CLASS 11 CHEMISTRY NCERT SOLUTIONS

NCERT Solutions for Class 11 Chemistry Chapter 7: Redox Reactions – Free PDF Download

Redox Reactions is the first chapter of Part 2 of the NCERT Class 11 Chemistry textbook (2026-27 rationalised edition), and it builds the electron-transfer and oxidation-number toolkit that students will reuse throughout electrochemistry, p-block chemistry, and inorganic qualitative analysis in Class 12. The chapter...

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NCERT Solutions for Class 11 Chemistry Chapter 7: Redox Reactions

Q7.1: Assign oxidation numbers to the underlined elements

Assign oxidation numbers to the underlined element in each of the following species:

(a) NaH_2PO_4 (P): Na = +1, each O = -2. Let P = x. 1 + 2(1) + x + 4(-2) = 0 → 1 + 2 + x − 8 = 0 → x = +5.
(b) $NaHSO_4$ (S): 1 + 1 + x − 8 = 0 → x = +6.
(c) $H_4P_2O_7$ (P): 4(+1) + 2x + 7(-2) = 0 → 4 + 2x − 14 = 0 → 2x = 10 → x = +5 per P atom.
(d) K_2MnO_4 (Mn): 2(+1) + x + 4(-2) = 0 → 2 + x − 8 = 0 → x = +6.
(e) CaO_2 (O, calcium peroxide): Ca is always +2, and CaO_2 contains a peroxide (O-O) linkage, so each O = -1 (not -2). Check: +2 + 2(-1) = 0. → O = -1.
(f) $NaBH_4$ (B): H is bonded to a less electronegative element (B), so H = -1 (hydridic hydrogen). 1 + x + 4(-1) = 0 → 1 + x − 4 = 0 → x = +3.
(g) $H_2S_2O_7$ (S, pyrosulphuric acid/oleum): 2(+1) + 2x + 7(-2) = 0 → 2 + 2x − 14 = 0 → 2x = 12 → x = +6 per S atom.
(h) $KAl(SO_4)_2 \cdot 12H_2O$ (S, potash alum): each SO_4^{2-} group has O = -2 × 4 = -8, so S + (-8) = -2 → S = +6.

Final answer: P = +5, S = +6, P = +5, Mn = +6, O = -1, B = +3, S = +6, S = +6 respectively (worked from first principles above).

Q7.2: Oxidation numbers that need structural rationalisation

What are the oxidation numbers of the underlined elements in the following, and how do you rationalise your results?

(a) KI_3: In KI_3, the I_3^- ion is not three equivalent iodine atoms — it is an I_2 molecule (both atoms at oxidation state 0) coordinated to an I^- ion (oxidation state -1). The formula-based "average" calculation gives -1/3 per I atom [K = +1, so 3(I) = -1, average I = -1/3], but this average has no physical meaning; the true picture is two I atoms at 0 and one I atom at -1.
(b) $H_2S_4O_6$ (tetrathionate): Average from the formula: 2(+1) + 4x + 6(-2) = 0 → 2 + 4x − 12 = 0 → x = +2.5. Structurally, tetrathionate has a chain $O_3S\text{-}S\text{-}S\text{-}SO_3$: the two central sulphur atoms (bonded only to other S atoms) are at 0, and the two terminal (SO_3-type) sulphur atoms are at +5 each. Average = (0+0+5+5)/4 = +2.5 — matches the formula average, but the real oxidation states are 0, 0, +5, +5, not four identical +2.5 atoms.
(c) Fe_3O_4 (magnetite): Formula average: 3x + 4(-2) = 0 → x = +8/3. Structurally Fe_3O_4 is $FeO \cdot Fe_2O_3$, i.e., one Fe^{2+} and two Fe^{3+}; average = (2+3+3)/3 = +8/3, confirming the mix of +2 and +3 iron.
(d) CH_3CH_2OH (ethanol): Formula average: 2x + 6(+1) + 1(-2) = 0 → 2x = -4 → x = -2 per C. By bond-by-

bond counting (C-H gives C = -1, C-O gives C = +1, C-C gives 0): the methyl carbon (3 C-H, 1 C-C) = -3; the CH2OH carbon (2 C-H, 1 C-O, 1 C-C) = -1. Average = (-3 + -1)/2 = -2, matching the formula value, but the two carbons individually are -3 and -1, not both -2.
(e) CH3COOH (acetic acid): Formula average: 2x + 4(+1) + 2(-2) = 0 → 2x = 0 → x = 0 per C. By structure: the methyl carbon = -3 (3 C-H bonds); the carboxyl carbon (one C=O counted as two O-bonds, one C-OH, one C-C) = +2 + 1 + 0 = +3. Average = (-3 + 3)/2 = 0, matching the formula, but individually the carbons are -3 and +3, not both 0.

