Class 11 Chemistry · Chapter 7: Redox Reactions

Answer key and marking scheme · 20 marks · about 40 minutes · based on NCERT

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1. In the classical sense (addition or removal of oxygen, hydrogen or an electronegative element), which of the following conversions is a reduction? [1]

(a) $2FeCl_2 \rightarrow 2FeCl_3$, on passing chlorine

(b) $CH_3CH_2OH \rightarrow CH_3CHO$, on removing two hydrogen atoms

(c) $ZnO \rightarrow Zn$, on heating zinc oxide with carbon

(d) $2Cu \rightarrow 2CuO$, on heating copper in air

Answer: (c) Removal of oxygen from ZnO is reduction. Adding chlorine to $FeCl_2$, removing hydrogen from ethanol and adding oxygen to copper are all oxidations in the classical sense.

2. Which of the following decomposition reactions is NOT a redox reaction? [1]

(a) $2Ag_2O(s) \rightarrow 4Ag(s) + O_2(g)$

(b) $2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$

(c) $NH_4Cl(s) \rightarrow NH_3(g) + HCl(g)$

(d) $(NH_4)_2Cr_2O_7(s) \rightarrow N_2(g) + Cr_2O_3(s) + 4H_2O(g)$

Answer: (c) In $NH_4Cl \rightarrow NH_3 + HCl$, N stays −3, H +1 and Cl −1, and no element is formed. Ag_2O gives Ag (+1 → 0) and O_2 (−2 → 0); $NaNO_3$ gives N +5 → +3; ammonium dichromate gives N −3 → 0 and Cr +6 → +3.

3. Vanadium pentoxide, V_2O_5, written in Stock notation is [1]

(a) $V_2(X)O_5$

(b) $V_2(V)O_5$

(c) $V(II)_2O_5$

(d) $V_2(II)O_5$

Answer: (b) Each V is +5 ($2x - 10 = 0$), and Stock notation puts that number as a Roman numeral after the metal's symbol: $V_2(V)O_5$. X would be the total for both atoms; II is the subscript, not an oxidation number.

4. Hydrogen has an oxidation number of −1 in [1]

(a) H_2S

(b) NH_4^+

(c) HF

(d) MgH_2

Answer: (d) Hydrogen is −1 only when bonded to a metal in a binary compound, as in the hydride MgH_2. In H_2S, NH_4^+ and HF it is bonded to non-metals and is +1.

5. On the electron-transfer view, an oxidising agent is a species that [1]

(a) accepts electrons and is itself reduced

(b) donates electrons and is itself oxidised

(c) accepts electrons and is itself oxidised

(d) donates electrons and is itself reduced

Answer: (a) An oxidising agent is an acceptor of electrons; gain of electrons is reduction, so the oxidant is itself reduced. A donor of electrons is a reducing agent.

6. No indicator is added when $KMnO_4$ is used as the titrant in a redox titration, but one is needed with $K_2Cr_2O_7$. Explain, naming that indicator. [2]

MnO_4^- is intensely coloured, so it is its own indicator: the first lasting pink tinge after the last reductant is used up marks the end point. Dichromate shows no such sharp colour change, so diphenylamine is added; $Cr_2O_7^{2-}$ oxidises it just after the equivalence point to an intense blue.

- permanganate is a self indicator (first lasting pink) (1)
- dichromate needs diphenylamine, which turns intense blue (1)

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7. Find the oxidation number of chlorine in (i) HOCl and (ii) Cl_2O_7. Why is it positive in both? [2]

(i) HOCl: (+1) + x + (−2) = 0, so Cl = +1. (ii) Cl_2O_7: 2x + 7(−2) = 0, so Cl = +7. Chlorine is −1 only as a halide ion; combined with oxygen, which is more electronegative, it takes positive oxidation numbers, as in its oxoacids and oxides.

