Solutions to Volumetric Analysis & Redox Tutorial

(A) Discussion Questions (Volumetric Analysis)

Acid-Base Titration

1. 45 cm^3 of concentrated aqueous NH_3 was diluted to 250 cm^3 solution labelled as **FA 1**. Given that 10.0 cm^3 of **FA 1** required 23.20 cm^3 of 0.18 mol dm^{-3} of HNO_3 for complete neutralisation, calculate the concentration of the concentrated aqueous NH_3.

[2.32 mol dm^{-3}]

Eqn for rxn: $NH_3 + HNO_3 \rightarrow NH_4NO_3$

n_{NH_3} in 10.0 cm^3 FA 1 = n_{HNO_3} used = $0.18 \times \frac{23.20}{1000} = 4.176 \times 10^{-3}$ mol

n_{NH_3} in 250 cm^3 of FA 1 = $4.176 \times 10^{-3} \times \frac{250}{10} = 0.1044$ mol

$[NH_3]_{\text{conc soln}} = 0.1044 / (\frac{45}{1000}) = \underline{2.32\text{ mol dm}^{-3}}$

2. A solution of a dibasic acid contains 7.30 g dm^{-3} of $HOOC–(CH_2)_{\boldsymbol{n}}–COOH$. 20.0 cm^3 of this acid solution was titrated with 25.00 cm^3 of NaOH(aq) containing 1.36 g of hydroxide ion per dm^3. Calculate

(a) the relative molecular mass of the acid;

[146.0]

Eqn for rxn: $HO_2C–(CH_2)_n–CO_2H + 2NaOH \rightarrow NaO_2C–(CH_2)_n–CO_2Na + 2H_2O$

$[OH^-] = \frac{1.36}{17.0} = 0.08$ mol dm^{-3}

Amount of NaOH reacted = $\frac{25}{1000} \times 0.08 = 2.00 \times 10^{-3}$ mol

Amount of dibasic acid reacted = $\frac{1}{2} \times 2.00 \times 10^{-3} = 1.00 \times 10^{-3}$ mol

$$[\text{dibasic acid}] = \frac{1.00 \times 10^{-3}}{20/1000} = 0.0500 \text{ mol dm}^{-3}$$

M_r of dibasic acid = $\frac{7.30}{0.0500} = \underline{146.0}$

(b) the value of $\boldsymbol{n}$ in the formula.

[4]

M_r of dibasic acid = $2(12.0 + 16.0 \times 2 + 1.0) + \boldsymbol{n}(12.0 + 2.0) = 146.0$

$\boldsymbol{n} = 4.07 = \underline{4}$ (to nearest integer)

3. **FA 8** solution contains 20.2 g of the acid HZO_4 per dm^3 of solution.
FA 9 is a 0.100 mol dm^{-3} NaOH solution. In a titration, 20.0 cm^3 of **FA 8** solution reacted with 21.05 cm^3 of **FA 9** solution. Calculate the relative atomic mass of element **Z** and identify **Z**.

[126.9; I]

NaOH ≡ HZO_4 (since HZO_4 is monobasic with 1 H^+ per molecule)

Amount of NaOH = $0.100 \times \frac{21.05}{1000} = 2.105 \times 10^{-3}$ mol

Amount of HZO_4 in 20.0 cm^3 FA 8 = 2.105×10^{-3} mol

$[HZO_4] = \frac{2.105\times10^{-3}}{20/1000} = 0.1053$ mol dm^{-3}

M_r of $HZO_4 = \frac{20.2}{0.1053} = 191.9$

A_r of **Z** = $191.9 - 1.0 - 4 \times 16.0 = \underline{126.9}$

Z is Iodine.

