Year 12 Chemistry Exam Semester 1, 2002 Solutions

PART 1

1c, 2a, 3c, 4b, 5a, 6a, 7a, 8c, 9d, 10d, 11d, 12a, 13b, 14d, 15d, 16b, 17c, 18d, 19c, 20a, 21b, 22a, 23b, 24c, 25d, 26c, 27a, 28c, 30a

PART 2

Question 1

Write equations for any reactions that occur in the following procedures. If no reaction occurs write 'no reaction'.

In each case describe **in full** what you would observe, including any
- colours
- odours
- precipitates (give the colour)
- gases evolved (give the colour or describe as colourless).

If no change is observed, you should state this.

(a) Solid sodium hydroxide is added to dilute hydrochloric acid.

Equation $NaOH(s) + H^+(aq) \longrightarrow H_2O(\ell) + Na^+(aq)$

Observation Solid dissolves

[3 marks]

(b) Dilute ammonia solution is added to sodium hydroxide solution.

Equation NO VISIBLE REACTION

Observation

[3 marks]

(c) Iron(III) nitrate solution is added slowly to sodium bromide solution.

Equation NO VISIBLE REACTION

Observation

[3 marks]

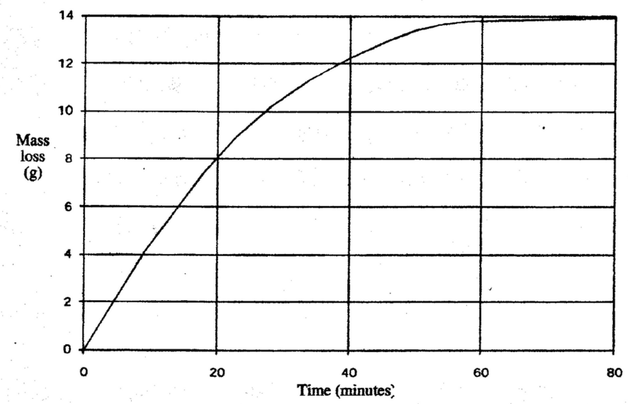
(d) Dilute sulfuric acid is added to a solution of potassium chromate.

Equation $2CrO_4^{2-}(aq) + 2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(\ell)$

Observation Yellow solution turns orange.

Question 2

An excess of $CaCO_3$ in the form of large pieces of marble is reacted with 500 mL hydrochloric acid in an open flask standing on a balance. The loss in mass of the contents of the flask as a function of time is shown graphically below.



(a) Write the equation for the reaction which occurs. Why is there a loss in mass of the flask and its contents?

[2 marks]

1 $CaCO_3(s) + 2H^+(aq) \longrightarrow CO_2(g) + H_2O(\ell) + Ca^{2+}(aq)$

1 There is a loss in mass of the flask and its contents because $CO_2(g)$ is leaving the system to the atmosphere

(b) Calculate the average rate of reaction between 0–20 minutes and also between 20–40 minutes (Take care with units). Explain any difference between these two rates of reaction.

[2 marks]

0.5 Average Rate (0-20mins) = $\frac{8\text{ g } CO_2 \text{ loss}}{20\text{mins}}$

= 0.4g/min mass loss

0.5 Average Rate (20-40mins) = $\frac{4.4\text{ g } CO_2 \text{ loss}}{20\text{mins}}$

= 0.22g/min mass loss

1 As time progresses the reactant HCl becomes less concentrated + so rate decreases

(c) Why does the mass loss of the contents of the flask not increase after 70 minutes?

[1 mark]

THE HYDROCHLORIC ACID IS TOTALLY CONSUMED SINCE THE CALCIUM CARBONATE IS IN EXCESS. HENCE THE REACTION STOPS AND NO MORE $CO_2(g)$ IS PRODUCED AND LOST

(d) Explain *clearly* what would have been the effect on the initial rate of reaction if small chips of marble had been used?

[1 mark]

BECAUSE THE REACTION RATE INCREASES WITH INCREASED SURFACE AREA OF REACTANT, SMALLER MARBLE CHIPS WILL CAUSE A HIGHER INITIAL RATE OF REACTION

Question 3

A chemical engineer has been given the task of maximising the yield of a chemical process, while minimising the cost of the process.

