# Solutions to Volumetric Analysis & Redox Tutorial

## (A) Discussion Questions (Volumetric Analysis)

## Acid-Base Titration

 45 cm<sup>3</sup> of concentrated aqueous NH<sub>3</sub> was diluted to 250 cm<sup>3</sup> solution labelled as FA 1. Given that 10.0 cm<sup>3</sup> of FA 1 required 23.20 cm<sup>3</sup> of 0.18 mol dm<sup>-3</sup> of HNO<sub>3</sub> for complete neutralisation, calculate the concentration of the concentrated aqueous NH<sub>3</sub>.

[2.32 mol dm<sup>-3</sup>]

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

$$\eta_{NH_3}$$
 in 10.0 cm<sup>3</sup> FA 1 =  $\eta_{HNO_3}$  used = 0.18 x  $\frac{23.20}{1000}$  = 4.176 x 10<sup>-3</sup> mol

 $\eta_{\text{NH}_3}$  in 250 cm<sup>3</sup> of FA 1 = 4.176 × 10<sup>-3</sup> ×  $\frac{250}{10}$  = 0.1044 mol [NH<sub>3</sub>]<sub>conc soln</sub> = 0.1044 / ( $\frac{45}{1000}$ ) = <u>2.32 mol dm<sup>-3</sup></u>

- 2. A solution of a dibasic acid contains 7.30 g dm<sup>-3</sup> of HOOC–(CH<sub>2</sub>)*n*–COOH. 20.0 cm<sup>3</sup> of this acid solution was titrated with 25.00 cm<sup>3</sup> of NaOH(aq) containing 1.36 g of hydroxide ion per dm<sup>3</sup>. Calculate
  - (a) the relative molecular mass of the acid;

[146.0]

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

 $[OH^{-}] = \frac{1.36}{17.0} = 0.08 \text{ mol } dm^{-3}$ Amount of NaOH reacted =  $\frac{25}{1000} \times 0.08 = 2.00 \times 10^{-3} \text{ mol}$ Amount of dibasic acid reacted =  $\frac{1}{2} \times 2.00 \times 10^{-3} = 1.00 \times 10^{-3} \text{ mol}$ 

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

 $M_{\rm r}$  of dibasic acid =  $\frac{7.30}{0.0500} = \frac{146.0}{0.0500}$ 

(b) the value of *n* in the formula.

 $M_{\rm r}$  of dibasic acid = 2(12.0 + 16.0 × 2 + 1.0) + **n**(12.0 + 2.0) = 146.0 **n** = 4.07 = <u>4</u> (to nearest integer) [4]

**3. FA 8** solution contains 20.2 g of the acid HZO<sub>4</sub> per dm<sup>3</sup> of solution.

**FA 9** is a 0.100 mol dm<sup>-3</sup> NaOH solution. In a titration, 20.0 cm<sup>3</sup> of **FA 8** solution reacted with 21.05 cm<sup>3</sup> 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<sup>+</sup> per molecule)

Amount of NaOH =  $0.100 \times \frac{21.05}{1000} = 2.105 \times 10^{-3}$  mol Amount of HZO<sub>4</sub> in 20.0 cm<sup>3</sup> FA 8 =  $2.105 \times 10^{-3}$  mol

 $[HZO_4] = \frac{2.105 \times 10^{-3}}{20/1000} = 0.1053 \text{ mol dm}^{-3}$   $M_r \text{ of } HZO_4 = \frac{20.2}{0.1053} = 191.9$  $A_r \text{ of } Z = 191.9 - 1.0 - 4 \times 16.0 = 126.9$ 

Z is lodine.

4. Washing soda has the formula Na<sub>2</sub>CO<sub>3</sub>·nH<sub>2</sub>O. A mass of 1.43 g of washing soda was made up to 250 cm<sup>3</sup> with water. 25.0 cm<sup>3</sup> of this solution was neutralised by 20.00 cm<sup>3</sup> of 0.050 mol dm<sup>-3</sup> 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 HCl =  $0.050 \times \frac{20}{1000} = 1.00 \times 10^{-3}$  mol Amount of Na<sub>2</sub>CO<sub>3</sub> in 25.0 cm<sup>3</sup> solution =  $\frac{1}{2} \times 1.00 \times 10^{-3} = 5.00 \times 10^{-4}$  mol Amount of Na<sub>2</sub>CO<sub>3</sub> in 250 cm<sup>3</sup> solution =  $5.00 \times 10^{-4} \times \frac{250}{25} = 5.00 \times 10^{-3}$  mol