Final answer: The formula-based averages (-1/3, +2.5, +8/3, -2, 0) are mathematically correct but physically misleading, since each of these species contains chemically non-equivalent atoms of the same element at different real oxidation states, as shown above.

Q7.3: Justify that these reactions are redox reactions

Justify that the following reactions are redox reactions:

(a) CuO(s) + H2(g) → Cu(s) + H2O(g): Cu goes from +2 (in CuO) to 0 (reduction); H goes from 0 (in H2) to +1 (in H2O) (oxidation). Both oxidation and reduction occur, so it is a redox reaction.
(b) Fe2O3(s) + 3CO(g) → 2Fe(s) + 3CO2(g): Fe goes from +3 to 0 (reduction); C goes from +2 (in CO) to +4 (in CO2) (oxidation). Redox reaction.
(c) 4BCl3(g) + 3LiAlH4(s) → 2B2H6(g) + 3LiCl(s) + 3AlCl3(s): assigning carefully, in B2H6, H is bonded to less electronegative B, so H = -1 and B = +3 (since 2B + 6(-1) = 0 → B = +3, unchanged from BCl3). The actual electron transfer is on hydrogen: H in LiAlH4 is -1 (hydridic), and H in B2H6 is also -1, so hydrogen is transferred as a hydride ion without a net oxidation-state change either. This is essentially a metathesis (double-displacement) reaction with no change in oxidation number for any atom, so despite appearances it is *not* a genuine electron-transfer redox reaction — it proceeds by hydride-ion transfer, not electron transfer between different oxidation states.
(d) 2K(s) + F2(g) → 2K+F-(s): K goes from 0 to +1 (oxidation); F goes from 0 to -1 (reduction). Redox reaction.
(e) 4NH3(g) + 5O2(g) → 4NO(g) + 6H2O(g): N goes from -3 (in NH3) to +2 (in NO) (oxidation); O goes from 0 to -2 (reduction). Redox reaction.

Final answer: Reactions (a), (b), (d) and (e) are genuine redox reactions with a clear change in oxidation number of at least two elements; reaction (c) shows no net oxidation-state change on careful analysis and is better classified as a hydride-transfer metathesis reaction, a subtlety many secondary sources skip over.

Q7.4: Is the reaction of fluorine with ice a redox reaction?

Fluorine reacts with ice: H2O(s) + F2(g) → HF(g) + HOF(g). Assign oxidation numbers: in H2O, H = +1, O = -2. In F2, F = 0. In HF, H = +1, F = -1. In HOF (hypofluorous acid), fluorine (more electronegative than O) takes -1 as usual, and since H(+1) + O(x) + F(-1) = 0 for the neutral molecule, oxygen must be assigned 0 here (pushed toward a less negative state because it is bonded to the more electronegative F). So oxygen's oxidation number changes from -2 (in H2O) to 0 (in HOF) — an oxidation — while fluorine goes from 0 (in F2) to -1 (in both HF and HOF) — a reduction.

Final answer: Yes, it is a redox reaction: oxygen is oxidised from -2 to 0, and fluorine is reduced from 0 to -1.

Q7.5: Oxidation numbers of S, Cr and N — and the "fallacy"

Calculate the oxidation number of sulphur, chromium and nitrogen in H2SO5, Cr2O7^2- and NO3^-. Suggest the structures of these compounds and account for the fallacy.

H2SO5 (peroxomonosulphuric acid / Caro's acid): A naive calculation treating all five oxygens as -2 gives 2(+1) + x + 5(-2) = 0 → x = +8, which is chemically impossible since sulphur (Group 16) cannot exceed +6. This is the "fallacy." The correct structure is HO-SO2-O-O-H: two oxygens are normal double-bonded oxo oxygens (-2 each), one oxygen is a normal -OH oxygen (-2), and two oxygens form a peroxide (-O-O-) linkage (-1 each, as in H2O2). Recomputing: x + 2(+1 for the 2 H) + 2(-2) + 1(-2) + 2(-1) = 0 → x + 2 − 4 − 2 − 2 = 0 → x = +6.
Cr2O7^2- (dichromate): 2x + 7(-2) = -2 → 2x = 12 → x = +6 per Cr. No peroxide linkage is present here, so this is a genuine, unproblematic +6 (the well-known Cr(VI) state).
NO3^- (nitrate): x + 3(-2) = -1 → x = +5. This is nitrogen's maximum possible oxidation state (Group 15) and is likewise unproblematic.

Final answer: S in H2SO5 = +6 (not the naive +8, corrected for the O-O peroxide linkage), Cr in Cr2O7^2- = +6, N in NO3^- = +5. The "fallacy" is applying a blanket -2 to every oxygen even when a peroxide bond is present; only H2SO5 needed this correction.

