Practice Questions

5 **(a)**

+5 −2 +3 0

$2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$

This is a redox reaction where the oxidation number of N decreases from +5 in KNO_3 to +3 in KNO_2, and the oxidation number of O increases from -2 in KNO_3 to 0 in O_2.

(b)

+6 +6

$2CrO_4^{2-}(aq) + 2H^+(aq) \rightarrow Cr_2O_7^{2-}(aq) + H_2O(l)$

This is not a redox reaction since there is no change in initial and final oxidation states for all elements.

(c)

+1 0 +2

$Cu_2O(s) + H_2SO_4(aq) \rightarrow Cu(s) + CuSO_4(aq) + H_2O(l)$

This is a redox reaction where Cu is both reduced and oxidised.

(d)

-4 0 +2 -2

$2CH_4(g) + 2NH_3(g) + 3O_2(g) \rightarrow 2HCN(g) + 6H_2O(g)$

This is a redox reaction where the oxidation number of O decreases from 0 in O_2 to -2 in H_2O, and the oxidation number of C increases from -4 in CH_4 to +2 in HCN.

6 **(a)**

Oxidation: $2I^- \rightarrow I_2 + 2e^-$

Reduction: $H_2SO_4 + 8H^+ + 8e^- \rightarrow H_2S + 4H_2O$

Overall: $\mathbf{8I^-(aq) + H_2SO_4(aq) + 8H^+(aq) \rightarrow 4I_2(g) + H_2S(g) + 4H_2O(l)}$

(b)

Oxidation: $MnO_4^{2-} \rightarrow MnO_4^- + e^-$

Reduction: $MnO_4^{2-} + 4H^+ + 2e^- \rightarrow MnO_2 + 2H_2O$

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

This is another example of a disproportionation reaction.

(c)

Note: Both the cation and the anion in FeC_2O_4 undergo oxidation.

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

$C_2O_4^{2-}(aq) \rightarrow 2CO_2(g) + 2e^-$

Oxidation: $FeC_2O_4 \rightarrow Fe^{3+} + 2CO_2 + 3e^-$

Reduction: $Ce^{3+}(aq) + e^- \rightarrow Ce^{2+}(aq)$

Overall: $\mathbf{FeC_2O_4(aq) + 3Ce^{3+}(aq) \rightarrow Fe^{3+}(aq) + 2CO_2(g) + 3Ce^{2+}(aq)}$

7 (a) **Oxidation:** $Mn(OH)_2 + 2OH^- \rightarrow MnO_2 + 2H_2O + 2e^-$

Reduction: $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$

Overall: $\mathbf{ClO^-(aq) + Mn(OH)_2(s) \rightarrow Cl^-(aq) + MnO_2(s) + H_2O(l)}$

(b) **Oxidation:** $C_2O_4^{2-} + 4OH^- \rightarrow 2CO_3^{2-} + 2H_2O + 2e^-$

Reduction: $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$

Overall: $\mathbf{2MnO_4^-(aq) + 3C_2O_4^{2-}(aq) + 4OH^-(aq) \rightarrow 2MnO_2(s) + 6CO_3^{2-}(aq) + 2H_2O(l)}$

(c) **Oxidation:** $ClO^- + 4OH^- \rightarrow ClO_3^- + 2H_2O + 4e^-$

Reduction: $ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$

Overall: $\mathbf{3ClO^-(aq) \rightarrow 2Cl^-(aq) + ClO_3^-(aq)}$

This is another example of a disproportionation reaction.

8 (a) $3SO_2 + Cr_2O_7^{2-} + 2H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-} + H_2O$

[Oxidation: $SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$

& Reduction: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$]

(b) $3SO_2 + Cr_2O_7^{2-} + 2H^+ \rightarrow 3SO_4^{2-} + 2Cr^{3+} + H_2O$

$Na_2S_2O_5 + 2HCl \rightarrow 2NaCl + 2SO_2 + H_2O$

Ratio of $Cr_2O_7^{2-}$: SO_2 : $Na_2S_2O_5$ = $1:3:\frac{3}{2}$

Amount of + $Cr_2O_7^{2-}$ reacted = $14.20 \times 10^{-3} \times 0.0100 = 1.420 \times 10^{-4}$ mol

Amount $Na_2S_2O_5$ reacted = $1.420 \times 10^{-3} \times \frac{3}{2} = 2.130 \times 10^{-4}$ mol

$$\text{Mass of } Na_2S_2O_5 \text{ present} = 2.130 \times 10^{-4} \times (2 \times 23.0 + 2 \times 32.1 + 5 \times 16.0) = 0.04051\text{ g}$$

$$\text{Concentration of } Na_2S_2O_5 \text{ in ppm} = \frac{0.04051}{100} \times 10^6 = 405 \text{ ppm}$$

Mass of $Na_2S_2O_5$ / g	:	Mass of preserved meat / g
0.04051	:	100
$\frac{0.04051}{100}$	:	1
$\frac{0.04051}{100} \times 10^6 = 405$	:	10^6

9 Since

Oxidation: $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$

or $4H_2O_2 \rightarrow 4O_2 + 8H^+ + 8e^-$

$n_{Cl_2O_7}$	:	n_{e^-}	:	$n_{H_2O_2}$
1	:		:	4
1	:	8	:	4

Reduction (unbalanced): $\overset{+7}{Cl_2O_7} + 8e^- \rightarrow$?