Mass of Na<sub>2</sub>CO<sub>3</sub> in 250 cm<sup>3</sup> solution =  $5.00 \times 10^{-3} \times (2 \times 23.0 + 12.0 + 3 \times 16.0)$ = 0.530 g

∴ mass of H<sub>2</sub>O in 1.43 g washing soda = 1.43 - 0.530 = 0.900 g ∴ amount of H<sub>2</sub>O in 1.43 g washing soda =  $\frac{0.900}{18.0} = 0.0500$  mol

$$\frac{\eta_{\rm H_{20}}}{\eta_{\rm Na_{2}CO_{3}}} = \frac{0.0500}{5.00 \times 10^{-3}} = 10$$

10 mol of water combine with 1 mol of Na<sub>2</sub>CO<sub>3</sub> ∴ n = 10; formula of washing soda is Na<sub>2</sub>CO<sub>3</sub>·<u>10</u>H<sub>2</sub>O.

### **Back Titration**

 3.92 g of an oxide of formula MO was completely dissolved in 30.0 cm<sup>3</sup> of 2.00 mol dm<sup>-3</sup> sulfuric acid. The resulting solution was made up to 100 cm<sup>3</sup>. 25.0 cm<sup>3</sup> of this solution was neutralised by 27.50 cm<sup>3</sup> of 0.10 mol dm<sup>-3</sup> NaOH(aq).

What is the relative atomic mass of M? Identify the metal.

[55.9; Fe]



- **6.** 5.00 g of ammonium chloride contaminated with sodium chloride was boiled with 100.0 cm<sup>3</sup> of 2 mol dm<sup>-3</sup> NaOH solution until no ammonia was evolved. The residual solution was made up to 250 cm<sup>3</sup> with water. 25.0 cm<sup>3</sup> of this solution required 22.40 cm<sup>3</sup> of 0.50 mol dm<sup>-3</sup> HC*l* for neutralisation.
  - (a) What was the mass of sodium chloride in the ammonium chloride sample?

[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 =  $\frac{0.292}{5.00} \times 100\% = 5.84\%$ 

[5.58 g]

[44.2 %]

### **Redox Titration**

FA 4 contains 10.0 g of Fe<sup>2+</sup> and Fe<sup>3+</sup>, dissolved in 500 cm<sup>3</sup> of solution. FA 5 contains 0.015 mol dm<sup>-3</sup> KMnO<sub>4</sub>.

In an experiment, 10.0 cm<sup>3</sup> 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<sup>3</sup> of **FA 5** was required to reach end-point.

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

$$\begin{split} &\mathsf{MnO_{4^-}+8H^++5Fe^{2+}} \to 5Fe^{3+}+\mathsf{Mn^{2+}+4H_2O} \\ &\mathsf{Amount} \text{ of }\mathsf{KMnO_4} \text{ reacted} = 0.015\times\frac{26.65}{1000} = 3.998\times10^{-4} \text{ mol} \\ &\mathsf{Amount} \text{ of } Fe^{2+} \text{ in } 10.0 \text{ cm}^3 \text{ FA } 4 = 5\times3.998\times10^{-4} = 1.999\times10^{-3} \text{ mol} \\ &\mathsf{Amount} \text{ of } Fe^{2+} \text{ in } 500 \text{ cm}^3 \text{ FA } 4 = 1.999\times10^{-3}\times\frac{500}{10} = 0.09995 \text{ mol} \\ &\mathsf{Mass} \text{ of } Fe^{2+} \text{ in } 500 \text{ cm}^3 \text{ FA } 4 = 0.09995\times55.8 = \underline{5.58} \text{ g} \end{split}$$

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

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

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<sup>2+</sup> oxidized to Fe<sup>3+</sup> before combining the 2 half-equations]

 $ClO_{3}^{-} + 6H^{+} + 6Fe^{2+} \rightarrow Cl^{-} + 3H_{2}O + 6Fe^{3+}$ 

- (b) In an experiment, 25.0 cm<sup>3</sup> of a solution of KClO<sub>3</sub> was titrated with a solution of iron(II) sulfate containing 6.72 g dm<sup>-3</sup> of Fe<sup>2+</sup>. In the titration, 20.0 cm<sup>3</sup> 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<sup>+</sup>; dilute H<sub>2</sub>SO<sub>4</sub>

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 KC/O<sub>3</sub> in mol dm<sup>-3</sup>.