4. Washing soda has the formula $Na_2CO_3{\cdot}nH_2O$. A mass of 1.43 g of washing soda was made up to 250 cm^3 with water. 25.0 cm^3 of this solution was neutralised by 20.00 cm^3 of 0.050 mol dm^{-3} dilute hydrochloric acid. The equation for the reaction is:

$$Na_2CO_3(aq) + 2HCl\,(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$$

Find the value of n, and hence the chemical formula of the washing soda.
[The equation for the dissolution of washing soda in water is: $Na_2CO_3{\cdot}nH_2O(s) \rightarrow Na_2CO_3(aq) + nH_2O$]

[10]

Amount of HC*l* = $0.050 \times \frac{20}{1000} = 1.00 \times 10^{-3}$ mol

Amount of Na_2CO_3 in 25.0 cm^3 solution = ½ × $1.00 \times 10^{-3} = 5.00 \times 10^{-4}$ mol

Amount of Na_2CO_3 in 250 cm^3 solution = $5.00 \times 10^{-4} \times \frac{250}{25} = 5.00 \times 10^{-3}$ mol

Mass of Na_2CO_3 in 250 cm^3 solution = $5.00 \times 10^{-3} \times (2\times23.0 + 12.0 + 3\times16.0)$
= 0.530 g

∴ mass of H_2O in 1.43 g washing soda = 1.43 – 0.530 = 0.900 g

∴ amount of H_2O in 1.43 g washing soda = $\frac{0.900}{18.0} = 0.0500$ mol

$$\frac{n_{H_2O}}{n_{Na_2CO_3}} = \frac{0.0500}{5.00 \times 10^{-3}} = 10$$

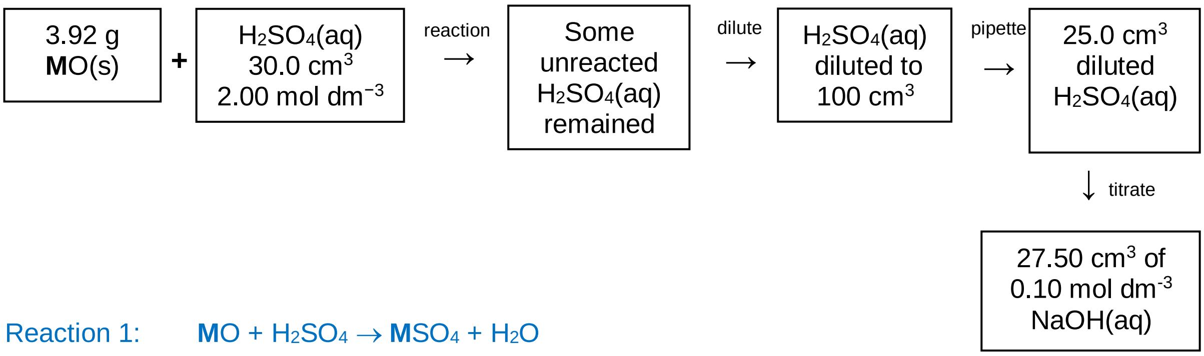
10 mol of water combine with 1 mol of Na_2CO_3

∴n = 10; formula of washing soda is $Na_2CO_3{\cdot}\underline{10}H_2O$.

Back Titration

5. 3.92 g of an oxide of formula **M**O was completely dissolved in 30.0 cm^3 of 2.00 mol dm^{-3} sulfuric acid. The resulting solution was made up to 100 cm^3. 25.0 cm^3 of this solution was neutralised by 27.50 cm^3 of 0.10 mol dm^{-3} NaOH(aq).
What is the relative atomic mass of **M**? Identify the metal.