Assess the range of factors that need to be considered by the engineer. In your answer refer to the Haber process.

[5 marks]

THE ENGINEER NEEDS TO CONSIDER TEMPERATURE, PRESSURE AND CATALYST USE.

TEMPERATURE: REACTION IS EXOTHERMIC + SO HIGH YIELDS COME FROM LOW TEMPERATURES. 1 BUT LOW TEMPERATURES MEAN REACTION IS TOO SLOW — NOT ECONOMICALLY VIABLE. 1 BY USING A CATALYST, A COMPROMISE TEMPERATURE OF 500°C IS THE BEST 1

PRESSURE: HIGH PRESSURES INCREASE RATE AND YIELD. 1 HOWEVER, HIGH PRESSURES ARE EXPENSIVE TO MAINTAIN BY USING A CATALYST (IRON/IRON OXIDE) A PRESSURE OF ABOUT 350 atm IS BEST. 1

Question 4

Identify by name or formula an example of each of the following

Description	Name or Formula
A positively charged complex ion	$[Ag(NH_3)_2]^+$
A halogen which is liquid at room temperature and pressure	Bromine
A salt that dissolves in water to give an acidic solution	NH_4Cl
A weak acid other than acetic acid	CARBONIC ACID
A substance that can be used as a primary standard for redox titrations	$H_2C_2O_4$
A diprotic acid	H_2SO_4
A secondary standard used in many acid-base titrations	Hydrochloric Acid.

[7 marks]

Question 5

(a) Consider the following reaction:

$$3Cl_{2(g)} + 6NaOH_{(aq)} \rightarrow 5NaCl_{(aq)} + NaClO_{3(aq)} + 3H_2O$$

[4 marks]

What is the oxidation state of

(i) Cl in $Cl_{2(g)}$ 0

(ii) Cl in $NaCl_{(aq)}$ .... −1

(iii) Cl in $NaClO_{3(aq)}$ .... +5

(iv) Explain why the above reaction is classified as a disproportionation reaction.

BECAUSE Cl_2 UNDERGOES BOTH OXIDATION AND REDUCTION

(b) $SO_{2(g)}$ is bubbled into an aqueous solution containing $Fe^{3+}_{(aq)}$ ions. Write half-equations and a balanced overall equation for the chemical changes occurring.

Name the oxidising agent and the reducing agent. [4 marks]

1 $SO_2(g) + H_2O(l) \rightarrow \cancel{H_2SO_3(aq)}$

1 $H_2O + \cancel{H_2SO_3(aq)} \rightarrow HSO_4^-(aq) + 3H^+(aq) + \cancel{2e^-}$

1 $\cancel{2e^-} + 2Fe^{3+}(aq) \rightarrow 2Fe^{2+}$

1/2 $SO_2(g) + 2Fe^{3+}(aq) + 2H_2O(l) \longrightarrow HSO_4^-(aq) + 3H^+(aq) + 2Fe^{2+}(aq)$

1/2 reducing agent   oxidising agent

Question 6

Many industrial chemical processes involve equilibrium reactions. For example, the gas carbon oxyfluoride (COF_2) decomposes to the gas carbon tetrafluoride (CF_4) and carbon dioxide.

(i) Write a balanced equation for the reaction, and hence write the expression for the equilibrium constant for this reaction. [2 marks]

1 $2COF_2(g) \rightleftharpoons CF_4(g) + CO_2(g)$

1 $K = \dfrac{[CF_4][CO_2]}{[COF_2]^2}$

The reaction is carried out at 200°C in a fixed-volume 10 litre container. Initially, there are 2.0 moles of carbon oxyfluoride gas in the flask. At equilibrium, 80% of the carbon oxyfluoride has decomposed. [3 marks]

(ii) Determine the value of the equilibrium constant.