Q7.6: Write formulae for the following compounds

(a) Mercury(II) chloride: HgCl2
(b) Nickel(II) sulphate: NiSO4
(c) Tin(IV) oxide: SnO2
(d) Thallium(I) sulphate: Tl2SO4
(e) Iron(III) sulphate: Fe2(SO4)3
(f) Chromium(III) oxide: Cr2O3

Final answer: HgCl2, NiSO4, SnO2, Tl2SO4, Fe2(SO4)3 and Cr2O3, obtained by balancing each metal's stated Roman-numeral charge against the standard charge of the anion (Cl- = -1, SO4^2- = -2, O^2- = -2).

Q7.7: Substances showing carbon (-4 to +4) and nitrogen (-3 to +5) oxidation states

Carbon shows every integral oxidation state from -4 to +4 through: CH4 (-4), C2H2 (-1), H2C=CH2 (-2), CH3OH (-2), HCHO (0), HCOOH (+2), CO2 (+4), CCl4 (+4). Nitrogen shows every integral state from -3 to +5 through: NH3 (-3), N2H4 (-2), NH2OH (-1), N2 (0), N2O (+1), NO (+2), N2O3/HNO2 (+3), NO2 (+4), N2O5/HNO3 (+5).

Final answer: Carbon spans -4 (CH4) to +4 (CO2, CCl4) through intermediate compounds like C2H2, CH3OH, HCHO and HCOOH; nitrogen spans -3 (NH3) to +5 (HNO3) through N2H4, NH2OH, N2, N2O, NO, HNO2 and NO2, as listed above.

Q7.8: Why are SO2 and H2O2 both oxidants and reductants, while O3 and HNO3 are only oxidants?

Sulphur in SO2 is at +4, an intermediate oxidation state for sulphur (whose range is -2 to +6) — it can rise to +6 (acting as a reducing agent, e.g. being oxidised to SO4^2-) or fall to lower states such as 0 or -2

(acting as an oxidising agent). Similarly, oxygen in H_2O_2 is at -1, intermediate between 0 (in O_2) and -2 (in H_2O), so H_2O_2 can be oxidised to O_2 (acting as reductant) or reduced to H_2O (acting as oxidant). In contrast, in O_3 the "active" oxygen has no easy access to a still-higher positive oxygen state, and in HNO_3 nitrogen is at +5, its maximum possible oxidation state — it can only decrease (never increase), so HNO_3 can only act as an oxidising agent, never as a reducing agent.

Final answer: SO_2 (S = +4) and H_2O_2 (O = -1) sit at intermediate oxidation states and can move either up or down, giving them dual oxidising/reducing character, whereas O_3 and HNO_3 (N at its ceiling of +5) can only be reduced further, so they act solely as oxidants.

Q7.9: Why write the water-balanced form of these reactions, and how to trace the mechanism

(a) $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$ is more correctly written as $6CO_2(g) + 12H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6H_2O(l) + 6O_2(g)$ because, mechanistically, the oxygen gas evolved in photosynthesis comes entirely from water molecules (not from CO_2), so 12 whole water molecules must be shown as reactants to supply the 12 oxygen atoms needed to form 6 O_2, with 6 water molecules regenerated as products.
(b) Similarly, $O_3(g) + H_2O_2(l) \rightarrow H_2O(l) + 2O_2(g)$ is better written as $O_3(g) + H_2O_2(l) \rightarrow H_2O(l) + O_2(g) + O_2(g)$, keeping the two product O_2 molecules distinct because they originate from two different sources: one O_2 forms from the peroxide oxygens of H_2O_2, and the other forms from ozone.
The technique used to establish which atoms end up where is **isotopic (heavy-isotope) labelling** — e.g., using water enriched with the oxygen-18 isotope ($H_2{}^{18}O$) and tracking whether the ^{18}O ends up in the O_2 product or remains as H_2O, thereby tracing the actual reaction path/mechanism.

Final answer: Both equations are rewritten to reflect that all the product O_2 in each case is mechanistically traceable to specific reactant oxygen atoms (water in photosynthesis; H_2O_2 and O_3 separately in the second case), a fact established using isotopic (^{18}O) tracer studies.

Q7.10: Why is AgF_2 a very strong oxidising agent?

In AgF_2, silver is forced into the unusual +2 oxidation state (Ag^{2+}). This state is highly unstable for silver, because silver's normal, far more stable oxidation state is +1 (Ag^+, with a stable filled d^{10} configuration). Ag^{2+} therefore has a very strong tendency to gain one electron and revert to the stable Ag^+ state, i.e., to get reduced. A species with such a strong drive to be reduced is, by definition, a powerful oxidising agent.

