> or <u>below</u> which the reaction is spontaneous. [2]

(i) $\frac{1}{2}H_2(g) + \frac{1}{2}N_2(g) + C(graphite) \rightarrow HCN(g)$ $\Delta H_f^o = + 135\ kJ/mol$

(2)

(ii) $\Delta G = \Delta H - T\Delta S$ (1)

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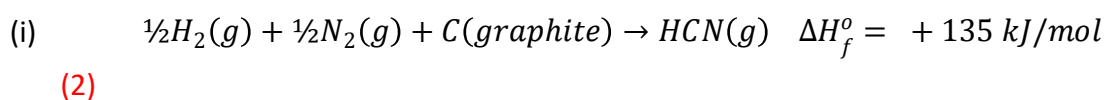
$$\Delta S = \Sigma n_J S^o(J) = 202 - \frac{1}{2}(218) - \frac{1}{2}(413) - 716 = -829.5 \, JK^{-1}mol^{-1}(1)$$
$$\Delta G = 135 kJ/mol - 298 Kx(-0.8295 \, kJ \, K^{-1}/mol) = 382 \text{ kJ/mol } (1)$$

(iii) For reaction to be spontaneous ΔG must be negative.

In this case since $\Delta H > 0$ and $\Delta S < 0$ so ΔG is always positive and reaction is not spontaneous at any temperature. (2)

SECTION B: ANSWER AT LEAST TWO (20) QUESTIONS FROM THIS SECTION

Question 4 (20 marks) (gases, stoichiometry) (NON QUOTA)

(a) A 200. mL sample of hydrogen chloride at 690. Torr and 20. °C is dissolved in 100. mL of water. The solution was titrated to the stoichiometric point with 15.7 mL of sodium hydroxide solution.

    (i) Write a balanced equation for the titration reaction. [1]

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)(1)$
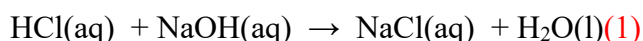
    (ii) What is the molar concentration of the NaOH solution? (4)   n

$$n(HCl) = \frac{PV}{RT} = \frac{690 \, Torr \, x \, 0.200 \, L}{62..36 \, L \, atm \, K^{-1}mol^{-1}} = 7.55 \, x \, 10^{-3} mol \text{ HCl} \qquad (2)$$

From balanced equation 1 mol HCl$\hat{=}$ 1 $mol \, NaOH$

Hence $n(NaOH) = 7.55 \, x \, 10^{-3} mol$ NaOH

$$M = \frac{n}{V} = \frac{7.55 \, x \, 10^{-3} mol}{0.0157 \, L} = 0.481 \, mol \, L^{-1} \quad (2)$$

(b) The analysis of a hydrocarbon revealed that it was 85.7% C and 14.3% H by mass. When 1.77 g of the gas was stored in a 1.500 L flask at 17 °C, it exerted a pressure of 508 Torr. Determine

    (i) the empirical formula of the hydrocarbon [3]

        In 100 g of sample there are   85.7 g C and   14.3 g H

        Converting to mols $85.7 gx\frac{1 \, mol}{12.01 \, g}$ $14.3 \, g \, x\frac{1 \, mol}{1.008 \, g}$

                                  $= 7.134 \, mol \, C$ $=14.19$ mol H

        Divide by the smallest   1.000   1.9888hence the mole ratio is 1 C:2 H

        and empirical formula is CH_2. (3)

    (ii) the molar mass of the hydrocarbon [3]

$$\text{Molar mass } M = \frac{mRT}{pV} = \frac{1.77 \, g \, x \, 62.36 \, L \, atm \, Torr^{-1}K^{-1}x290 \, K}{508 \, Torr \, x \, 1.500 \, L} = 42.0 \, g \, mol^{-1}(3)$$

   (iii) the molecular formula of the hydrocarbon? [1]

        Molar mass of CH_2 is 14.02g/mol

        Hence there are 42.0/14.0 =3 CH_2 units per molecule

        The molecular formula is C_3H_6(1)

(c) Iron pyrite, FeS_2, is the form in which much of the sulphur occurs in coal. In the combustion of the coal, oxygen reacts with iron pyrite to produce iron(III) oxide and sulphur dioxide.