$\Rightarrow$ Each Cl atom accepted 4 e^-.

$\Rightarrow$ Oxidation number of Cl decreases from +7 to +3.

A $\overset{-1}{Cl^-}$ **B** $\overset{+1}{ClO^-}$ **C** $\overset{+3}{ClO_2^-}$ **D** $\overset{+5}{ClO_3^-}$

Answer: **C**

10

n_{NaNO_2}	:	$n_{(NH_4)_2SO_4}$
0.010	:	0.005

$n_{NO_2^-}$	:	n_{e^-}	:	$n_{NH_4^+}$
0.010	:		:	0.010
1	:	x	:	1

Since $\overset{+3}{NO_2^-} \xrightarrow{\text{Accept } x \text{ mol of } e^-} \overset{?}{\text{Final Product}} \xleftarrow{\text{Lose } x \text{ mol of } e^-} \overset{-3}{NH_4^+}$

$\Rightarrow$ New oxidation number of N = 0

A $\overset{0}{N_2}$ **B** $\overset{+2}{NO}$ **C** $\overset{+1}{N_2O}$ **D** $\overset{-1}{NH_2OH}$

Answer: **A**

11 **(a)** $2NH_2OH \rightarrow N_2 + 2H_2O + 2H^+ + 2e^-$

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

(b) Amount of Fe^{3+} reduced $= 10 \times 10^{-3} \times 0.50 = 5.00 \times 10^{-3}$ mol

Amount of NH_2OH reacted $= 50 \times 10^{-3} \times 0.050 = 2.50 \times 10^{-3}$ mol

n_{NH_2OH}	:	n_{e^-}	:	$n_{Fe^{3+}}$
2.50×10^{-3}	:		:	5.00×10^{-3}
1	:		:	2
1	:	2	:	2

Since
Reduction: $Fe^{3+} + e^- \rightarrow Fe^{2+}$
or $2Fe^{3+} + 2e^- \rightarrow 2Fe^{2+}$
Oxidation: $NH_2OH \rightarrow$ product $+ 2e^-$

Using $NH_2OH \rightarrow$ product $+ 2e^-$,
oxidation number of N in the reactant = oxidation number of N in product + (–2)
Oxidation number of N in product = –1 – (–2) = +1
NH_2OH is oxidised to N_2O.

The relevant oxidation half-equation is therefore

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

in which 1 mole of NH_2OH loses 2 mol of electrons.

NH_2OH is oxidised to N_2O.
Overall equation: $2NH_2OH + 4Fe^{3+} \rightarrow N_2O + 4Fe^{2+} + H_2O + 4H^+$

12 **(a)**

$$[Na_2S_2O_3.5H_2O] \text{ in } \mathbf{FA1} = \frac{12.1}{2\times23.0+2\times32.1+3\times16.0+5(2\times1.0+16.0)}$$
$$= 0.04875 \text{ mol dm}^{-3}$$
$$= [S_2O_3^{2-}] \text{ in } \mathbf{FA1}$$

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

Amount of $S_2O_3^{2-}$ used $= 24.40 \times 10^{-3} \times 0.04875 = 1.190 \times 10^{-3}$ mol

Amount of I_2 liberated $= 1.190 \times 10^{-3} \div 2 = 5.948 \times 10^{-4}$ mol

Concentration of XO_3^- ions in **FA2** = 7.95×10^{-3} mol dm^{-3}

Amount of XO_3^- which reacted with I^- to form $I_2 = 25.0 \times 10^{-3} \times 7.95 \times 10^{-3}$

$= 1.988 \times 10^{-4}$ mol

$n_{XO_3^-}$	:	n_{I_2}
1.988×10^{-4}	:	5.948×10^{-4}
1	:	**3**

Hence, **<u>3 mol</u>** of I_2 is liberated by 1 mol of XO_3^-.
(Note: It must be an integer.)

(b) Since

Oxidation: $2I^- \rightarrow I_2 + 2e^-$

or $6I^- \rightarrow 3I_2 + 6e^-$

n_{I_2}	:	n_{e^-}	:	$n_{XO_3^-}$
3	:		:	1
3	:	6	:	1

Reduction (unbalanced): $\overset{+5}{X}O_3^- + 6e^- \rightarrow$?

$\Rightarrow$ Oxidation number of X decreases from +5 to –1.

Hence, the product obtained is X^- where the oxidation number of X is –1.