$2COF_2(g) \longrightarrow CF_4(g) + CO_2(g)$

	$2COF_2$	CF_4	CO_2	
initial	2	0	0	1
equilibrium	2-1.6	0.8	0.8	1

$K = \dfrac{[CF_4][CO_2]}{[COF_2]^2} = \dfrac{0.8 \times 0.8}{(2-1.6)^2}$ 1

$= \underline{\underline{4}}$

(iii) The enthalpy change for the decomposition of carbon oxyfluoride is –24 kJ mol⁻¹. Explain the effect of an increase in temperature on the equilibrium constant. [1 mark]

$2COF_2(g) \rightarrow CF_4(g) + CO_2(g) + \text{heat}$

$K = \dfrac{[CF_4][CO_2]}{[COF_2]^2}$

IF HEAT APPLIED TO SYSTEM EQUILIBRIUM IS RE-ESTABLISHED BY AN INCREASE IN REVERSE REACTION

∴ $[COF_2]$ increases and $[CF_4]$ + $[CO_2]$ decreases

∴ New Equilibrium constant is smaller

Question 7

Use the information given below to identify the elements X and Y. Justify your choices.

The oxides have formulae of XO and YO_2.
XO is a solid which dissolves in water producing a basic solution.
YO_2 is a colourless, odourless gas which dissolves in water producing an acidic solution.
The element X is the third member of its group in the periodic table.

[3 marks]

1 X must form X^{2+} since XO exists

1 X must be Ca since this is the 3rd member of group 2 and CaO reacts with water to form $Ca(OH)_2(aq)$

1 Y is C since CO_2 is a colourless, odourless gas to form $H_2CO_3(aq)$ in water

Question 8

Consider the equilibrium reaction between hydrogen chloride gas and oxygen gas using a suitable catalyst:

$$4HCl(g) + O_2(g) \rightleftharpoons 2H_2O(g) + 2Cl_2(g) \qquad \Delta H = -280 \text{ kJ mol}^{-1}$$

Complete the table below by:

(a) indicating the change in the **amount** of hydrogen chloride gas (write either "more", "less" or "unchanged") after each of the following changes to conditions at equilibrium have been made.

(b) giving also a brief **reason** for your choice based on accepted chemical principles.

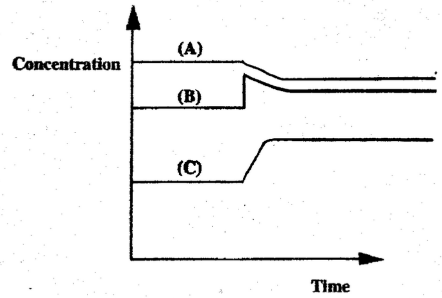
Change imposed on the system	Effect on **amount** of HCl(g) present
1. The temperature of the system is raised.	(a) Amount Increases (b) Reason. REACTION IS EXOTHERMIC ∴ IF HEAT GIVEN, SYSTEM REESTABLISHES EQUILIBRIUM BY REVERSE REACTION WHICH OPPOSES THE CHANGE
2. The pressure of the system is doubled by halving the volume of the reaction vessel.	(a) Amount Decreases (b) Reason. SYSTEM TRIES TO REDUCE PRESSURE BY LOWERING NUMBER OF PARTICLES. THEREFORE FORWARD REACTION (5 particles → 4 particles) is favoured
3. The surface area of the **catalyst** is increased	(a) NO CHANGE (b) Reason. CATALYST EFFECTS RATE OF REACTIONS NOT EQUILIBRIUM YIELDS OF SUBSTANCES

Question 9

Consider the gaseous system $A_{(g)} + B_{(g)} \rightleftharpoons C_{(g)}$. At equilibrium, a concentration/time graph would look like the graph below.

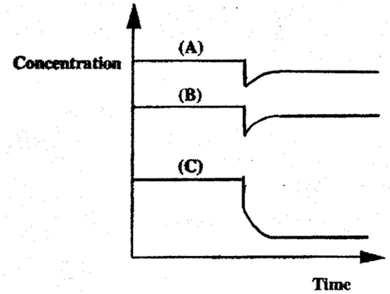
(a) Draw a concentration/time graph to reflect changes if the partial pressure of gas (B) is increased by adding gas (B) at constant volume.