[55.9; Fe]



Reaction 1: $MO + H_2SO_4 \rightarrow MSO_4 + H_2O$

Reaction 2: $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

Total starting amount of H_2SO_4 (S) $= 2.00 \times \frac{30}{1000} = 0.0600$ mol

For Reaction 2, Amount of NaOH reacted $= 0.10 \times \frac{27.50}{1000} = 2.75 \times 10^{-3}$ mol

Amount of H_2SO_4 in the 25 cm^3 solution $= 2.75 \times 10^{-3} \times \frac{1}{2} = 1.375 \times 10^{-3}$ mol

Amount of H_2SO_4 in the 100 cm^3 solution (E) $= 1.375 \times 10^{-3} \times \frac{100}{25} = 5.50 \times 10^{-3}$ mol

For back titration, $R = S - E$

Amount of H_2SO_4 that reacted with **M**O (R) $= 0.0600 - 5.50 \times 10^{-3} = 0.0545$ mol

Amount of **M**O = 0.0545 mol

M_r of **M**O $= \frac{3.920}{0.0545} = 71.9$

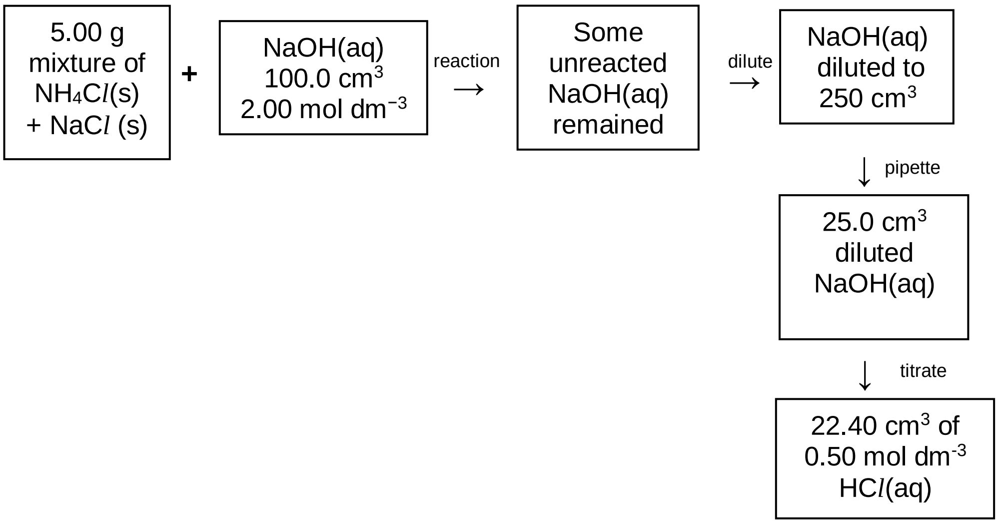
A_r of **M** $= 71.9 - 16.0 = 55.9$

M is Fe (A_r of 55.8).

6. 5.00 g of ammonium chloride contaminated with sodium chloride was boiled with 100.0 cm^3 of 2 mol dm^{-3} NaOH solution until no ammonia was evolved. The residual solution was made up to 250 cm^3 with water. 25.0 cm^3 of this solution required 22.40 cm^3 of 0.50 mol dm^{-3} HCl for neutralisation.

(a) What was the mass of sodium chloride in the ammonium chloride sample?

[0.292 g]



Reaction 1: $NH_4Cl + NaOH \rightarrow NH_3 + NaCl + H_2O$
Reaction 2: $NaOH + HCl \rightarrow NaCl + H_2O$

For reaction 2, amount of HCl = $\frac{22.40}{1000} \times 0.5 = 0.0112$ mol
Amount of unreacted NaOH in 25 cm^3 = 0.0112 mol
Amount of unreacted NaOH in 250 cm^3 (E) = $0.0112 \times \frac{250}{25} = 0.112$ mol

Initial amount of NaOH added (S) = $\frac{100}{1000} \times 2.0 = 0.2$ mol

For back titration, $R = S - E$
Amount of NaOH reacted with NH_4Cl (R) = 0.2 – 0.112 = 0.088 mol
Mass of NH_4Cl = 0.088 × 53.5 = 4.708 g
Mass of $NaCl$ = 5.00 – 4.708 = 0.292 g

(b) Hence, calculate the percentage by mass of sodium chloride in the ammonium chloride sample.