Final answer: AgF_2 contains silver in the unstable +2 state, which readily gains an electron to fall back to the far more stable +1 state; this strong drive toward reduction makes AgF_2 (where formed) a very strong oxidising agent.

Q7.11: Excess reductant vs. excess oxidant — three illustrations

When a redox reaction is carried out with the reducing agent in excess, the product tends to be the one with the lower oxidation state of the variable element; when the oxidising agent is in excess, the product has the higher oxidation state.

(i) Sulphur with limited O_2 gives SO_2 (S = +4); with excess O_2/further oxidation, SO_3 (S = +6) is favoured.
(ii) Phosphorus with limited Cl_2 gives PCl_3 (P = +3); with excess Cl_2, PCl_5 (P = +5) is formed.

(iii) Copper reacting with dilute HNO_3 (a comparatively weaker oxidant) gives NO (N = +2); with concentrated HNO_3 (stronger oxidising conditions), NO_2 (N = +4) is the major product — illustrating the same principle for the oxidant's own reduction product.

Final answer: Excess reductant → lower oxidation state product (e.g., PCl_3, SO_2, NO); excess oxidant → higher oxidation state product (e.g., PCl_5, SO_3, NO_2), as shown in the three illustrations above.

Q7.12: Alkaline $KMnO_4$ for benzoic acid; HCl gas vs Br_2 vapour with conc. H_2SO_4

(a) Toluene is oxidised to benzoic acid using alkaline $KMnO_4$ (followed by acidification) because the strongly oxidising, alkaline permanganate medium is powerful enough to oxidise the benzylic -CH_3 side chain all the way to -COOH (via the intermediate potassium benzoate) without over-oxidising or degrading the aromatic ring. Balanced equation:
$C_6H_5CH_3 + 2KMnO_4 \rightarrow C_6H_5COOK + 2MnO_2 + KOH + H_2O$
(Atom check: C 7=7, H 8=8, O 8=8, K 2=2, Mn 2=2 — balanced.) The potassium benzoate formed is then acidified to yield free benzoic acid.
(b) When concentrated H_2SO_4 is added to a chloride salt, colourless HCl gas is evolved because Cl^- is too weak a reducing agent to be oxidised by H_2SO_4, so only a simple (non-redox) acid-displacement occurs. But with a bromide salt, Br^- is a stronger reducing agent than Cl^-, so concentrated H_2SO_4 can additionally oxidise the HBr formed to red-brown Br_2 vapour (H_2SO_4 itself being reduced, typically to SO_2): $2HBr + H_2SO_4 \rightarrow Br_2 + SO_2 + 2H_2O$.

Final answer: (a) Alkaline $KMnO_4$ fully oxidises the -CH_3 group to -COOK without destroying the ring, per the balanced equation above. (b) Chloride only undergoes acid displacement (HCl gas) because Cl^- resists oxidation, while bromide's stronger reducing character lets conc. H_2SO_4 additionally oxidise HBr to Br_2 vapour.

Q7.13: Identify substance oxidised/reduced and oxidising/reducing agents

(a) $2AgBr(s) + C_6H_6O_2(aq) \rightarrow 2Ag(s) + 2HBr(aq) + C_6H_4O_2(aq)$: Ag^+ (in AgBr) is reduced to Ag(0); hydroquinone ($C_6H_6O_2$) is oxidised to quinone ($C_6H_4O_2$). Oxidising agent = AgBr; reducing agent = hydroquinone.
(b) $HCHO(l) + 2[Ag(NH_3)_2]^+(aq) + 3OH^-(aq) \rightarrow 2Ag(s) + HCOO^-(aq) + 4NH_3(aq) + 2H_2O(l)$: HCHO is oxidised to $HCOO^-$; Ag^+ (as the diammine complex) is reduced to Ag(0). Oxidising agent = $[Ag(NH_3)_2]^+$; reducing agent = HCHO.
(c) $HCHO(l) + 2Cu^{2+}(aq) + 5OH^-(aq) \rightarrow Cu_2O(s) + HCOO^-(aq) + 3H_2O(l)$: HCHO is oxidised to $HCOO^-$; Cu^{2+} is reduced to Cu^+ (in Cu_2O). Oxidising agent = Cu^{2+}; reducing agent = HCHO.
(d) $N_2H_4(l) + 2H_2O_2(l) \rightarrow N_2(g) + 4H_2O(l)$: N (-2 in N_2H_4) is oxidised to N_2 (0); O (-1 in H_2O_2) is reduced to -2 (in H_2O). Oxidising agent = H_2O_2; reducing agent = N_2H_4.
(e) $Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$: Pb(0) is oxidised to Pb^{2+} (in $PbSO_4$); Pb (+4, in PbO_2) is reduced to Pb^{2+}. Oxidising agent = PbO_2; reducing agent = Pb(s). (This is the discharge reaction of the lead storage battery.)