[3 marks]



(b) Draw a concentration/time graph to reflect the changes if the total pressure is decreased by increasing the volume (temperature is held constant).

[3 marks]



Question 10

Consider the equilibrium:

$$BaSO_{4(s)} + H_2O \rightleftharpoons Ba^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

(a) Why is this described as a **dynamic equilibrium?**

[2 marks]

"Dynamic" because the forward and reverse reactions are always occurring.
"Equilibrium" because the forward and reverse reactions occur at the same rate rendering the macroscopic properties as constant.

(b) Given a small amount of $BaSO_{4(s)}$ containing radioactive barium, explain how this could be used in an experiment to show that the equilibrium above is dynamic.

[2 marks]

AT EQUILIBRIUM THE CONCENTRATION OF RADIOACTIVE $Ba^{2+}(aq)$ MAY BE DETERMINED
IF A NUMBER OF TESTS ARE PERFORMED AT VARIOUS TIMES DURING EQUILIBRIUM, THE CONCENTRATION OF RADIOACTIVE BARIUM IONS WILL ALWAYS BE DIFFERENT EVEN THOUGH EQUILIBRIUM OCCURS, THIS IS BECAUSE THE RADIOACTIVE BARIUM IS MOVING WITH NORMAL BARIUM BETWEEN SOLID BARIUM SULFATE AND THE IONS.
HENCE, DYNAMIC PROCESS

Question 11

What is the function of carbon in the CIP process?

[1 mark]

Adsorbs $[Au(CN)_2]^-$ FROM THE ORE PULP WITHIN THE ADSORPTION TANKS.

Question 12

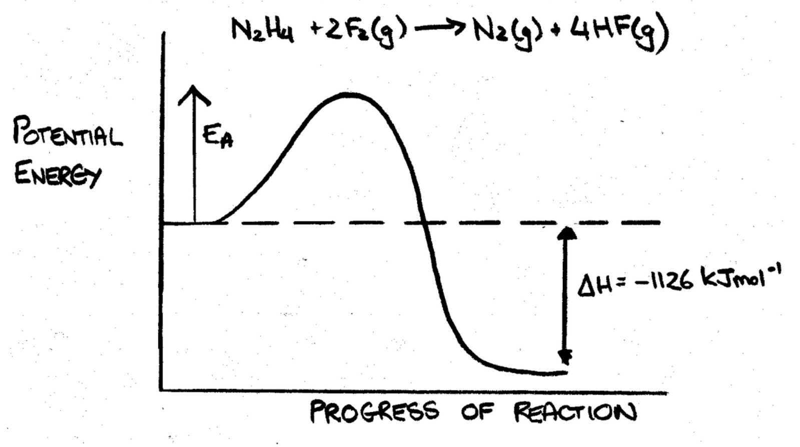
Hydrazine, N_2H_4, is sometimes used as a rocket fuel. If $F_{2(g)}$ is used as an oxidant, its heat of combustion is:

$$N_2H_4 + 2F_{2(g)} \rightarrow N_{2(g)} + 4HF_{(g)} \quad \Delta H = -1126 kJ\ mol^{-1}.$$

Draw a potential energy diagram for the above reaction.

Show all information on the diagram.

[3 marks]

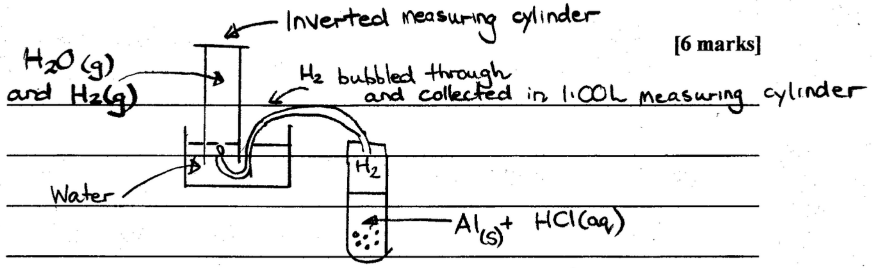


PART 3

Question 1

Maria designed an experiment to measure the molar volume of hydrogen gas. To do this she planned to dissolve a carefully weighed piece of pure aluminium in excess hydrochloric acid and collect the resulting hydrogen gas. Since a 1.00 L measuring cylinder is the largest she has at her disposal, determine the maximum mass of pure aluminium Maria should use. Assume the laboratory temperature is 25 °C and the gas is collected at 100.7 kPa. Water has a vapour pressure of 3.17 kPa at these conditions.

