DPP – 11

BATTERIES AND CORROSION

Chapter: Electrochemistry • Complete Solutions

"Great results are not built in one day. They are built in quiet moments like this — when you choose to sit and try again. Start with Question 1."

Quick Answer Key

1. (1) 2. (3) 3. (3) 4. (3) 5. (1) 6. (4) 7. (4) 8. (4) 9. (1) 10. (3) 11. (2) 12. (2) 13. (3)

Quick Reference: Common Batteries

Battery	Type	Anode	Cathode	EMF
Dry cell (Leclanche)	Primary	Zn	MnO_2 / NH_4Cl paste	∼1.5 V
Mercury cell	Primary	Zn(Hg) amalgam	HgO	∼1.35 V
Lead storage	Secondary	Pb	PbO_2	∼2 V/cell
Ni–Cd	Secondary	Cd	$Ni(OH)_3$	∼1.2 V

TYPE–1 : Battery Components & Structure

1. In lead storage battery, cathode is made up of [NCERT Pg. 89]

Explanation

The **lead storage battery** (car battery) has:

Electrode	Material	Reaction (Discharge)
Anode (–)	Pb grid packed with Pb	$Pb \longrightarrow Pb^{2+} + {}_2e^-$
Cathode (+)	Pb grid packed with PbO_2	$PbO_2 + {}_4H^+ + SO_4^{2-} + {}_2e^- \longrightarrow PbSO_4 + {}_2H_2O$
Electrolyte	38% H_2SO_4 (aq)	—

The cathode is the **Pb grid packed with PbO_2** (lead dioxide). PbO_2 is the oxidising agent—it gets *reduced* (gains electrons) at the cathode. ✓

Approach & Analogy

Memory aid: In a lead storage battery, think "**Pb** at both ends". The difference is: Anode = plain Pb (grey, dull); Cathode = PbO_2 (dark chocolate brown). The brown one is the cathode. When both convert to white $PbSO_4$ on discharge, the battery is dead.

Answer

Option (1) — Cathode is **Pb grid packed with PbO_2**.

2. In a mercury cell, product at anode is [NCERT Pg. 88]

Explanation

In the **mercury cell**:

- **Anode:** Zn–Hg amalgam oxidises in KOH electrolyte:

$$Zn(Hg) + {}_2OH^- \longrightarrow ZnO(s) + H_2O + {}_2e^-$$

  Product at anode = **ZnO(s)**
- **Cathode:** HgO is reduced:

$$HgO(s) + H_2O + {}_2e^- \longrightarrow Hg(l) + {}_2OH^-$$

  Product at cathode = Hg(l)

The electrolyte is paste of KOH / ZnO. The key point: in alkaline medium, Zn forms ZnO (not $ZnSO_4$ as in acidic dry cell).

Approach & Analogy

Analogy: In the mercury cell, the anode (Zn) is like a sacrificial soldier—it oxidises and leaves behind ZnO as its "grave marker." The cathode (HgO) releases mercury liquid as it gets reduced—like melting into a silvery puddle. Anode product = ZnO, cathode product = Hg.

Answer

Option (3) — Product at anode of mercury cell is **ZnO(s)**.

3. Which of the following is a secondary cell?

Explanation

Cell	Type	Rechargeable?
Dry cell (Leclanche)	**Primary**	No
Mercury cell	**Primary**	No
Lead storage battery	**Secondary**	Yes

A **secondary cell** can be recharged by passing current in reverse—the electrode reactions are reversed, regenerating the original reactants. The lead storage battery is the classic example.

Dry cell and mercury cell are **primary cells**—once discharged, the reactants are consumed and cannot be regenerated by electrical means.

Approach & Analogy

Analogy: Primary cell = a *single-use* pen (write once, throw away). Secondary cell = a *refillable* pen (you can refill the ink = recharge). The lead battery is recharged every time you run the car's alternator!

Answer

Option (3) — **Lead storage battery** is the secondary cell.

TYPE–2 : Electrochemical Reactions in Batteries

4. The chemical species undergoing reduction in dry cell is

Explanation

In the **dry cell (Leclanche cell)**:

- **Anode (–):** Zn container is oxidised:

$$Zn(s) \longrightarrow Zn^{2+}(aq) + {}_2e^-$$

- **Cathode (+):** MnO_2 is **reduced** (gains electrons):

$$MnO_2(s) + NH_4\equiv(aq) + e^- \longrightarrow MnO(OH)(s) + NH_3(g)$$

Manganese goes from Mn^{4+} in MnO_2 to Mn^{3+} in MnO(OH)—this is reduction.

NH_4^+ acts as the electrolyte/oxidant helper but it is MnO_2 that formally undergoes reduction. Zn undergoes *oxidation* (not reduction). NH_3 is a product, not the reducing species.