- Cl = +1 in HOCl and +7 in Cl_2O_7, with the sums shown (1)
- positive because chlorine is bonded to the more electronegative oxygen (1)

8. Ammonia is formed from its elements: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. No ions are formed, yet the reaction is called a redox reaction. Explain, with oxidation numbers. [2]

N goes from 0 in N_2 to −3 in NH_3 (reduced); H goes from 0 in H_2 to +1 in NH_3 (oxidised). NH_3 is covalent and the electrons are only shifted towards nitrogen, but for book-keeping the bonding electrons are taken as wholly transferred to the more electronegative atom, N. Since oxidation numbers change, the reaction is redox.

- N 0 → −3, reduced; H 0 → +1, oxidised (1)
- electron shift treated as complete transfer to the more electronegative N, for book-keeping (1)

9. A student classifies three reactions as follows: (i) $2Ag_2O(s) \rightarrow 4Ag(s) + O_2(g)$, a redox decomposition; (ii) $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$, a metal displacement; (iii) $Cl_2(g) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(l)$, a disproportionation, since one halogen is oxidised and the other reduced. Using oxidation numbers, say whether each classification is right, and correct any that is wrong. [3]

(i) Right: one compound breaks into elements, Ag +1 → 0 and O −2 → 0, so it is a redox decomposition. (ii) Wrong: Al 0 → +3 and H +1 → 0; aluminium displaces hydrogen, a non-metal, so it is a hydrogen displacement, not a metal displacement. (iii) Wrong: Cl 0 → −1 and Br −1 → 0 are two different elements. Disproportionation needs one element in one state to be both oxidised and reduced; this is a halogen displacement, chlorine displacing bromide.

- (i) right, with Ag +1 → 0 and O −2 → 0 (1)
- (ii) wrong: hydrogen (non-metal) displacement, Al 0 → +3, H +1 → 0 (1)
- (iii) wrong: two different elements change, so halogen displacement, not disproportionation (1)

10. In basic solution, iodate oxidises chromium(III) hydroxide to chromate: $Cr(OH)_3(s) + IO_3^-(aq) \rightarrow CrO_4^{2-}(aq) + I^-(aq)$. Balance the equation by the half reaction method. [3]

Oxidation half: $Cr(OH)_3 + 5OH^- \rightarrow CrO_4^{2-} + 4H_2O + 3e^-$ (Cr +3 → +6). Reduction half: $IO_3^- + 3H_2O + 6e^- \rightarrow I^- + 6OH^-$ (I +5 → −1). Multiply the oxidation half by 2 and add: $2Cr(OH)_3 + 10OH^- + IO_3^- + 3H_2O \rightarrow 2CrO_4^{2-} + 8H_2O + I^- + 6OH^-$. Cancel: $2Cr(OH)_3(s) + IO_3^-(aq) + 4OH^-(aq) \rightarrow 2CrO_4^{2-}(aq) + I^-(aq) + 5H_2O(l)$. Check: O 13 = 13, H 10 = 10, charge −5 = −5.

- oxidation half balanced in base with $3e^-$ (1)
- reduction half balanced in base with $6e^-$ (1)
- halves added in 2 : 1, cancelled to the final equation (1)

11. Aluminium burns in sulphur vapour: $2Al(s) + 3S(s) \rightarrow Al_2S_3(s)$. Write the two half reactions, and name the oxidising agent and the reducing agent. [3]

Al_2S_3 is ionic, $(Al^{3+})_2(S^{2-})_3$. Oxidation half: $2Al \rightarrow 2Al^{3+} + 6e^-$. Reduction half: $3S + 6e^- \rightarrow 3S^{2-}$. Their sum gives the overall reaction. Aluminium donates electrons, so it is the reducing agent; sulphur accepts them, so it is the oxidising agent.

- oxidation half $2Al \rightarrow 2Al^{3+} + 6e^-$ (1)
- reduction half $3S + 6e^- \rightarrow 3S^{2-}$ (1)
- reductant Al, oxidant S (1)

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