[0.0161 mol dm<sup>-3</sup>]

 $[Fe^{2+}] \text{ in mol } dm^{-3} = \frac{6.72}{55.8} = 0.1204 \text{ mol } dm^{-3}$ Amount of Fe<sup>2+</sup> reacted = 0.1204 ×  $\frac{20}{1000}$  = 2.409 × 10<sup>-3</sup> mol
Amount of KC/O<sub>3</sub> in 25.0 cm<sup>3</sup> solution =  $\frac{1}{6}$  × 2.409 × 10<sup>-3</sup> = 4.014 × 10<sup>-4</sup> mol  $[KC/O_3] = \frac{4.014 \times 10^{-4}}{25.0/1000} = \frac{0.0161 \text{ mol } dm^{-3}}{25.0/1000}$ 

(iii) Calculate the mass of potassium in 1 dm<sup>3</sup> of the KClO<sub>3</sub> solution.

[0.630 g]

Amount of K in 1 dm<sup>3</sup> of solution = 0.0161 mol Mass of K in 1 dm<sup>3</sup> of solution =  $39.1 \times 0.0161 = 0.630$  g

- 9. FA 2 is a solution containing 6.70 g dm<sup>-3</sup> of a metal ethanedioate, M<sub>2</sub>C<sub>2</sub>O<sub>4</sub>.
  FA 3 is a solution containing 0.010 mol dm<sup>-3</sup> of manganate(VII) ions, MnO<sub>4</sub><sup>-</sup>.
  10.0 cm<sup>3</sup> 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<sup>3</sup> 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 hot water bath / heat the 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  $M_2C_2O_4$  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<sub>4</sub> reacted =  $0.010 \times \frac{20}{1000} = 2.00 \times 10^{-4}$  mol Amount of C<sub>2</sub>O<sub>4</sub><sup>2-</sup> in 10.0 cm<sup>3</sup> **FA 2** =  $\frac{5}{2} \times 2.00 \times 10^{-4} = 5.00 \times 10^{-4}$  mol [C<sub>2</sub>O<sub>4</sub><sup>2-</sup>] in **FA 2** =  $\frac{5.00 \times 10^{-4}}{10.0/1000} = 0.0500$  mol dm<sup>-3</sup> *M*<sub>r</sub> of M<sub>2</sub>C<sub>2</sub>O<sub>4</sub> =  $\frac{6.70}{0.0500} = 134.0$ *A*<sub>r</sub> of M =  $\frac{1}{2}$  [134.0 - (2 × 12.0 + 16.0 × 4)] = <u>23.0</u> **10.** Hydroxylamine, NH<sub>2</sub>OH, can be oxidised to nitrogen according to the following half-equation: 2NH<sub>2</sub>OH  $\rightarrow$  N<sub>2</sub> + 2H<sub>2</sub>O + 2H<sup>+</sup> + 2e<sup>-</sup>

Hydroxylamine can also be oxidised to dinitrogen oxide, N<sub>2</sub>O, according to the following half-equation:  $2NH_2OH \rightarrow N_2O + H_2O + 4H^+ + 4e^-$ 

(a) In an experiment, 40.00 cm<sup>3</sup> of 0.05 mol dm<sup>-3</sup> NH<sub>2</sub>OH(aq) was reacted with 10.0 cm<sup>3</sup> of 0.40 mol dm<sup>-3</sup> Fe<sup>3+</sup>(aq). In the reaction, the Fe<sup>3+</sup> was reduced to Fe<sup>2+</sup>. Calculate the amount of Fe<sup>3+</sup> that reacted with 1 mole of NH<sub>2</sub>OH.

[2]

[R] Fe<sup>3+</sup> + e<sup>-</sup>  $\rightarrow$  Fe<sup>2+</sup> Amount of Fe<sup>3+</sup> reacted = 0.40 ×  $\frac{10}{1000}$  = 4.00 × 10<sup>-3</sup> mol Amount of NH<sub>2</sub>OH reacted = 0.05 ×  $\frac{40}{1000}$  = 2.00 × 10<sup>-3</sup> mol  $\therefore$  <u>2</u> mol of Fe<sup>3+</sup> reacted with 1 mol of NH<sub>2</sub>OH

(b) Hence deduce whether the NH<sub>2</sub>OH is oxidised to N<sub>2</sub> or N<sub>2</sub>O by the Fe<sup>3+</sup> and write a balanced ionic equation for the reaction.

If NH<sub>2</sub>OH was oxidised to N<sub>2</sub>: Overall eqn will be  $2NH_2OH + 2Fe^{3+} \rightarrow N_2 + 2H_2O + 2H^+ + 2Fe^{2+}$ 1 mol of Fe<sup>3+</sup> should react with 1 mol of NH<sub>2</sub>OH; <u>inconsistent</u> with answer in obtained in (a).