[5.84%]

Percentage by mass of $NaCl$ in the sample = $\frac{0.292}{5.00} \times 100\% = 5.84\%$

Redox Titration

7. **FA 4** contains 10.0 g of Fe^{2+} and Fe^{3+}, dissolved in 500 cm^3 of solution. **FA 5** contains 0.015 mol dm^{-3} $KMnO_4$.

In an experiment, 10.0 cm^3 of solution of **FA 4** was pipetted into a titration flask. Excess dilute sulfuric acid was added and the mixture titrated with **FA 5**. 26.65 cm^3 of **FA 5** was required to reach end-point.

(a) Calculate the mass of Fe^{2+} in 500 cm^3 of **FA 4**.

[5.58 g]

$MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$

Amount of $KMnO_4$ reacted = $0.015 \times \frac{26.65}{1000} = 3.998 \times 10^{-4}$ mol

Amount of Fe^{2+} in 10.0 cm^3 **FA 4** = $5 \times 3.998 \times 10^{-4} = 1.999 \times 10^{-3}$ mol

Amount of Fe^{2+} in 500 cm^3 **FA 4** = $1.999 \times 10^{-3} \times \frac{500}{10} = 0.09995$ mol

Mass of Fe^{2+} in 500 cm^3 **FA 4** = $0.09995 \times 55.8 = \underline{5.58\text{ g}}$

(b) Calculate the percentage by mass of Fe^{3+} ions in **FA 4**.

[44.2 %]

% by mass of Fe^{3+} in **FA 4** = $\frac{10.00 - 5.58}{10} \times 100 = \underline{44.2\ \%}$

8. Chlorate(V) ions, ClO_3^-, act as an oxidising agent according to the following half-equation:

$$ClO_3^- + 6H^+ + 6e^- \rightarrow Cl^- + 3H_2O$$

(a) Chlorate(V) ions are reduced by Fe^{2+} ions. Write a balanced equation for the reaction between Fe^{2+} and ClO_3^-.
[Hint: Write the half equation for Fe^{2+} oxidized to Fe^{3+} before combining the 2 half-equations]

$ClO_3^- + 6H^+ + 6Fe^{2+} \rightarrow Cl^- + 3H_2O + 6Fe^{3+}$

(b) In an experiment, 25.0 cm^3 of a solution of $KClO_3$ was titrated with a solution of iron(II) sulfate containing 6.72 g dm^{-3} of Fe^{2+}. In the titration, 20.0 cm^3 of iron(II) sulfate was used for the reaction.

(i) What other chemical is required for the titration? Suggest a reactant that can be added to the reaction mixture for the titration.

acid to provide H^+; dilute H_2SO_4

Note: HNO_3 and HCl are not suitable acids as they will take part in redox reaction. HNO_3 is an oxidising agent (NO_3^- can undergo reduction) and Cl^- can undergo oxidation.

(ii) Calculate the concentration of $KClO_3$ in mol dm^{-3}.

[0.0161 mol dm^{-3}]

$[Fe^{2+}]$ in mol dm^{-3} = $\frac{6.72}{55.8}$ = 0.1204 mol dm^{-3}

Amount of Fe^{2+} reacted = 0.1204 × $\frac{20}{1000}$ = 2.409×10^{-3} mol

Amount of $KClO_3$ in 25.0 cm^3 solution = $\frac{1}{6}$ × 2.409×10^{-3} = 4.014×10^{-4} mol

$[KClO_3] = \frac{4.014 \times 10^{-4}}{25.0/1000}$ = $\underline{0.0161 \text{ mol dm}^{-3}}$

(iii) Calculate the mass of potassium in 1 dm^3 of the $KClO_3$ solution.

[0.630 g]

Amount of K in 1 dm^3 of solution = 0.0161 mol

Mass of K in 1 dm^3 of solution = 39.1 × 0.0161 = $\underline{0.630 \text{ g}}$

9. **FA 2** is a solution containing 6.70 g dm^{-3} of a metal ethanedioate, $\mathbf{M}_2C_2O_4$.
FA 3 is a solution containing 0.010 mol dm^{-3} of manganate(VII) ions, MnO_4^-.
10.0 cm^3 of **FA 2** was pipetted into a conical flask and dilute sulfuric acid was added in excess. The mixture was then titrated with **FA 3** and the end-point was reached when 20.00 cm^3 of **FA 3** had been added.