Final answer: Oxidising/reducing agent pairs are (a) AgBr / hydroquinone, (b) $[Ag(NH_3)_2]^+$ / HCHO, (c) Cu^{2+} / HCHO, (d) H_2O_2 / N_2H_4, (e) PbO_2 / Pb, as identified from the oxidation-number changes above.

Q7.14: Why does thiosulphate react differently with iodine and bromine?

2S2O3^2-(aq) + I2(s) → S4O6^2-(aq) + 2I-(aq) and S2O3^2-(aq) + 4Br2(l) + 5H2O(l) → 2SO4^2-(aq) + 8Br-(aq) + 10H+(aq). In thiosulphate, S is at an average +2. Iodine (I2/I-, E° = +0.54 V) is a comparatively mild oxidising agent, so it can only partially oxidise thiosulphate, taking sulphur from +2 to +2.5 (tetrathionate, S4O6^2-). Bromine (Br2/Br-, E° = +1.09 V) is a much stronger oxidising agent, so it can fully oxidise sulphur all the way to +6 (sulphate, SO4^2-).

Final answer: Iodine, being a weaker oxidant, converts thiosulphate only to tetrathionate (S: +2 → +2.5), while the much stronger oxidant bromine drives the oxidation all the way to sulphate (S: +2 → +6).

Q7.15: Fluorine as best oxidant; hydroiodic acid as best reductant

Among the halogens, fluorine has the highest standard reduction potential (F2/F-, E° = +2.87 V) due to its small atomic size, very high electronegativity, weak F-F bond, and very high hydration enthalpy of F-, all of which favour F2 readily accepting electrons — e.g., F2 oxidises even water: 2F2 + 2H2O → 4HF + O2, something no other halogen does. Among the hydrohalic acids, HI is the best reductant because the H-I bond is the weakest (largest, most polarisable iodine atom, longest and weakest bond), so I- is oxidised most easily of all the halide ions — e.g., HI reduces concentrated H2SO4 all the way to H2S, whereas HCl cannot reduce H2SO4 at all.

Final answer: F2 is the best oxidant due to its high electronegativity, small size, weak F-F bond and high hydration energy of F- (highest E°); HI is the best reductant because its weak H-I bond lets I- lose an electron most readily of all the halides.

Q7.16: XeO6^4- reacting with fluoride — what it reveals about Na4XeO6

XeO6^4-(aq) + 2F-(aq) + 6H+(aq) → XeO3(g) + F2(g) + 3H2O(l). Xenon falls from +8 (in the perxenate ion, XeO6^4-) to +6 (in XeO3) — a reduction — while fluoride is oxidised from -1 to 0 (F2). Since F2/F- already has the highest known standard reduction potential of any common couple (+2.87 V), for perxenate to be able to oxidise F- to F2 at all, the XeO6^4-/XeO3 couple must have an even higher reduction potential than F2/F-.

Final answer: The reaction shows that perxenate (as in Na4XeO6) is an exceptionally powerful oxidising agent — stronger even than elemental fluorine — since it is able to oxidise F- ions to F2 gas.

Q7.17: What Ag+ vs Cu2+ reactions reveal about their oxidising strength

H3PO2 reduces both [Ag(NH3)2]+ and Cu2+ to the metal/lower oxide. Benzaldehyde (C6H5CHO, an aromatic aldehyde) reduces [Ag(NH3)2]+ (Tollens' reagent) to metallic silver, but shows no reaction at all with Cu2+ (Fehling-type reagent). This shows that Ag+ is a distinctly stronger oxidising agent than Cu2+ — strong enough to oxidise even a comparatively resistant aromatic aldehyde, which Cu2+ cannot do.

Final answer: Ag+ (as [Ag(NH3)2]+) is a stronger oxidising agent than Cu2+, since it can oxidise even the less reactive aromatic aldehyde (benzaldehyde), which fails to react with Cu2+ at all.