If NH<sub>2</sub>OH is oxidised to N<sub>2</sub>O: Overall eqn will be  $2NH_2OH + 4Fe^{3+} \rightarrow N_2O+ H_2O + 4H^+ + 4Fe^{2+}$ 2 mol of Fe<sup>3+</sup> should react with <u>1</u> mol of NH<sub>2</sub>OH; <u>consistent</u> with answer in obtained in (a).

 $\therefore$  NH<sub>2</sub>OH was oxidised to N<sub>2</sub>O in the reaction.

[+5]

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

[2]

- A solution of vanadium ions, VO<sub>2</sub><sup>+</sup>(aq) is blue. The VO<sub>2</sub><sup>+</sup>(aq) are reduced by zinc powder and acid to violet V<sup>z+</sup>(aq) ions.
  - (a) What is the oxidation number of the vanadium in  $VO_2^+$ ?

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

- (b) The V<sup>z+</sup>(aq) can be oxidised quantitatively to VO<sub>2</sub><sup>+</sup>(aq) by acidified MnO<sub>4</sub><sup>-</sup>(aq). It was found experimentally that 20.0 cm<sup>3</sup> of 0.10 mol dm<sup>-3</sup> V<sup>z+</sup>(aq) reacted with 24.00 cm<sup>3</sup> of 0.050 mol dm<sup>-3</sup> MnO<sub>4</sub><sup>-</sup>(aq).
  - (i) Calculate the number of moles of  $V^{z_+}(aq)$  that react with 1 mol of acidified MnO<sub>4</sub>-(aq).

Amount of KMnO<sub>4</sub> reacted =  $0.050 \times \frac{24}{1000} = 1.20 \times 10^{-3}$  mol Amount of V<sup>z+</sup> reacted =  $0.10 \times \frac{20}{1000} = 2.00 \times 10^{-3}$  mol

Amt of KMnO<sub>4</sub>: Amount of V<sup>z+</sup> =  $1.20 \times 10^{-3}$ :  $2.00 \times 10^{-3}$ = 3 : 5

:. Amount of V<sup>z+</sup> that reacted with 1 mol of MnO<sub>4</sub><sup>-</sup> =  $\frac{5}{3}$  mol

(ii) Hence deduce the number of moles of electrons transferred per mole of V<sup>z+</sup>(aq) in the reaction and value of z on V<sup>z+</sup>.

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

During reduction, one mole of MnO<sub>4</sub><sup>-</sup> gains 5e<sup>-</sup>; During oxidation,  $\frac{5}{3}$  mole of V<sup>z+</sup> loses 5e<sup>-</sup>

∴1 mole of V<sup>z+</sup> loses 3e<sup>-</sup>

 $\therefore$  no of electrons transferred (lost) per mole of V<sup>z+</sup> is <u>3</u>.

O.S. of V in VO<sub>2</sub><sup>+</sup> = +5 During oxidation, V<sup>z+</sup>  $\rightarrow$  VO<sub>2</sub><sup>+</sup> after losing 3 e<sup>-</sup>

Oxidation number of V increases by 3 units from +2 (in V<sup>z+</sup>) to +5 (in VO<sub>2</sub><sup>+</sup>)

Hence z = 2

- A 25.0 cm<sup>3</sup> sample of 0.0210 mol dm<sup>-3</sup> potassium peroxodisulfate(VI),K<sub>2</sub>S<sub>2</sub>O<sub>8</sub>, was treated with an excess of potassium iodide. The iodine liberated reacted with 21.00 cm<sup>3</sup> of 0.0500 mol dm<sup>-3</sup> thiosulfate.
  - (a) Calculate the amount of  $S_2O_8^{2-}$  used.

Amount of  $S_2O_8^{2-}$  used = 0.0210 ×  $\frac{25}{1000}$  = 5.25 × 10<sup>-4</sup> mol

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

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 × 10<sup>-4</sup> mol = Amt of  $I_2$  formed in Reaction 1

Amount of I<sup>-</sup> reacted in 1<sup>st</sup> rxn to form I<sub>2</sub> = 2 × 5.25 × 10<sup>-4</sup> = 1.05 × 10<sup>-3</sup> mol

 $\frac{\text{Amount of I}^{-}}{\text{Amount of S}_2 O_8^{2-}} = \frac{1.05 \times 10^{-3}}{5.25 \times 10^{-4}} = \frac{2}{1}$ 

 $\therefore$  2 mol of I<sup>-</sup> reacted with 1 mol of S<sub>2</sub>O<sub>8</sub><sup>2-</sup>