[6 marks]



P_{TOT} (measuring cylinder) = $p(H_2O(g)) + p(H_2(g))$

$\therefore p(H_2(g)) = p_{TOT} - p(H_2O(g))$

$= 100.7\text{ kPa} - 3.17\text{ kPa}$

$= 97.53\text{ kPa}$ ①

$\therefore P_1 = 97.53\text{ kPa}$ $P_2 = 101.3\text{ kPa}$

$V_1 = 1.00\text{ L}$ $V_2 = ?$ STP

$T_1 = 298\text{ K}$ $T_2 = 273\text{ K}$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}$$

$$= \frac{97.53 \times 1 \times 273}{101.3 \times 298}$$

$= 0.8820\text{ L}$ ①

Now $n = \frac{V}{22.4}$ (at STP)

$\therefore n = \frac{0.8820}{22.4} = 0.0394$ moles. ①

Now $2Al_{(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_3(aq) + 3H_2(g)$ ①

$n(Al) = \frac{2\,n(H_2)}{3} = \frac{2 \times 0.0394}{3} = 0.02625$ moles ①

$m(Al) = n(Al) \times M = 0.0262 \times 26.98 = 0.708\text{g}$ (3 sig fig) ①

Maximum Mass of Al is 0.708g

Question 2

Hydrogen sulfide gas (H_2S) readily dissolves in a sodium hydroxide solution to produce a sodium sulfide solution and water. In an experiment 4.92 g of $H_2S(g)$ is bubbled through 155 mL of 1.100 mol L^{-1} NaOH solution. What maximum mass of Na_2S could be produced in this experiment?

[8 marks]

$H_2S(g) + 2NaOH(aq) \longrightarrow Na_2S(aq) + 2H_2O(\ell)$ 1

$n(H_2S) = \frac{m}{M} = \frac{4.92}{34.086} = 0.14434$ moles 1

$n(NaOH) = cV = 1.100 \times 0.155 = 0.1705$ moles 1

Now 0.1705 moles of NaOH needs $\left(\frac{0.1705}{2}\right)$ moles of H_2S for total consumption. This equals 0.08525 moles of H_2S

We have 0.14434 moles of H_2S which is more than enough. Hence, H_2S is excess reagent and NaOH is limiting reagent 3

$n(Na_2S) = \frac{n(NaOH)}{2} = \frac{0.1705}{2} = 0.08525$ moles 1

$m(Na_2S) = n \times M = 0.08525 \times 78.05 = 6.65$ g 1

∴ Maximum mass of Na_2S is 6.65 g (3 sig figs)

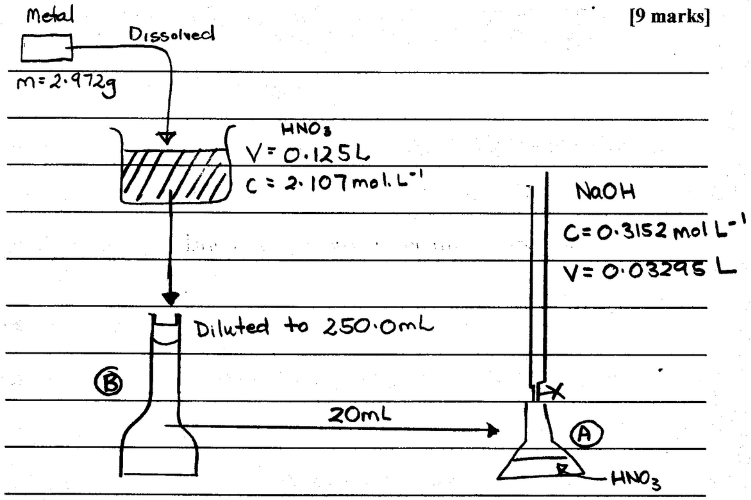
Question 3

A 2.972 g sample of an unknown metallic element was dissolved in 125.0 mL of 2.107 mol L^{-1} $HNO_3(aq)$. The reaction produced a salt with formula $M(NO_3)_4$.