(a) **(i)** How do you know when the end-point was reached?

Colour change from colourless to first permanent pale pink colour (after adding 1 excess drop of MnO_4^-)

(ii) What other condition was required for the titration to be successful?

Conduct the titration in a <u>hot water bath</u> / <u>heat the</u> reaction mixture.

The reaction occurs too slowly at room temperature due to repulsion between like charges of the two reacting species, $C_2O_4^{2-}$ and MnO_4^- (may result in lower than actual titre). Heat is required to overcome the high activation energy barrier.

(b) Calculate the relative formula mass of $\mathbf{M}_2C_2O_4$ and the relative atomic mass of **M**.

[134.0; 23.0]

Overall eqn: $2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2$

Amount of $KMnO_4$ reacted = 0.010 × $\frac{20}{1000}$ = 2.00×10^{-4} mol

Amount of $C_2O_4^{2-}$ in 10.0 cm^3 **FA 2** = $\frac{5}{2}$ × 2.00×10^{-4} = 5.00×10^{-4} mol

$[C_2O_4^{2-}]$ in **FA 2** = $\frac{5.00 \times 10^{-4}}{10.0/1000}$ = 0.0500 mol dm^{-3}

M_r of $M_2C_2O_4$ = $\frac{6.70}{0.0500}$ = 134.0

A_r of M = ½ [134.0 – (2 × 12.0 + 16.0 × 4)] = $\underline{23.0}$

10. Hydroxylamine, NH_2OH, can be oxidised to nitrogen according to the following half-equation:

$$2NH_2OH \rightarrow N_2 + 2H_2O + 2H^+ + 2e^-$$

Hydroxylamine can also be oxidised to dinitrogen oxide, N_2O, according to the following half-equation:

$$2NH_2OH \rightarrow N_2O + H_2O + 4H^+ + 4e^-$$

(a) In an experiment, 40.00 cm^3 of 0.05 mol dm^{-3} $NH_2OH(aq)$ was reacted with 10.0 cm^3 of 0.40 mol dm^{-3} $Fe^{3+}(aq)$. In the reaction, the Fe^{3+} was reduced to Fe^{2+}. Calculate the amount of Fe^{3+} that reacted with 1 mole of NH_2OH.

[2]

[R] $Fe^{3+} + e^- \rightarrow Fe^{2+}$

Amount of Fe^{3+} reacted $= 0.40 \times \frac{10}{1000} = 4.00 \times 10^{-3}$ mol

Amount of NH_2OH reacted $= 0.05 \times \frac{40}{1000} = 2.00 \times 10^{-3}$ mol

$\therefore$ $\underline{2}$ mol of Fe^{3+} reacted with 1 mol of NH_2OH

(b) Hence deduce whether the NH_2OH is oxidised to N_2 or N_2O by the Fe^{3+} and write a balanced ionic equation for the reaction.

If NH_2OH was oxidised to N_2:

Overall eqn will be $2NH_2OH + 2Fe^{3+} \rightarrow N_2 + 2H_2O + 2H^+ + 2Fe^{2+}$

1 mol of Fe^{3+} should react with 1 mol of NH_2OH; <u>inconsistent</u> with answer in obtained in (a).

If NH_2OH is oxidised to N_2O:

Overall eqn will be $2NH_2OH + 4Fe^{3+} \rightarrow N_2O + H_2O + 4H^+ + 4Fe^{2+}$

2 mol of Fe^{3+} should react with $\underline{1}$ mol of NH_2OH; <u>consistent</u> with answer in obtained in (a).

$\therefore$ NH_2OH was oxidised to N_2O in the reaction.

