Pages 92–93 Exam practice questions

1 a) $2H^+(aq) + 2e^- \rightarrow H_2(g)$ [1]

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

c) $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)$ [1]

2 a) $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ [1]

b) $Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$ [1]

c) $2I^-(aq) \rightarrow I_2(aq) + 2e^-$ [1]

3 a) $2Fe^{3+}(aq) + Sn^{2+}(aq)$ [1] $\rightarrow 2Fe^{2+}(aq) + Sn^{4+}(aq)$ [1]

b) $Mg(s) + 2H^+(aq)$ [1] $\rightarrow Mg^{2+}(aq) + H_2(g)$ [1]

c) $H_2O_2(aq) + 2H^+(aq) + 2I^-(aq)$ [1] $\rightarrow 2H_2O(l) + I_2(aq)$ [1]

4 Oxidation numbers: –1, +1, +3, +5, +7 (*3 marks for all 5, lose one for each mistake*)

5 Oxidation numbers: 0, –3, –2, +5, +3, –1, +3 (*4 marks for all 7, lose one for each mistake*)

6 a) $Co^{2+}(aq) \rightarrow Co^{3+}(aq) + e^-$ [1] (oxidation) [1]

b) $SO_2(aq) + 6H^+(aq) + 6e^- \rightarrow H_2S(g) + 2H_2O(l)$ (reduction) [1]

c) $4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$ [1] (oxidation) [1]

d) $H_2(g) \rightarrow 2H^+(aq) + 2e^-$ [1] (oxidation) [1]

7 a) Oxygen disproportionates [1] from –1 to –2 and 0. [1]

b) Chlorine disproportionates [1] from 0 to –1 and +1. [1]

c) Manganese disproportionates [1] from +6 to +7 and +4. [1]

8 a) $2Fe^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + 2e^-$ [1] (oxidised)

$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$ [1] (reduced) [1]

b) $2I^-(aq) \rightarrow I_2(aq) + 2e^-$ [1] (oxidised)

$Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$ [1] (reduced) [1]

c) $Zn(s) \rightarrow Zn^{2+}(aq)\ ) + 2e^-$ (oxidised) [1]

$2V^{3+}(aq) + 2e^- \rightarrow 2V^{2+}(aq)$ (reduced) [1]

9 a) $3CuO(s) + 2NH_3(g)$ [1] $\rightarrow N_2(g) + 3H_2O(l) + 3Cu(s)$ [1]

b) $8KI(s) + 5H_2SO_4(l)$ [1] $\rightarrow 4K_2SO_4(s) + 4I_2(s) + H_2S(g) + 4H_2O(l)$ [1]

c) $NaIO_3(aq) + 5NaI(aq) + 3H_2SO_4(aq)$ [1] $\rightarrow 3I_2(aq) + 3H_2O(l) + 3Na_2SO_4(aq)$ [1]

d) $Cu(s) + 4HNO_3(aq)$ [1] $\rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$ [1]

10 a) Br oxidised (–1 to 0) [1]

S reduced (+6 to +4) [1]

$2HBr(g) + H_2SO_4(l) \rightarrow Br_2(l) + SO_2(g) + 2H_2O(l)$ [1]

b) Fe oxidised (+2 to +3) [1]

Mn reduced (+7 to +2) [1]

$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$ [1]

c) Cr oxidised (+3 to +6) [1]

O in H_2O_2 reduced (–1 to –2) [1]

$2Cr^{3+}(aq) + 10OH^-(aq) + 3H_2O_2(aq) \rightarrow 2CrO_4^{2-}(aq) + 8H_2O(l)$ [1]

d) Cr reduced (+6 to +3) [1]

S oxidised (+4 to +6) [1]

$Cr_2O_7^{2-}(aq) + 2H^+(aq) + 3SO_2(aq) \rightarrow 2Cr^{3+}(aq) + H_2O(l) + 3SO_4^{2-}(aq)$ [1]

e) N in N_2H_4 oxidised (−2 to 0) [1]

N in N_2O_4 reduced (+4 to 0) [1]

$2N_2H_4(g) + N_2O_4(g) \rightarrow 3N_2(g) + 4H_2O(g)$ [1]