$3M(s) + 16HNO_3(aq) \rightarrow 3M(NO_3)_4(aq) + 4NO(g) + 8H_2O(l)$

The remaining solution which was known to contain excess $HNO_3(aq)$ was diluted to 250.0 mL. A 20 mL sample of the diluted solution containing excess $HNO_3(aq)$ was titrated to equivalence using 32.95 mL of 0.3152 mol L^{-1} $NaOH(aq)$. What is the molar mass of the metal? Use your answer to identify the metal element.

[9 marks]



need Molar mass of Metal

1 $NaOH(aq) + HNO_3(aq) \longrightarrow NaNO_3(aq) + H_2O(\ell)$

1 $n(NaOH) = cV = 0.3152 \times 0.03295 = 0.0103858$ moles

1 $n(HNO_3 \text{ at } A) = n(NaOH) = 0.0103858$

1 $c(HNO_3 \text{ at } B) = \frac{n}{V} = \frac{0.0103858}{0.020} = 0.519292$ mol.L^{-1}

1 $n(HNO_3 \text{ at } B) = n(HNO_3$ remaining after metal addition)

$= cV = 0.519292 \times 0.250$

$= 0.129823$ moles.

1 $n(HNO_3$ before metal addition$) = cV = 2.107 \times 0.125$

$= 0.263375$ moles

(1) $n(HNO_3 \text{ reacting with metal})$

$= n(HNO_3 \text{ before metal addition})$

$- n(HNO_3 \text{ remaining after metal addition})$

$= 0.263375 - 0.129823$

$= 0.133552$ moles

(1) now $n(M) = \dfrac{3n(HNO_3)}{16} = \dfrac{3 \times 0.133552}{16}$

$= 0.025041$ moles

(1/2) Molar mass$(M) = \dfrac{m}{n} = \dfrac{2.972}{0.025041} = 118.68 \text{ g mol}^{-1}$

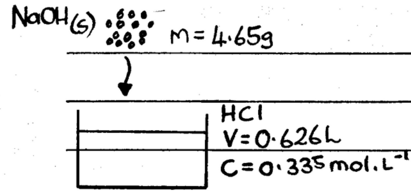
$\therefore$ Molar mass $= 118.7 \text{ g mol}^{-1}$ 119 (3 sig fig)

(1/2) The metal is Sn (Tin)

Question 4

A 4.65 g sample of pure NaOH*(s)* is added to 626 mL of 0.335 mol L^{-1} HCl*(aq)*. Assume the final volume of mixture is unchanged and determine the pH of the mixture when the reaction is complete.

[8 marks]



(1) $n(NaOH(s)) = \dfrac{m}{M} = \dfrac{4.65}{39.998} = 0.11626$

(1) $n(HCl) = cV = 0.335 \times 0.626 = 0.20971$

(1) $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(\ell)$

(2) 0.11626 mol NaOH needs 0.11626 mol HCl. We have 0.20971 mol of HCl which is more than enough. $\therefore$ HCl is excess reagent. NaOH is limiting reagent.

(1) $n(HCl \text{ after reaction}) = n(HCl \text{ original}) - n(HCl \text{ used})$

$= 0.20971 - 0.11626$

$= 0.09345$ moles

(1) $c(HCl \text{ after reaction}) = \dfrac{n}{V} = \dfrac{0.09345}{0.626} = 0.14928 \text{ mol L}^{-1}$