11. A solution of vanadium ions, $VO_2^+(aq)$ is blue. The $VO_2^+(aq)$ are reduced by zinc powder and acid to violet $V^{z+}(aq)$ ions.

(a) What is the oxidation number of the vanadium in VO_2^+?

[+5]

Let O.S. of V be a
a + 2(–2) = +1
a = +5

(b) The $V^{z+}(aq)$ can be oxidised quantitatively to $VO_2^+(aq)$ by acidified $MnO_4^-(aq)$. It was found experimentally that 20.0 cm^3 of 0.10 mol dm^{-3} $V^{z+}(aq)$ reacted with 24.00 cm^3 of 0.050 mol dm^{-3} $MnO_4^-(aq)$.

(i) Calculate the number of moles of $V^{z+}(aq)$ that react with 1 mol of acidified $MnO_4^-(aq)$.

$[\frac{5}{3}]$

Amount of $KMnO_4$ reacted = $0.050 \times \frac{24}{1000} = 1.20 \times 10^{-3}$ mol

Amount of V^{z+} reacted = $0.10 \times \frac{20}{1000} = 2.00 \times 10^{-3}$ mol

Amt of $KMnO_4$: Amount of V^{z+} = $1.20 \times 10^{-3} : 2.00 \times 10^{-3}$
= 3 : 5

$\therefore$ Amount of V^{z+} that reacted with 1 mol of MnO_4^- = $\frac{5}{3}$ mol

(ii) Hence deduce the number of moles of electrons transferred per mole of $V^{z+}(aq)$ in the reaction and value of **z** on V^{z+}.

[2]

[R]: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

During reduction, one mole of MnO_4^- gains $5e^-$;

During oxidation, $\frac{5}{3}$ mole of V^{z+} loses $5e^-$

$\therefore$ 1 mole of V^{z+} loses $3e^-$

$\therefore$ no of electrons transferred (lost) per mole of V^{z+} is 3 .

O.S. of V in VO_2^+ = +5
During oxidation, $V^{z+} \rightarrow VO_2^+$ after losing 3 e^-

Oxidation number of V **increases** by **3 units** from **+2** (in V^{z+}) to **+5** (in VO_2^+)

Hence z = 2

12. A 25.0 cm^3 sample of 0.0210 mol dm^{-3} potassium peroxodisulfate(VI), $K_2S_2O_8$, was treated with an excess of potassium iodide. The iodine liberated reacted with 21.00 cm^3 of 0.0500 mol dm^{-3} thiosulfate.

(a) Calculate the amount of $S_2O_8^{2-}$ used.

[5.25×10^{-4} mol]

Amount of $S_2O_8^{2-}$ used $= 0.0210 \times \frac{25}{1000} = 5.25 \times 10^{-4}$ mol

(b) Calculate the amount of I^- that reacted with 1 mol of $S_2O_8^{2-}$.

[2 mol]

Reaction 1: $S_2O_8^{2-} + 2I^- \rightarrow\ ? + I_2$ (unbalanced)

The iodine liberated is then reacted with thiosulfate.

Reaction 2: $2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$

Amount of $S_2O_3^{2-} = 0.0500 \times \frac{21.00}{1000} = 1.05 \times 10^{-3}$ mol

Amount of I_2 in Reaction 2 $= \frac{1}{2} \times 1.05 \times 10^{-3}$ mol $= 5.25 \times 10^{-4}$ mol

= Amt of I_2 formed in Reaction1

Amount of I^- reacted in 1st rxn to form $I_2 = 2 \times 5.25 \times 10^{-4} = 1.05 \times 10^{-3}$ mol

$$\frac{\text{Amount of } I^-}{\text{Amount of } S_2O_8^{2-}} = \frac{1.05 \times 10^{-3}}{5.25 \times 10^{-4}} = \frac{2}{1}$$

$\therefore$ 2 mol of I^- reacted with 1 mol of $S_2O_8^{2-}$

(c) Suggest a likely equation for the reaction between $K_2S_2O_8$ and KI, given that $S_2O_8^{2-}$ is reduced to SO_4^{2-}.