Q7.18: Balance the following redox reactions by the ion-electron method

(i) MnO4^- + I^- → MnO2 + I2 (basic medium). Mn: +7 → +4 (gain 3e-); I: -1 → 0 (lose 1e- per I, 2e- per I2). Reduction half: MnO4- + 2H2O + 3e- → MnO2 + 4OH-. Oxidation half: 2I- → I2 + 2e-. Equalising electrons (×2 and ×3) and adding:
2MnO4^- + 4H2O + 6I^- → 2MnO2 + 3I2 + 8OH^-
(ii) MnO4^- + SO2 → Mn2+ + HSO4^- (acidic medium). Mn: +7 → +2 (gain 5e-); S: +4 → +6 (lose 2e-). Reduction half: MnO4- + 8H+ + 5e- → Mn2+ + 4H2O. Oxidation half: SO2 + 2H2O → HSO4- + 3H+ + 2e-. Equalising electrons (×2 and ×5) and adding, then cancelling common H2O/H+:
2MnO4^- + 5SO2 + 2H2O + H^+ → 2Mn2+ + 5HSO4^-
(iii) H2O2 + Fe2+ → Fe3+ + H2O (acidic medium). O: -1 → -2 (gain 2e- per H2O2); Fe: +2 → +3 (lose 1e-). Reduction half: H2O2 + 2H+ + 2e- → 2H2O. Oxidation half: 2Fe2+ → 2Fe3+ + 2e-. Adding directly:
H2O2 + 2Fe2+ + 2H^+ → 2Fe3+ + 2H2O
(iv) Cr2O7^2- + SO2 → Cr3+ + SO4^2- (acidic medium). Cr: +6 → +3 per Cr (gain 3e- each, 6e- total); S: +4 → +6 (lose 2e-). Reduction half: Cr2O7^2- + 14H+ + 6e- → 2Cr3+ + 7H2O. Oxidation half: SO2 + 2H2O → SO4^2- + 4H+ + 2e-. Equalising electrons (×1 and ×3) and simplifying:
Cr2O7^2- + 3SO2 + 2H^+ → 2Cr3+ + 3SO4^2- + H2O

Final answer: The four balanced equations are as boxed above, each verified atom-by-atom and charge-by-charge using the ion-electron (half-reaction) method.

Q7.19: Balance in basic medium by both methods; identify oxidising and reducing agents

(a) P4(s) + OH^-(aq) → PH3(g) + H2PO2^-(aq). This is a disproportionation of P4 (oxidation state 0): one P is reduced to -3 (in PH3, gaining 3e-) and another P is oxidised to +1 (in H2PO2-, losing 1e-). Balancing the two half-reactions and combining in a 1:3 ratio of reduced to oxidised phosphorus gives, after simplification:
P4 + 3OH^- + 3H2O → PH3 + 3H2PO2^-
(checked: P 4=4, O 6=6, H 9=9, charge -3=-3). Since P4 is simultaneously oxidised and reduced, it is both the oxidising agent and the reducing agent.
(b) N2H4(l) + ClO3^-(aq) → NO(g) + Cl^-(aq). N: -2 → +2 per N (oxidation, 4e- per N, 8e- per N2H4); Cl: +5 → -1 (reduction, 6e-). Oxidation half: N2H4 + 8OH- → 2NO + 6H2O + 8e-. Reduction half: ClO3- + 3H2O + 6e- → Cl- + 6OH-. Equalising electrons (LCM 24: ×3 and ×4) and simplifying:
3N2H4 + 4ClO3^- → 6NO + 4Cl^- + 6H2O
Oxidising agent = ClO3-; reducing agent = N2H4.
(c) Cl2O7(g) + H2O2(aq) → ClO2^-(aq) + O2(g) + H^+. Cl: +7 → +3 (reduction, 4e- per Cl, 8e- per Cl2O7); O (peroxide): -1 → 0 (oxidation, 2e- per H2O2). Reduction half: Cl2O7 + 6H+ + 8e- → 2ClO2- + 3H2O. Oxidation half: H2O2 → O2 + 2H+ + 2e-. Equalising electrons (×1 and ×4) and simplifying:
Cl2O7 + 4H2O2 → 2ClO2^- + 4O2 + 3H2O + 2H^+
Oxidising agent = Cl2O7; reducing agent = H2O2.

Final answer: (a) P4 + 3OH- + 3H2O → PH3 + 3H2PO2- (P4 is both oxidant and reductant); (b) 3N2H4 + 4ClO3- → 6NO + 4Cl- + 6H2O (oxidant: ClO3-, reductant: N2H4); (c) Cl2O7 + 4H2O2 → 2ClO2- + 4O2 + 3H2O + 2H+ (oxidant: Cl2O7, reductant: H2O2).

Q7.20: What information can you draw from the cyanogen disproportionation reaction?

(CN)2(g) + 2OH-(aq) → CN-(aq) + CNO-(aq) + H2O(l). Assigning oxidation numbers (N = -3 throughout, O = -2): in cyanogen, (CN)2, carbon is at +3 (average); in the product cyanide ion, CN-, carbon is at +2; in the product cyanate ion, CNO-, carbon is at +4. So carbon splits from +3 into +2 (reduced) and +4 (oxidised) — a disproportionation. This is exactly analogous to the disproportionation of a halogen in alkali, e.g., Cl2 + 2OH- → Cl- + OCl- + H2O, which is why (CN)2 is classed as a pseudohalogen.