Approach & Analogy

Quick rule: Reduction happens at the cathode. In a dry cell, the cathode is the central carbon rod surrounded by MnO_2 paste. Whatever is *at the cathode* undergoes reduction. That's MnO_2—its Mn goes from +4 to +3.

Answer

Option (3) — $\mathbf{MnO_2}$ undergoes reduction at the cathode of the dry cell.

5. The net reaction in Ni–Cd cell is

Explanation

In the **Nickel–Cadmium (Ni–Cd) cell**:

- **Anode (–):** Cd is oxidised:

$$Cd(s) + {}_2OH-(aq) \longrightarrow Cd(OH)_2(s) + {}_2e^-$$

- **Cathode (+):** $Ni(OH)_3$ (nickel oxyhydroxide) is reduced:

$$_2Ni(OH)_3(s) + {}_2e^- \longrightarrow {}_2Ni(OH)_2(s) + {}_2OH^-$$

Net reaction (add anode + cathode, cancel OH^-):

$$Cd(s) + {}_2Ni(OH)_3(s) \longrightarrow CdO(s) + {}_2Ni(OH)_2(s) + H_2O(l)$$

Note: Cd forms CdO (or $Cd(OH)_2$ depending on conditions; the net equation given in option (1) with CdO is the standard NCERT/JEE version).

Approach & Analogy

Memory trick: In Ni–Cd: "**Cd goes up** (oxidised), **Ni comes down** (reduced)." Cd → CdO; $Ni(OH)_3 \rightarrow Ni(OH)_2$. The Ni drops one OH group (from 3 to 2) — like going down one floor in a building.

Answer

Option (1) $Cd + {}_2Ni(OH)_3 \longrightarrow CdO + {}_2Ni(OH)_2 + H_2O$

6. Rusting on the surface of iron involves

Explanation

Rusting of iron is an electrochemical process involving **three stages**:
Stage 1 — Anodic oxidation (iron dissolves):

$$Fe(s) \longrightarrow Fe^{2+}(aq) + {}_2e^- \quad \text{(at anodic sites)}$$

Stage 2 — Cathodic reduction (oxygen is reduced in presence of moisture):

$$O_2(g) + {}_4H{\equiv}(aq) + {}_4e^- \longrightarrow {}_2H_2O(l) \quad \text{(at cathodic sites)}$$

Stage 3 — Formation of rust (Fe_2O_3):

$${}_4Fe^{2+}(aq) + O_2(g) + {}_4H_2O(l) \longrightarrow {}_2Fe_2O_3(s) + {}_8H^+$$

Fe_2O_3 further hydrates to give the familiar red-brown rust ($Fe_2O_3 \cdot xH_2O$).
All three stages occur simultaneously—**all of these** are involved in rusting.

Approach & Analogy

Analogy: Rusting is like a tiny galvanic cell on the iron surface. Impurities or grain boundaries create micro-anodes and micro-cathodes. The anode (iron) dissolves (like a metal electrode); the cathode reduces oxygen (like the solution electrode). The Fe_2^+ then gets oxidised further by dissolved O_2 to form the red oxide. It's electrochemistry happening silently on your car bonnet!

Answer

Option (4) — Rusting involves **all three** reactions: anodic Fe dissolution, cathodic O_2 reduction, and Fe_2O_3 formation.

7. When a lead storage battery is recharged

Explanation

During **discharge**, both electrodes form $PbSO_4$ and H_2SO_4 is consumed:

$$\text{Discharge: } Pb + PbO_2 + {}_2H_2SO_4 \longrightarrow {}_2PbSO_4 + {}_2H_2O$$

During **recharging** (passing external current in reverse), the reaction reverses:

$$\text{Recharge: } {}_2PbSO_4 + {}_2H_2O \longrightarrow Pb + PbO_2 + {}_2H_2SO_4$$

So on recharging:
- Pb is regenerated at the anode (Option 1 — True)
- H_2SO_4 is regenerated—i.e. H_2SO_4 is *formed* (Option 2 — True)
- Option (3) says "H_2SO_4 is consumed"—this is what happens during *discharge*, not recharge. So option (3) is false.

Since both (1) and (2) are correct: **Option (4): Both (1) & (2)**.

Approach & Analogy

Key idea: Discharge = forward reaction (acid consumed, $PbSO_4$ formed). Recharge = reverse reaction (acid regenerated, Pb and PbO_2 reformed). The density of H_2SO_4 in the battery is a practical indicator—low density = discharged; high density = charged. Mechanics check battery acid density to assess charge state!

Answer

Option (4) — On recharging: **both Pb is formed AND H_2SO_4 is formed** (the discharge reaction runs in reverse).