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

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

[2 mol]

[5.25 × 10<sup>-4</sup> mol]

**13.** A solution contained 5.53 g dm<sup>-3</sup> of a metal iodate(VII), **M**IO<sub>4</sub>. 25.0 cm<sup>3</sup> 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<sup>-3</sup> S<sub>2</sub>O<sub>3</sub><sup>2–</sup>, where starch was added only when the solution turned pale yellow. It was found that 33.30 cm<sup>3</sup> of S<sub>2</sub>O<sub>3</sub><sup>2–</sup> 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 <u>all</u>  $IO_4^-$  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<sub>2</sub>] at start of titration, some of the I<sub>2</sub> forms an insoluble complex with starch. Hence, not all the I<sub>2</sub> liberated reacts with S<sub>2</sub>O<sub>3</sub><sup>2–</sup>, titre value is lesser than actual.  $\therefore$  [IO<sub>4</sub>–] calculated is thus inaccurate

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

[85.9; Rb]

Amount of  $S_2O_3^{2-} = 0.120 \times \frac{33.30}{1000} = 3.996 \times 10^{-3} \text{ mol}$ Amount of  $I_2 = \frac{1}{2} \times 3.996 \times 10^{-3} = 1.998 \times 10^{-3} \text{ mol}$ Amount of  $IO_4^-$  in 25.0 cm<sup>3</sup> solution  $= \frac{1}{4} \times 1.998 \times 10^{-3} = 4.995 \times 10^{-4} \text{ mol}$ 

 $[\mathbf{MIO}_4] = \frac{4.995 \times 10^{-4}}{25.0/1000} = 0.01998 \text{ mol } \text{dm}^{-3}$   $M_r \text{ of } \mathbf{MIO}_4 = \frac{5.53}{0.01998} = 276.8$   $A_r \text{ of } \mathbf{M} = 276.8 - 126.9 - 4 \times 16.0 = 85.9$ **M** is Rb.

#### (B) Discussion Questions (Redox Practice)

| 14 | 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?                  |   |   |   |   |   |   |   |
|    | Α  | 1 | В | 2 | С | 3 | D | 4 |

Ans: (B)

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

2 moles of e<sup>--</sup> will be gained by 1 mole of NO<sub>2</sub>.

- 15 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:  $3MnO_4^{2-}(aq) + 2H_2O(l) \rightarrow MnO_2(s) + 2MnO_4^-(aq) + 4OH^-(aq)$
- 16 Write the half-equations and hence the full balanced equation for each reaction below.
- (a) Sodium sulfite, Na<sub>2</sub>SO<sub>3</sub>, reduces an orange acidified solution of dichromate(VI) ions, Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup>, to green solution containing Cr<sup>3+</sup> 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<sup>2+</sup>(aq), & brown fumes of nitrogen dioxide are produced.

Half eqn: [O] Cu  $\rightarrow$  Cu<sup>2+</sup> + 2e<sup>-</sup>------ (1) [R] NO<sub>3</sub><sup>-</sup> + 2H<sup>+</sup> + e<sup>-</sup>  $\rightarrow$  NO<sub>2</sub> + H<sub>2</sub>O ------- (2) x2

Overall eqn:  $Cu + 2NO_3^- + 4H^+ \rightarrow Cu^{2+} + 2NO_2 + 2H_2O_3$ 

(c) When potassium iodide (KI) is added to acidified hydrogen peroxide, brown I<sub>2</sub> solution is formed.

Half eqn: [O]  $2I^{-} \rightarrow I_{2} + 2e^{-}$  ......(1) [R]  $H_{2}O_{2} + 2H^{+} + 2e^{-} \rightarrow 2H_{2}O$  .....(2) Overall eqn:  $H_{2}O_{2} + 2I^{-} + 2H^{+} \rightarrow I_{2} + 2H_{2}O$ 

(d) When heated, phosphorus acid,  $H_3PO_3$ , gives phosphoric acid,  $H_3PO_4$ , and phosphine gas, PH<sub>3</sub>. Half eqn: [O]  $H_3PO_3 + H_2O \rightarrow H_3PO_4 + 2H^+ + 2e^- ------(1) \times 3$ [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<sub>3</sub>(aq), silver solid forms on the surface of zinc. Zn<sup>2+</sup> ions are formed.

Half eqn: [O]  $Zn \rightarrow Zn^{2+} + 2e^{-}$  ------ (1) [R]  $Ag^+ + e^- \rightarrow Ag$  ------- (2) x2 Overall eqn:  $Zn + 2Ag^+ \rightarrow Zn^{2+} + 2Ag$