$S_2O_8^{2-} + 2I^- \rightarrow 2SO_4^{2-} + I_2$

13. A solution contained 5.53 g dm^{-3} of a metal iodate(VII), $\mathbf{M}IO_4$. 25.0 cm^3 of the solution was added to excess acidified potassium iodide. The iodate(VII) ions reacted with the iodide according to the following equation:

$$IO_4^- + 7I^- + 8H^+ \rightarrow 4I_2 + 4H_2O$$

The iodine liberated was then titrated with 0.120 mol dm^{-3} $S_2O_3^{2-}$, where starch was added only when the solution turned pale yellow. It was found that 33.30 cm^3 of $S_2O_3^{2-}$ was required for the dark-blue colour to be discharged.

(a) Explain why it is necessary for the potassium iodide to be added in excess.

To ensure all IO_4^- reacted completely to give I_2.
(or else its amount determined by titrating with $S_2O_3^{2-}$ will be inaccurate)

(b) Why was the starch indicator not added at the beginning of the titration of iodine with thiosulfate?

high $[I_2]$ at start of titration, some of the I_2 forms an insoluble complex with starch.
Hence, not all the I_2 liberated reacts with $S_2O_3^{2-}$, titre value is lesser than actual.
$\therefore [IO_4^-]$ calculated is thus inaccurate

(c) Calculate the relative atomic mass of element **M**. Suggest the identity of **M**.

[85.9; Rb]

Reaction 1: $IO_4^- + 7I^- + 8H^+ \rightarrow 4I_2 + 4H_2O$
Reaction 2: $2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$

Amount of $S_2O_3^{2-} = 0.120 \times \frac{33.30}{1000} = 3.996 \times 10^{-3}$ mol

Amount of $I_2 = \frac{1}{2} \times 3.996 \times 10^{-3} = 1.998 \times 10^{-3}$ mol

Amount of IO_4^- in 25.0 cm^3 solution $= \frac{1}{4} \times 1.998 \times 10^{-3} = 4.995 \times 10^{-4}$ mol

$[\mathbf{M}IO_4] = \frac{4.995 \times 10^{-4}}{25.0/1000} = 0.01998$ mol dm^{-3}

M_r of $\mathbf{M}IO_4 = \frac{5.53}{0.01998} = 276.8$

A_r of $\mathbf{M} = 276.8 - 126.9 - 4 \times 16.0 = 85.9$

M is Rb.

(B) Discussion Questions (Redox Practice)

1 Nitrogen dioxide reacts with iodine under acidic conditions according to the equation shown.

$$NO_2 + 2I^- + 2H^+ \rightarrow NO + I_2 + H_2O$$

How many moles of electrons are gained by one mole of the nitrogen dioxide?

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

Ans: (B)

Half equation: $2I^- \rightarrow I_2 + 2e^-$

2 moles of e^- will be gained by 1 mole of NO_2.

2 Write the half-equations and balance them in the respective mediums.

(a) **$As_2O_3(s) + VO_2^+(aq) \rightarrow As_2O_5(s) + VO^{2+}(aq)$ (in acidic medium)**

Half eqn: [O] $As_2O_3 + 2H_2O \rightarrow As_2O_5 + 4H^+ + 4e^-$ ------------ (1)

[R] $VO_2^+ + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$ ------------------ (2) x4

Overall eqn: $As_2O_3(s) + 4VO_2^+(aq) + 4H^+(aq) \rightarrow As_2O_5(s) + 4VO^{2+}(aq) + 2H_2O(l)$

(b) **$H_2O_2(aq) + Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + H_2O(l)$ (in acidic medium)**

Half eqn: [R] $H_2O_2 + 2H^+ + 2e^- \rightarrow H_2O + H_2O$ ------------ (1)

[O] $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$ ----------------------------- (2)