(c)

Oxidation: $2I^- \rightarrow I_2 + 2e^-$

Reduction: $XO_3^- + 6H^+ + 6e^- \rightarrow X^- + 3H_2O$

Overall: $XO_3^-(aq) + 6I^-(aq) + 6H^+(aq) \rightarrow X^-(aq) + 3H_2O(l) + 3I_2(aq)$

13 Explanation for formation of CuI:

Reduction: $Cu^{2+} + e^- \longrightarrow Cu^+$ Oxidation: $2I^- \longrightarrow I_2 + 2e^-$ **Redox**: $2Cu^{2+} + 2I^- \longrightarrow 2Cu^+ + I_2$	However, **precipitation** of CuI also occurs: $Cu^+ + I^- \longrightarrow CuI(s)$

As these 2 processes take place simultaneously,
The overall equation is:
$2Cu^{2+}(aq) + 4I^-(aq) \longrightarrow 2CuI(s) + I_2(aq)$

•

Amount of thiosulfate ions = $0.0500 \times \dfrac{24.80}{1000}$ = <u>1.24×10^{-3} mol</u>

Mole ratio of $Cu^{2+} : I_2 : S_2O_3^{2-} = 2 : 1 : 2$

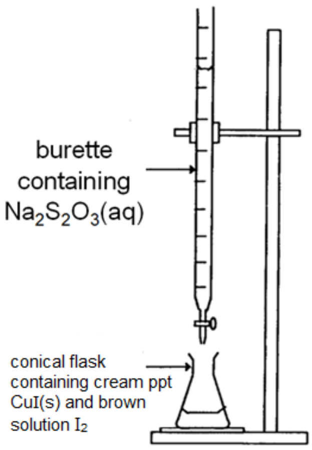
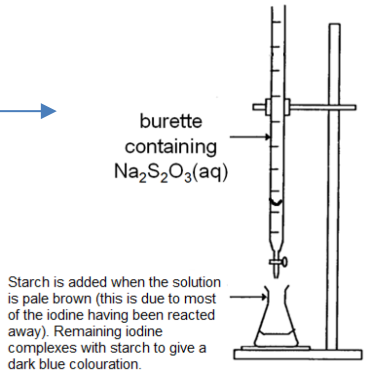
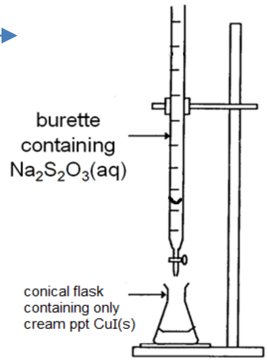
Amount of Cu^{2+} in 25.0 cm^3 of diluted solution = amount of $S_2O_3^{2-}$

$= 1.24 \times 10^{-3}$ mol

$[Cu^{2+}]$ in diluted solution = $\dfrac{1.24\times10^{-3}}{25.0\times10^{-3}}$ = 0.0496 mol dm^{-3}

$[Cu^{2+}]$ in original undiluted solution = $0.0496 \times \dfrac{250}{41}$ = <u>0.302 mol dm^{-3}</u>

Colour changes as titration progresses

Just before titration begins.	Addition of starch indicator during titration.	End point of titration. Colour of solution changes from dark blue to colourless, but pale cream ppt remains.
		

14 **(a)** $H_2C_2O_4(aq) + 2NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2H_2O(l)$

Amount of NaOH used $= 14.75 \times 10^{-3} \times 0.100 = 1.475 \times 10^{-3}$ mol

Amount of $H_2C_2O_4$ present $= 1.475 \times 10^{-3} \div 2 = 7.375 \times 10^{-4}$ mol

$$[H_2C_2O_4] \text{ in solution} = \frac{7.375 \times 10^{-4}}{25.0 \times 10^{-3}} = 0.0295 \text{ mol dm}^{-3}$$

(b) **(i)** oxidation: $C_2O_4^{2-}(aq) \rightarrow 2CO_2(g) + 2e^-$

reduction: $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

Overall:

$2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) + 16H^+(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(l)$

(ii) Amount of MnO_4^- used $= 32.00 \times 10^{-3} \times 0.0205 = 6.560 \times 10^{-4}$ mol

$$\text{Total amount of } C_2O_4^{2-} \text{ present} = \frac{6.560 \times 10^{-4}}{2} \times 5 = 1.64 \times 10^{-3} \text{ mol}$$

(c) Amount of $C_2O_4^{2-}$ from $H_2C_2O_4 = 7.375 \times 10^{-4}$ mol

Amount of $C_2O_4^{2-}$ from $Na_2C_2O_4 = 1.640 \times 10^{-3} - 7.375 \times 10^{-4} = 9.025 \times 10^{-4}$ mol

$= $ Amount of $Na_2C_2O_4$

$$\text{Hence, } [Na_2C_2O_4] \text{ in solution} = \frac{9.025 \times 10^{-4}}{25.0 \times 10^{-3}} = 0.0361 \text{ mol dm}^{-3}$$