$c(H^+) = c(HCl \text{ after reaction}) = 0.14928 \text{ mol L}^{-1}$

(1) $pH = -\log(c(H^+)) = -\log(0.14928)$

$= 0.826$

Question 5

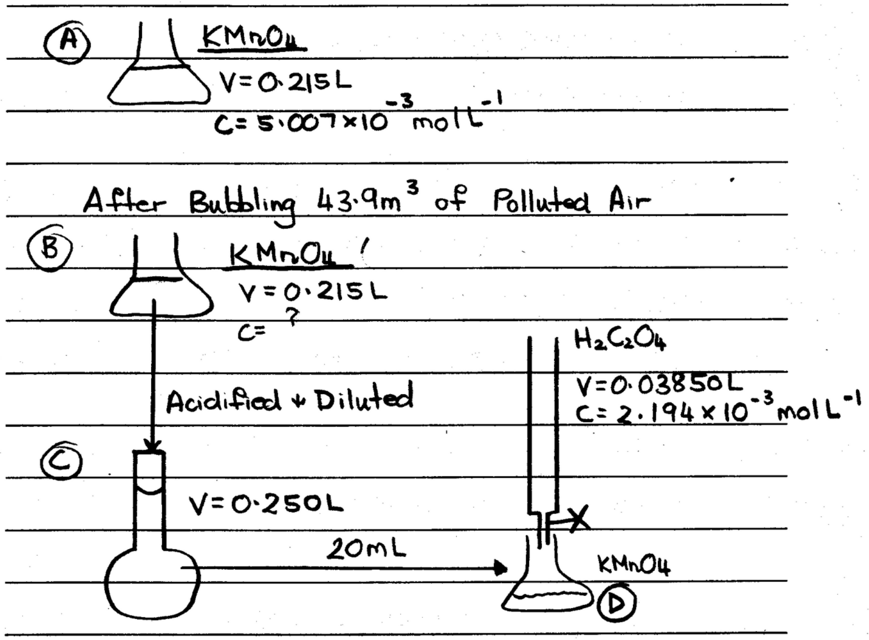
The concentration of the atmospheric pollutant sulfur dioxide (SO_2) can be found by bubbling air through a dilute $KMnO_4(aq)$ solution of known concentration.

$$5SO_2(g) + 2MnO_4^-(aq) + 2H_2O(l) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 4H^+(aq)$$

The concentration of the remaining $KMnO_4(aq)$ can be found by titration with standardised oxalic acid. This allows the amount of $KMnO_4$ reacting with sulfur dioxide to be found and thus its concentration in the air sample can be calculated. In such a procedure 43.9 m^3 of SO_2 polluted air was bubbled through 215.0 mL of 5.007 x 10^{-3} mol L^{-1} $KMnO_4(aq)$. The unreacted $KMnO_4$ was acidified and diluted to a volume of 250.0 mL. 20.00 mL samples of this $KMnO_4$ solution were titrated to equivalence with 38.50 mL of 2.194 x 10^{-3} mol L^{-1} oxalic acid solution.
What is the concentration of the pollutant $SO_2(g)$ in ppm if the air has a density of 1.18 kg m^{-3}?

[12 marks]

~~Bub~~ Before Bubbling



Need ppm of $SO_2(g)$ in air

density of air = 1.18 Kg m^{-3}

① $n(H_2C_2O_4) = cV = 2.194\times10^{-3} \times 0.03850 = 8.4469\times10^{-5}$

② $5H_2C_2O_4(aq) \longrightarrow \cancel{10e^-} + 10CO_2(g) + \cancel{10H^+}(aq)$

$2MnO_4^-(aq) + \cancel{16}H^+(aq) + \cancel{10e^-} \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l)$

$5H_2C_2O_4(aq) + 2MnO_4^-(aq) + 6H^+ \longrightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(l)$

① $n(MnO_4^- \text{ at } D) = \dfrac{2\,n(H_2C_2O_4)}{5} = \dfrac{2\times8.4469\times10^{-5}}{5}$

$= 3.37876\times10^{-5}$ mol

① $c(MnO_4^- \text{ at } C) = \dfrac{n}{V} = \dfrac{3.37876\times10^{-5}}{0.020} = 1.68938\times10^{-3}$ mol L^{-1}

① $n(MnO_4^- \text{ at } C) = cV = 1.68938\times10^{-3}\times0.250 = 4.22345\times10^{-4}$

