CHEMISTRY (CODE - 043)

CHAPTER TEST - STEP-BY-STEP SOLUTIONS AND MARKING SCHEME

Class XI (2026-27) | Chapter 7: Redox Reactions

Max. Marks: 35

SECTION A (12 × 1 = 12 marks)

Each correct answer carries 1 mark.

Q1. Ans: B. +5 [1]

$H_4P_2O_7$: 4(+1) + 2x + 7(−2) = 0 → 2x = 10 → x = **+5**.

Q2. Ans: D. T [1]

T is ClO_4^-: Cl is in its highest oxidation state (+7), so it can only be reduced (oxidising agent only; it also cannot disproportionate). P (Cl^-, −1) can only be oxidised; Q, R and S are in intermediate states and can do both.

Q3. Ans: A. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ [1]

No element changes its oxidation number (Ca +2, C +4, O −2 on both sides). The other three involve changes (Cl +5 → −1, Na +1 → 0, Zn 0 → +2).

Q4. Ans: C. R and S [1]

A metal with a negative E° is a stronger reducing agent than H_2 and displaces hydrogen from acids. R (−0.44 V) and S (−0.76 V) are negative; P and Q (+0.80 V, +0.34 V) are positive and do not.

Q5. Ans: D. −½ [1]

Superoxide ion is O_2^-: 2x = −1 → x = **−½** for each oxygen.

Q6. Ans: C. Cu [1]

Copper has a positive E° (+0.34 V), so it is a weaker reducing agent than H_2 and does not displace H_2 from dilute HCl.

Q7. Ans: A. +2.5 [1]

4x + 6(−2) = −2 → 4x = 10 → x = **+2.5** (an average value; the actual states are +5, 0, 0, +5).

Q8. Ans: A. $Cl_2(g) + 2OH^-(aq) \rightarrow ClO^-(aq) + Cl^-(aq) + H_2O(l)$ [1]

Chlorine (0) is simultaneously oxidised to ClO^- (+1) and reduced to Cl^- (−1) → disproportionation. (B is combination, C is displacement, D is decomposition.)

Q9. Ans: B. $KMnO_4$ [1]

The permanganate ion is itself intensely coloured (purple); the first lasting pink tinge marks the end point, so no external indicator is needed.

Q10. Ans: A. Both A and R are true, and R is the correct explanation of A. [1]

Electrons lost in oxidation must be gained by some species (reduction), so the two always occur together.

Q11. Ans: D. A is false but R is true. [1]

In OF_2 fluorine (most electronegative) is −1, so oxygen is **+2**, not −2. A is false; R is true.

Q12. Ans: B. Both A and R are true, and R is not the correct explanation of A. [1]

Both statements are true, but fluorine does not disproportionate because, being the most electronegative element, it cannot exhibit any positive oxidation state – not simply because it is a strong oxidant.

SECTION B (4 × 2 = 8 marks)

Q13. [2]

(i) $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$. [½] Zn is **oxidised** (0 → +2); Cu^{2+} is **reduced** (+2 → 0). [½]

(ii) Reducing power: **Zn > Cu > Ag**. [½] In Q, Cu is a weaker reducing agent than Zn, so it cannot release electrons to Zn^{2+}; the equilibrium lies far on the side of the reactants and there is no visible change. [½]

Q14. [2]

(i) The larger the positive E°, the stronger the oxidising agent: strongest oxidising agent = $\mathbf{F_2}$ (+2.87 V). The couple with the lowest E° has the strongest reducing form: strongest reducing agent = $\mathbf{I^-}$ (E° = +0.54 V). [½ + ½]

(ii) Cl_2 (E° = +1.36 V) is a stronger oxidant than Br_2 (+1.09 V), so $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$ occurs: **yes**. [½] I_2 (+0.54 V) is a weaker oxidant than Br_2, so it cannot oxidise Br^-: **no**. [½]

Q15. [2]

(i) $H_2S_2O_7$: 2(+1) + 2x + 7(−2) = 0 → x = **+6** [½]

(ii) $NaBH_4$: (+1) + x + 4(−1) = 0 (H is −1 in a metal hydride) → x = **+3** [½]

(iii) CaO_2 (peroxide): (+2) + 2x = 0 → x = **−1** [½]

(iv) K_2MnO_4: 2(+1) + x + 4(−2) = 0 → x = **+6** [½]

Q16. [2]

(a) Combination (Mg and N_2 combine to give Mg_3N_2). (b) Decomposition ($KClO_3$ breaks into KCl and O_2). (c) Displacement – non-metal (halogen) displacement (Cl_2 displaces Br^-). (d) Disproportionation (N in NO_2 goes +4 → +3 and +5). [½ each]

OR

In HNO_3 nitrogen is in +5, its highest oxidation state (group number 15 − 10 = 5), so it can only decrease its oxidation number (only an oxidant). [1] In SO_2 sulphur is in +4, an intermediate state between −2 and +6, so it can be oxidised (to +6) as well as reduced (to lower states), acting as a reductant or an oxidant. [1]