Final answer: The reaction shows that (CN)2 undergoes disproportionation (carbon splits from +3 into +2 in CN- and +4 in CNO-) in exactly the same pattern as a halogen reacting with alkali, confirming cyanogen's classification as a "pseudohalogen."

Q7.21: Balanced equation for Mn3+ disproportionation

Mn3+ is unstable in solution and disproportionates to Mn2+ and MnO2. One Mn3+ is reduced to Mn2+ (gains 1e-); another Mn3+ is oxidised to Mn4+ as MnO2 (loses 1e-, and needs O supplied by water, releasing H+). Combining 1:1 and supplying oxygen/hydrogen via water:
2Mn3+(aq) + 2H2O(l) → Mn2+(aq) + MnO2(s) + 4H+(aq)
Check: Mn 2=2; O 2=2; H 4=4; charge LHS +6, RHS (+2) + (+4 from 4H+) = +6. Balanced.

Final answer: 2Mn3+(aq) + 2H2O(l) → Mn2+(aq) + MnO2(s) + 4H+(aq).

Q7.22: Oxidation-state behaviour of Cs, Ne, I and F

(a) Element showing only a negative oxidation state: **F** (the most electronegative element of all, never found in a positive oxidation state).
(b) Element showing only a positive oxidation state: **Cs** (a highly electropositive alkali metal, essentially always +1).
(c) Element showing both positive and negative oxidation states: **I** (iodine is -1 in iodides, but can be +1, +3, +5 or +7 in oxoacids, oxides and interhalogen compounds).
(d) Element showing neither a positive nor a negative oxidation state (stays at 0): **Ne** (a noble gas that forms essentially no stable compounds).

Final answer: F = only negative; Cs = only positive; I = both positive and negative; Ne = neither (stays at 0).

Q7.23: Balanced equation for removing excess chlorine from drinking water with SO2

Cl2 is reduced (0 → -1) and SO2 is oxidised (+4 → +6, forming sulphate/sulphuric acid), consuming water:
Cl2(aq) + SO2(g) + 2H2O(l) → 2HCl(aq) + H2SO4(aq)
Check: Cl 2=2; S 1=1; O 4=4; H 4=4. Balanced.

Final answer: Cl2(aq) + SO2(g) + 2H2O(l) → 2HCl(aq) + H2SO4(aq).

Q7.24: Non-metals and metals that show disproportionation

(a) Possible non-metals: elements with several accessible intermediate oxidation states, such as **Cl, Br, I (halogens: e.g. Cl2 + 2OH- → Cl- + OCl- + H2O), P (e.g. P4 → PH3 + H2PO2-, as in Q7.19a), and S (e.g. S8 in hot alkali giving sulfide- and thiosulfate-type products).**
(b) Three metals: **Cu (2Cu+ → Cu2+ + Cu, since Cu+ is unstable in aqueous solution), Au (3Au+ → Au3+ + 2Au), and Mn (Mn3+ → Mn2+ + MnO2, as in Q7.21).**

Final answer: Non-metals: Cl, Br, I, P, S; Metals: Cu, Au, Mn, each chosen because the element has an intermediate oxidation state that can simultaneously rise and fall.

Q7.25: Maximum mass of NO from the Ostwald process (limiting-reagent calculation)

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$. Molar mass NH_3 = 17 g/mol; O_2 = 32 g/mol; NO = 30 g/mol.
Moles of NH_3 = 10.00/17 = 0.588 mol. Moles of O_2 = 20.00/32 = 0.625 mol.
Required O_2 for all the NH_3 (ratio $NH_3:O_2$ = 4:5) = 0.588 × (5/4) = 0.735 mol, but only 0.625 mol O_2 is available → O_2 is the limiting reagent.
Moles of NO formed = moles O_2 used × (4/5) = 0.625 × 0.8 = 0.500 mol.
Mass of NO = 0.500 mol × 30 g/mol = 15.0 g.

Final answer: 15.0 g of NO is the maximum obtainable, since O_2 (not NH_3) is the limiting reagent.

Q7.26: Feasibility of reactions from standard electrode potentials

Using standard reduction potentials: Fe^{3+}/Fe^{2+} = +0.77 V; I_2/I^- = +0.54 V; Ag^+/Ag = +0.80 V; Cu^{2+}/Cu = +0.34 V; Br_2/Br^- = +1.09 V. A reaction is spontaneous when the species with the higher (more positive) reduction potential is reduced by the species with the lower one.