$n(MnO_4^- \text{ at } B) = 4.22345\times10^{-4}$

① $n(MnO_4^- \text{ at } A) = cV = 5.007\times10^{-3}\times0.215 = 1.076505\times10^{-3}$

① $n(MnO_4^- \text{ reacting with } SO_2) = n(MnO_4^- \text{ at } A) - n(MnO_4^- \text{ at } B)$

$= (1.076505\times10^{-3}) - (4.22345\times10^{-4})$

$= 6.5416\times10^{-4}$ moles

① $n(SO_2(g)) = \dfrac{5\,n(MnO_4^-)}{2} = \dfrac{5\times6.5416\times10^{-4}}{2} = 1.6354\times10^{-3}$

① $m(SO_2(g)) = nM = 1.6354\times10^{-3}\times64.07 = 0.10478$g

$= 104.78$ mg

① $m(\text{air}) = \text{density} \times \text{Volume} = 1.18\times43.9 = 51.802$ Kg

① $ppm(SO_2) = \dfrac{m(SO_2) \text{ in mg}}{m(\text{air}) \text{ in Kg}} = \dfrac{104.78}{51.802} = 2.02$ ppm

∴ concentration of SO_2 = 2.02 ppm

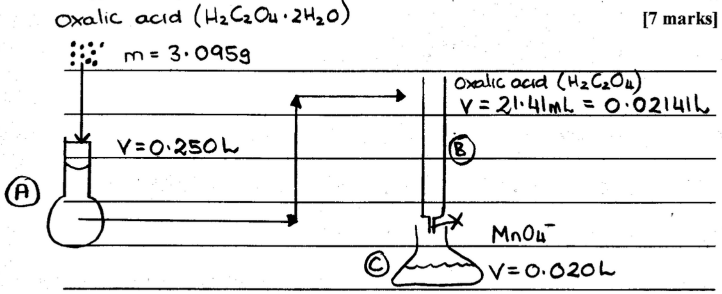
Question 6

In a laboratory procedure two students need to make a secondary standard solution of approximately 0.04 mol L^{-1} $KMnO_4(aq)$. To do this they dissolved 3 g of solid $KMnO_4$ in 500 mL of distilled water. The solution was then boiled, filtered through glass wool and stored in a dark bottle away from light.
A second solution of oxalic acid was prepared by dissolving 3.095 g of $H_2C_2O_4.2H_2O(s)$ in water and making its volume up to 250.0 mL in a volumetric flask. Finally, the oxalic acid solution was added to a burette and titrated into 20.00 mL samples of the potassium permanganate solution. The burette readings for the $H_2C_2O_4$ solution were as follows:

					IGNORE			
Final volume (mL)	22.25	21.38	22.91	21.45	24.75	22.00	22.38	21.40
Initial volume (mL)	0.87		1.46		2.75		0.98	

Using this information, determine the actual concentration of the approximately 0.04 mol L^{-1} $KMnO_4$ solution. (1)

[7 marks]



(1) $n(H_2C_2O_4 \cdot 2H_2O) = \frac{m}{M} = \frac{3.095}{126.02} = 0.0245595$

$\therefore n(H_2C_2O_4) = 0.0245595$

(1) $c(H_2C_2O_4 \text{ at } A) = \frac{n}{V} = \frac{0.0245595}{0.250} = 0.0982383 \text{ mol.L}^{-1}$

(1) $n(H_2C_2O_4 \text{ added from } B) = cV = 0.0982383 \times 0.02141$

$= 2.1032836 \times 10^{-3}$ moles

(1) $5H_2C_2O_4(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O$

(1) $n(MnO_4^- \text{ at } C) = \frac{2\, n(H_2C_2O_4)}{5} = \frac{2 \times 2.1032836 \times 10^{-3}}{5}$

$= 8.4131344 \times 10^{-4}$ moles

$c(MnO_4^-) = \frac{n}{V} = \frac{8.4131344 \times 10^{-4}}{0.020}$ (1)

$= \underline{\underline{0.0421 \text{ mol.L}^{-1}}}$ (3 sig fig)