SECTION C (2 × 3 = 6 marks)

Q17. [3]

(i) Oxidation: Fe^{2+} (+2) → Fe^{3+} (+3), increase of 1. Reduction: MnO_4^- (Mn +7) → Mn^{2+} (+2), decrease of 5. [1]

(ii) Oxidation: $Fe^{2+} \rightarrow Fe^{3+} + e^-$ Reduction: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$. (O balanced with $4H_2O$, H with $8H^+$, charge with $5e^-$: −1 + 8 − 5 = +2.) Multiply the oxidation half-reaction by 5: $5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$. [1]

(iii) Adding and cancelling $5e^-$: $\mathbf{MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)}$. Check: charge left = −1 + 8 + 10 = +17; right = +2 + 15 = +17. [1]

Q18. [3]

(i) Cl_2: 0; ClO^-: +1 (x − 2 = −1); Cl^-: −1. The same element (chlorine, intermediate state 0) is simultaneously oxidised (0 → +1) and reduced (0 → −1) → **disproportionation**. [1]

(ii) Oxidation: $Cl_2 + 4OH^- \rightarrow 2ClO^- + 2H_2O + 2e^-$. Reduction: $Cl_2 + 2e^- \rightarrow 2Cl^-$. Adding: $2Cl_2 + 4OH^- \rightarrow 2ClO^- + 2Cl^- + 2H_2O$, i.e. $Cl_2 + 2OH^- \rightarrow ClO^- + Cl^- + H_2O$. [1]

(iii) Fluorine is the most electronegative element and cannot exhibit any positive oxidation state, so it cannot form a higher state (+1) of the type ClO^-; only −1 and 0 are possible. (With alkali, F_2 instead gives F^- and OF_2.) [1]

SECTION D (1 × 4 = 4 marks)

Q19. [4]

(i) P = $KMnO_4$ solution (purple) in the burette; Q = Fe^{2+} solution with dilute H_2SO_4 in the conical flask. [½] The first lasting pink colour in R shows the **end point** – all Fe^{2+} has been oxidised and a slight excess of MnO_4^- is present. [½]

(ii) Mn: +7 → +2 (reduced); Fe: +2 → +3 (oxidised). [½] Oxidising agent = MnO_4^-; reducing agent = Fe^{2+}. [½]

(iii) Moles of MnO_4^- = M × V = 0.02 × (20.0/1000) = 0.02 × 0.020 = **4×10^{-4} mol** [½]

From the equation 1 mol MnO_4^- ≡ 5 mol Fe^{2+}: moles of Fe^{2+} = $5 \times 4 \times 10^{-4} = 2 \times 10^{-3}$ mol [½]

Molarity of Fe^{2+} = 2×10^{-3} mol / 0.025 L = **0.08 M** [1]

SECTION E (1 × 5 = 5 marks)

Q20. [5]

(i) R = **salt bridge** (U-tube containing KCl or NH_4NO_3 in agar-agar jelly). P (zinc, where oxidation occurs, negative) = **anode**; Q (copper, where reduction occurs, positive) = **cathode**. [1]

(ii) Electrons flow in the wire from P (Zn) to Q (Cu); the conventional current flows in the opposite direction, from Q to P. Inside the cell the current is carried by the **ions** (migrating through the solutions and the salt bridge). [1]

(iii) Anode (P): $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ [½] Cathode (Q): $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ [½] Overall: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$ [½]

(iv) The salt bridge completes the electrical circuit by allowing ions to migrate between the two solutions (maintaining electrical neutrality) without allowing the solutions to mix. [1]

(v) The standard electrode potential is the potential of an electrode when the concentration of each species in the electrode reaction is unity (gases at 1 atm pressure) at 298 K. [¼] For the hydrogen electrode it is **0.00 V**. [¼]

OR

(i) Pb_3O_4 is a stoichiometric mixture of 2 mol PbO (Pb +2) and 1 mol PbO_2 (Pb +4). [1]

(ii) PbO_2 (+4 lead) acts as an oxidant and oxidises Cl^- of HCl to Cl_2: $PbO_2 + 4HCl \rightarrow PbCl_2 + Cl_2 + 2H_2O$ (a **redox** reaction); PbO is basic and reacts in an acid–base reaction: $2PbO + 4HCl \rightarrow 2PbCl_2 + 2H_2O$. [1] HNO_3 itself is an oxidising agent, so PbO_2 cannot oxidise it; only the **acid–base** reaction of PbO with HNO_3 takes place (PbO_2 is passive). [1]

(iii) Oxidation: $SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$ (×3). Reduction: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$. [1]

Add: $Cr_2O_7^{2-} + 3SO_2 + 6H_2O + 14H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-} + 7H_2O + 12H^+ \rightarrow$ **$Cr_2O_7^{2-}(aq) + 3SO_2(g) + 2H^+(aq) \rightarrow 2Cr^{3+}(aq) + 3SO_4^{2-}(aq) + H_2O(l)$** [1]

***** END OF SOLUTIONS *****