    (i) Write a balanced equation for the reaction of iron pyrite with oxygen. [2]

        $4 \, FeS_2(s) + 11 \, O_2(g) \rightarrow 2 \, Fe_2O_3(s) + 8 \, SO_2(g)(2)$
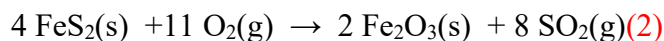
    (ii) Calculate the mass of Fe_2O_3 that is produced from the reaction of 75.0 L of oxygen at 2.33 atm and 150. °C with an excess of iron pyrite. [4]

6

$$n(O_2) = \frac{pV}{RT} = \frac{2.33\ Atm \times 75.0\ L}{0.08206\ L\ atm\ K^{-1}mol^{-1} \times 423\ K} = 5.03 \text{ mol}(2)$$

Number of mol Fe_2O_3 =5.03 $mol\ O_2 x \frac{2\ mol\ Fe_2O_3}{11\ mol\ O_2}$ =0.912 molFe_2O_3

Mass of Fe_2O_3 = 0.912 molFe2O3x159.68 g/mol= 1.46 x10²g(3)

(iii)   If the sulphur dioxide that is produced is dissolved to form 5.00 L of aqueous solution, what is the molar concentration of the resulting sulphurous acid, H_2SO_3, solution?   [2]

Number of mol SO_2 =5.03 $mol\ O_2 x \frac{8\ mol\ SO_2}{11\ mol\ O_2}$ = 3.66 mol SO_2

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3$.

Number of mol H_2SO_3 =Mol SO_2.

Hence molarity $= \frac{n}{V} = \frac{3.66\ mol}{5.00\ L} = 0.732\ M(2)$

OR

Question 4 (20 marks) (Biochemistry) (NATURAL RESOUCES)

Question 5 (20marks) (electron configuaton, Lewis symbols&structures, chem bonding and molecular geometry, intermolecular forces)

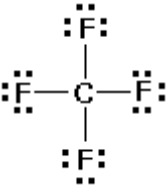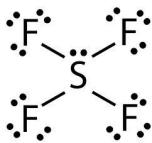(a)   Consider the following elements: magnesium, carbon, and chlorine.
   (i)   Write the ground state electron configuration of each element.   [3]

Mg: $1s^2 2s^2 2p^6 3s^2$ C: $1s^2 2s^2 2p^2$ Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$ [3]

   (ii)   Use an appropriate pair of the above elements and their Lewis symbols to illustrate covalent bond formation.   [2]
   (iii)   Use an appropriate pair of the above elements and their Lewis symbols to illustrate ionic bond formation.   [2]

(b) Consider the following molecules: CF₄ and SF₄
   (i) Write the Lewis structure of each. [3]
   (ii) Predict the shape of each molecule using VSEPR model. [2]
   (iii) Predict whether molecule is polar or non-polar. [2]

species	CF₄	SF₄
No. Of electrons	4+4x7=32 electrons or 16 e-pairs	6+4x7=44 electron or 17 paits
(i)Lewis structure	 (1.5)	 (1.5)
(ii)	4 electron pairs around central atom. Electron arrangement is tetrahedral and molecular shape tetrahedral (1)	5 electron pairs around central atom. Electron arrangement is trigonal bypyramidal with lone pair on equatorial position. molecular shape see-saw (1)

(iii)   CF₄ is non-polar while SF₄is polar(1)

(c) For the molecule N₂O
   (i) Write the Lewis structures that contribute to its resonance hybrid. (skeleton is N-N-O) [2]
   (ii) Calculate the formal charges on all atoms in the above structures. [3]
   (iii) Select the structure that is likely to make a dominant contribution to the resonance hybrid. [1]

(i)Resonance hybrids below (2)   (ii) formal charges also below (3)
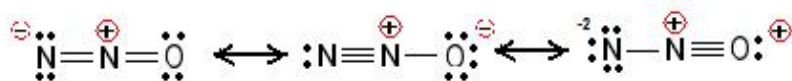
The three resonance structures are shown above with the formal charges on the atoms. The middle structure is most stable as it places the negative charge on O the more electronegative atom.(1)

Question 6 (20 marks) (Organic chemistry)

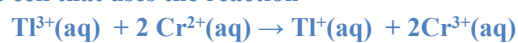## Question 7 (20 marks) (aqueous solutions and electrochemistry)

(a) **Assign oxidation numbers to underlined elements in the following compounds and ions,**