(a) Fe^{3+} + I^-: 0.77 V > 0.54 V, so Fe^{3+} can oxidise I^-. **Feasible**: $2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$.
(b) Ag^+ and Cu: 0.80 V > 0.34 V, so Ag^+ can oxidise Cu. **Feasible**: $2Ag^+ + Cu \rightarrow 2Ag + Cu^{2+}$.
(c) Fe^{3+} and Cu: 0.77 V > 0.34 V, so Fe^{3+} can oxidise Cu. **Feasible**: $2Fe^{3+} + Cu \rightarrow 2Fe^{2+} + Cu^{2+}$.
(d) Ag and Fe^{3+}: this would require 0.77 V > 0.80 V, which is false (E°_{cell} = 0.77 − 0.80 = -0.03 V). **Not feasible.**
(e) Br_2 and Fe^{2+}: 1.09 V > 0.77 V, so Br_2 can oxidise Fe^{2+}. **Feasible**: $2Fe^{2+} + Br_2 \rightarrow 2Fe^{3+} + 2Br^-$ (E°_{cell} = +0.32 V).

Final answer: Feasible: (a), (b), (c) and (e); Not feasible: (d), since Ag^+/Ag (0.80 V) has a higher reduction potential than Fe^{3+}/Fe^{2+} (0.77 V), so silver metal cannot reduce Fe^{3+}.

Q7.27: Products of electrolysis

(i) Aqueous $AgNO_3$ with silver electrodes: silver dissolves from the anode and deposits on the cathode, net effect being transfer of silver metal from anode to cathode with no change in solution concentration.
(ii) Aqueous $AgNO_3$ with platinum (inert) electrodes: Ag^+ is reduced at the cathode, depositing silver metal; at the anode, water is oxidised instead of NO_3^-: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$, releasing O_2 gas.
(iii) Dilute H_2SO_4 with platinum electrodes: essentially electrolysis of water, giving H_2 gas at the cathode and O_2 gas at the anode, in a 2:1 volume ratio.
(iv) Aqueous $CuCl_2$ with platinum electrodes: Cu^{2+} is reduced and deposited as copper metal at the cathode; Cl^- is oxidised at the anode to Cl_2 gas.

Final answer: (i) Ag transferred anode→cathode, solution unchanged; (ii) Ag deposited at cathode, O_2 evolved at anode; (iii) H_2 at cathode, O_2 at anode (2:1 by volume); (iv) Cu deposited at cathode, Cl_2 gas evolved at anode.

Q7.28: Arranging Al, Cu, Fe, Mg and Zn by displacement order

Using standard reduction potentials (more negative = more easily oxidised = more reactive): Mg^{2+}/Mg ≈ -2.36 V, Al^{3+}/Al ≈ -1.66 V, Zn^{2+}/Zn ≈ -0.76 V, Fe^{2+}/Fe ≈ -0.44 V, Cu^{2+}/Cu = +0.34 V.

Final answer: Mg > Al > Zn > Fe > Cu (most reactive/best displacer first) — each metal in this list displaces every metal to its right from solutions of their salts.

Q7.29: Increasing order of reducing power from given electrode potentials

Given: K+/K = -2.93 V, Ag+/Ag = +0.80 V, Hg2+/Hg = +0.79 V, Mg2+/Mg = -2.37 V, Cr3+/Cr = -0.74 V. Reducing power increases as the standard reduction potential becomes more negative.

Final answer: Ag < Hg < Cr < Mg < K (increasing reducing power), following directly from arranging the given E° values from most positive to most negative.

Q7.30: Depicting the Zn/Ag galvanic cell

Zn(s) + 2Ag+(aq) → Zn2+(aq) + 2Ag(s). Cell representation: Zn(s) | Zn2+(aq) || Ag+(aq) | Ag(s).

(i) The zinc electrode is the **anode** (where oxidation occurs), and in a galvanic cell the anode is the **negatively charged** electrode.
(ii) Current carriers: inside the cell, ions carry the current through the electrolyte; in the external circuit, electrons flowing from zinc to silver carry the current.
(iii) Electrode reactions: anode (Zn): Zn(s) → Zn2+(aq) + 2e- (oxidation). Cathode (Ag): 2Ag+(aq) + 2e- → 2Ag(s) (reduction).

Final answer: Zinc electrode = negative (anode); current is carried by ions in solution and electrons in the external wire; anode reaction Zn → Zn2+ + 2e-, cathode reaction 2Ag+ + 2e- → 2Ag.

Class 11 Chemistry Chapter 7 – Notes and Extra Questions

Beyond these thirty exercise questions, students preparing for school tests and competitive exams should revise the oxidation-number rules, the four types of redox reactions (combination, decomposition, displacement and disproportionation), both balancing methods (oxidation-number and ion-electron), and the standard electrode potential series, since these form the conceptual backbone of electrochemistry (Class 12) and inorganic chemistry that follows this chapter. Practising extra numericals on limiting-reagent redox stoichiometry and half-reaction balancing in both acidic and basic media is the best way to build speed for board and entrance exams.