Overall eqn: $H_2O_2(aq) + Sn^{2+}(aq) + 2H^+(aq) \rightarrow Sn^{4+}(aq) + 2H_2O(l)$

(c) **$Cl_2(aq) \rightarrow Cl^-(aq) + ClO_3^-(aq)$ (in alkaline medium)**

Half eqn: [R] $Cl_2 + 2e^- \rightarrow 2Cl^-$--- (1) x5

[O] $Cl_2 + 12OH^- \rightarrow 2ClO_3^- + 10e^- + 6H_2O$ ---------------- (2)

Overall eqn: $6Cl_2(aq) + 12OH^-(aq) \rightarrow 10Cl^-(aq) + 2ClO_3^-(aq) + 6H_2O(l)$

Simplified eqn: $3Cl_2(aq) + 6OH^-(aq) \rightarrow 5Cl^-(aq) + ClO_3^-(aq) + 3H_2O(l)$

(d) **$MnO_4^{2-}(aq) \rightarrow MnO_2(s) + MnO_4^-(aq)$ (in alkaline medium)**

Half eqn: [R] $MnO_4^{2-} + 2H_2O + 2e^- \rightarrow MnO_2 + 4OH^-$--------------------- (1)

[O] $MnO_4^{2-} \rightarrow MnO_4^- + e^-$ -- (2) x2

Overall eqn: $3MnO_4^{2-}(aq) + 2H_2O(l) \rightarrow MnO_2(s) + 2MnO_4^-(aq) + 4OH^-(aq)$

3 Write the half-equations and hence the full balanced equation for each reaction below.

(a) Sodium sulfite, Na_2SO_3, reduces an orange acidified solution of dichromate(VI) ions, $Cr_2O_7^{2-}$, to green solution containing Cr^{3+} ions. Sulfate ions are formed in the reaction.

Half eqn: [R] $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ ------------------ (1)

[O] $SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-$ ------------------------- (2) x3

Overall eqn: $Cr_2O_7^{2-} + 3SO_3^{2-} + 8H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-} + 4H_2O$

(b) When copper is added to concentrated nitric acid, the solution becomes pale blue due to Cu^{2+}(aq), & brown fumes of nitrogen dioxide are produced.

Half eqn: [O] $Cu \rightarrow Cu^{2+} + 2e^-$ ------------------------------------ (1)

[R] $NO_3^- + 2H^+ + e^- \rightarrow NO_2 + H_2O$ ------------------- (2) x2

Overall eqn: $Cu + 2NO_3^- + 4H^+ \rightarrow Cu^{2+} + 2NO_2 + 2H_2O$

(c) When potassium iodide (KI) is added to acidified hydrogen peroxide, brown I_2 solution is formed.

Half eqn: [O] $2I^- \rightarrow I_2 + 2e^-$ ------------------------------------ (1)

[R] $H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$ -------------------- (2)

Overall eqn: $H_2O_2 + 2I^- + 2H^+ \rightarrow I_2 + 2H_2O$

(d) When heated, phosphorus acid, H_3PO_3, gives phosphoric acid, H_3PO_4, and phosphine gas, PH_3.

Half eqn: [O] $H_3PO_3 + H_2O \rightarrow H_3PO_4 + 2H^+ + 2e^-$ ---------------- (1) x3

[R] $H_3PO_3 + 6H^+ + 6e^- \rightarrow PH_3 + 3H_2O$ ------------------- (2)

Overall eqn: $4H_3PO_3 \rightarrow 3H_3PO_4 + PH_3$

(e) When zinc is added to silver nitrate solution, $AgNO_3$(aq), silver solid forms on the surface of zinc. Zn^{2+} ions are formed.

Half eqn: [O] $Zn \rightarrow Zn^{2+} + 2e^-$ ------------------------------------ (1)

[R] $Ag^+ + e^- \rightarrow Ag$ -------------------------------------- (2) x2

Overall eqn: $Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$