11 Definitions of the terms: 'oxidation' (gain of oxygen, loss of hydrogen, loss of electrons, oxidation number becoming more positive) and 'reduction' (loss of oxygen, gain of hydrogen, gain of electrons, oxidation number becoming less positive) [4]

a) Up to 4 marks for any key points such as:

$H_2S \rightarrow S$: sulfur oxidised from −2 to 0 state. [1] Older definition of loss of hydrogen applies. [1] Half-equation to show loss of electrons. [1]

$SO_2 \rightarrow S$: sulfur reduced from +4 to 0 state. [1] Definition of loss of oxygen applies. [1] Half-equation to show gain of electrons. [1]

b) Up to 4 marks for any key points such as:

$H_2 \rightarrow H^-$: hydrogen reduced from 0 to −1 state. [1] This gives hydrogen in an unusual oxidation state. It is acting as an oxidising agent, not a reducing agent. [1]

Half-equation to show gain of electrons. [1]

$Na \rightarrow Na^+$: sodium oxidised from 0 to +1 state. [1] Older definition of gain of hydrogen does not apply. [1] Half-equation to show loss of electrons. [1]

c) Up to 4 marks for any key points such as:

$H_2O_2 \rightarrow H_2O$: oxygen reduced from −1 to the −2 state. [1] Oxygen has an unusual oxidation state in peroxides. Because atoms of the element are bonded to each other [1] Definition of reduction in terms of loss of oxygen suggests that hydrogen is reduced when it does not change oxidation state [1] Half-equation to show gain of electrons. [1]

$I^- \rightarrow I_2$: iodide oxidised from −1 to 0 state. [1] Older definitions in terms of oxygen and hydrogen do not apply. [1] Half-equation to show loss of electrons. [1]

d) Up to 4 marks for any key points such as:

This is an example of disproportionation [1]

$H_2O_2 \rightarrow H_2O$: as in (c) [1]

$H_2O_2 \rightarrow O_2$: oxygen oxidised from −1 to the 0 state. [1] Definition of oxidation in terms of loss of hydrogen could be applied. [1] Half-equation to show loss of electrons. [1]

12 This question assesses a student's ability to show a coherent and logically structured answer with linkages and fully sustained line of reasoning.

a) Up to 4 marks for key points such as:

Sulfite: S is +4. [1]

Four outer electrons used for bonding to a more electronegative atom. [1]

One lone pair not involved in bonding. [1]

Sulfate: S is +6. [1] All six outer electrons used for bonding. [1]

b) Assess the quality of the answer taking into account both the key points made (*up to 4 marks*) and the logic and coherence of the discussion (*up to 2 marks*).

Points to make in an answer:

Thiosulfate:

- Applying the normal rules to the thiosulfate ion, without reference to structure, gives an average value of +2.
- However, many chemists now regard the thiosulfate ion as analogous to the sulfate ion with all six outer electrons of sulfur involved in bonding and assign the central S atom to the +6 state and the other S atom to the −2 state.

Tetrathionate:

- The average value for the oxidation state of S in tetrathionate is +2.5.
- However, the presence of an S–S bond suggests that the two S atoms involved should be assigned oxidation state −1 by analogy with the –O–O– situation in peroxides.
- In which case the oxidation states of the other two S atoms are +6 corresponding the use of 6 outer electrons in bonding as in the sulfate ion.
- The oxidation number rules are, to some extent arbitrary, and work so long as they are applied consistently.

c) Assess the quality of the answer taking into account both the key points made (*up to 4 marks*) and the logic and coherence of the discussion (*up to 2 marks*).

Points to make in an answer:

- The equation for the reaction is: $S_2O_3^{2-}(aq) + 2H^+(aq) \rightarrow SO_2(aq) + S(s) + H_2O(l)$
- This can be regarded as disproportionation …
- … as S in the +2 state $\rightarrow$ S in the +4 and 0 states.
- Alternatively this is seen as a redox reaction in which ….
- … the sulfur atom in the +6 state is reduced to the +4 state …
- … as it oxidises the other sulfur atom from the −2 to the 0 state.