    (i) $\underline{Cs}_2O$ **[½]**   **(ii)** $H_3\underline{As}O_3$ **[½]**   **(iii)** $\underline{Mo}O_4^{2-}$ **[½]**   **(iv)** $\underline{Cl}F_3$ **[½]**

    (i)   Cs +1   (ii) As +3   (iii) Mo +6   (iv) Cl +3

(b) **Classify the following reactions according to the types of aqueous reactions you learnt in CH110/FO130.**

    (i) $Cl_2(g) + 2OH^-(aq) \rightarrow Cl^-(aq) + ClO^-(aq) + H_2O(l)$ **[1]**

        Redox reaction(1)

    (ii) $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$ **[1]**

        Precipitation reaction(1)

    (iii) $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$ **[1]**

        Neutralizationrreaction(1)

(c) **A 60.00 mL 0.513 M glucose ($C_6H_{12}O_6$) solution is mixed with 120 mL of 2.33 glucose solution. What is the concentration of the final solution?** *(Assume additive volumes)*   **[4]**

     Let moles, volume and molarity for first solution be n_1, V_1 and M_1, respectively and Let moles, volume and molarity for second solution be n_2, V_2 and M_2, respectively.

     The moles for each solution is $n_1 = V_1 \times M_1$ and $n_2 = V_2 \times M_2$ **[½]**

     Total moles of the combined solutions is $n = n_1 + n_2$ **[½]**
     Total solution volume $V = V_1 + V_2$ **[½]**
     Thus, molarity of combined solution is $M = \frac{n}{V}$ **[½]**
     Or $M = \frac{n_1 + n_2}{V_1 + V_2}$ **[½]** thatis, $M = \frac{(V_1 \times M_1) + (V_2 \times M_2)}{V_1 + V_2}$ **[½]**

     Substituting $V_1 = 60\,mL, M_1 = 0.513\frac{mol}{L}, V_2 = 120\,mL$ and $M_2 = 2.33\frac{mol}{L}$ gives

$$M = \frac{\left(60\,mL \times 0.513\frac{mol}{L}\right) + \left(120\,mL \times 2.33\frac{mol}{L}\right)}{60\,mL + 120\,mL} = 0.324\frac{mol}{L} \text{[1]}$$

(d) **Balance the following reaction occurring in acidic solution**

       $NO_2^-(aq) + Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq) + NO_3^-(aq)$ **[4]**

       Oxidation: $NO_2^- + H_2O \rightarrow NO_3^- + 2H^+ + 2e^-$ (1)

       Reduction: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + \;\; + 7H_2O$(1)

       Multiply the oxidation equation by 3, add to the reduction equation and simplify to get overall equation:

   $3NO_2^-(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3NO_3^-(aq) + 2Cr^{3+}(aq) + \;\; + 4H_2O(l)$(2)

(e) **A voltaic cell consists of a silver strip placed in a solution of AgNO₃(aq) and an iron strip immersed in FeCl₂(aq) solution. The overall cell reaction is**

       Fe(s) + 2 Ag⁺(aq) → Fe²⁺(aq) + 2 Ag(s)

    (i)   **Write the half reactions that occur in the two electrode compartments.**   **[2]**

(ii)   **Which electrode is the anode and which the cathode?**   **[2]**

Cathode: $2\ Ag^+(aq) + 2e^- \rightarrow 2\ Ag(s)$ (1)

Anode: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$(1)
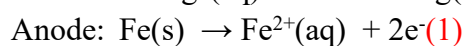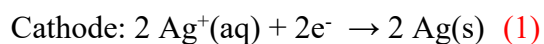
(f)   **A voltaic cell that uses the reaction**

$Tl^{3+}(aq) + 2\ Cr^{2+}(aq) \rightarrow Tl^+(aq) + 2Cr^{3+}(aq)$

has a measured standard cell potential of +1.19 V. Given that

$E^o_{red}(Cr^{2+}/Cr^{3+}) = -0.41$ **determine** E^o_{red} **for reduction of Tl³⁺(aq) to Tl⁺(aq) .**

[3]

In this reaction the Cr^{2+}/Cr^{3+} couple is anode and the Tl^{3+}/Tl^+ the cathode(1)

$E^o_{cell} = E^o_{cath} - E^o_{anod}$

Hence $E^o_{cathode} = E^o_{cell} + E^o_{anod} = +1.19 + (-0.41) = +0.78\ V$(2)