8. Which of the following reactions occurs at the cathode?

Explanation

Cathode = site of **reduction** (gain of electrons). Check each option:
- Option (1): $_2OH^- \longrightarrow H_2O + O + {}_2e^-$ — electrons are *lost* = oxidation = **anode**. Wrong.
- Option (2): $Zn \longrightarrow Zn^{2+} + {}_2e^-$ — electrons are lost = oxidation = **anode**. Wrong.
- Option (3): $Sn^{2+} \longrightarrow Sn^{4+} + {}_2e^-$ — Sn loses electrons, goes to higher oxidation state = oxidation = **anode**. Wrong.
- Option (4): $Cu^{2+} + {}_2e^- \longrightarrow Cu$ — Cu^{2+} *gains* electrons and is deposited as Cu metal = **reduction** = cathode. Correct.

Approach & Analogy

Quick filter: Look for the arrow pointing *towards* a neutral species (gaining electrons), or any reaction where oxidation state *decreases*. In option (4), Cu goes from +2 to 0—that's a decrease, so it's reduction. This is also what happens in copper electroplating—Cu^{2+} from solution deposits as shiny copper metal on the cathode.

Answer

Option (4) — $Cu^{2+} + {}_2e^- \longrightarrow Cu$ is a **reduction reaction**, occurring at the cathode.

TYPE–3 : Corrosion & Protection Methods

9. In electrochemical corrosion of metals, the metal undergoing corrosion

Explanation

In electrochemical corrosion, the corroding metal **loses electrons**—it is *oxidised*.
- **Oxidation** occurs at the **anode**.
- The corroding metal is therefore the **anode**.
- It is *not* reduced (reduction happens at the cathode, where O_2 is reduced).

For iron rusting: $Fe \longrightarrow Fe^{2+} + {}_2e^-$ — iron is the anode; it gets corroded.

Approach & Analogy

Analogy: Think of corrosion as a slow "dissolving" of the metal. The metal gives away electrons (like a generous donor) to oxygen or H^+ in the environment. The giver is the anode. The receiver (oxygen) is at the cathode. The metal always acts as anode in corrosion.

Answer

Option (1) — The metal undergoing corrosion **acts as the anode** and is oxidised.

10. The most convenient method to protect the bottom of a ship made of iron is

Explanation

The hull of a ship is constantly in contact with seawater (an electrolyte) and is highly prone to corrosion. The method used is **cathodic protection** (sacrificial anode method):

- A more *active* (lower reduction potential) metal is attached to the iron hull.
- This metal acts as the **sacrificial anode**—it corrodes preferentially.
- Iron is forced to act as the cathode and is protected.
- **Mg** (magnesium) is ideal: it has a much lower reduction potential than Fe ($E^\circ(Mg^{2+}/Mg) = -2.37$ V vs. $E^\circ(Fe^{2+}/Fe) = -0.44$ V).

Evaluating options:

- Option (1): Red lead oxide coating—useful paint but not the most convenient for submerged hulls; requires maintenance.
- Option (2): Tin plating—tin is *less* active than iron; if the coating is scratched, Fe corrodes faster (galvanic corrosion).
- **Option (3): Connecting with Mg block**—cathodic protection using a sacrificial anode. Most convenient and reliable for ship hulls.
- Option (4): Pb block—Pb is less active than Mg and actually *less* active than Fe in some conditions; would not protect Fe effectively.

Approach & Analogy

Analogy: The Mg block is like a *bodyguard who takes the hit* for the iron hull. Mg is more eager to give away electrons than Fe, so the seawater "attacks" Mg first, leaving the iron unharmed. Sailors literally bolt Mg/Zn "sacrificial anodes" to the hull—visible as lumpy grey blocks on ship bottoms.

Answer

Option (3) — **Connecting with a Mg block** (cathodic/sacrificial anode protection) is the most convenient and effective method for ship hulls.

11. To protect iron against corrosion, the most suitable metal plating on it is

Explanation

To protect iron by plating, the coating metal must be **more active** than iron (more negative E°) so that even if the coating is scratched, the coating metal corrodes preferentially (cathodic protection of iron):

Plating Metal	E° **(V)**	**Effect on Fe if scratched**
Cu	+0.34	Cu is less active than Fe $\rightarrow$ Fe corrodes faster (**worse**)
Zn	**–0.76**	Zn is more active than Fe $\rightarrow$ Zn corrodes, Fe protected (**best**)
Ni	–0.25	Less active than Fe (–0.44 V) $\rightarrow$ Fe still exposed
Sn	–0.14	Less active than Fe $\rightarrow$ Fe corrodes if scratched

Zinc plating (galvanisation) is the answer. Even if the Zn coating is damaged, Zn (more reactive) acts as the sacrificial anode and protects the underlying iron.

Approach & Analogy

Real-world example: Galvanised iron sheets (zinc-coated) are used everywhere—roofing, buckets, fencing. Even if scratched, Zn preferentially corrodes. Tin cans (food tins) are tin-plated iron—but if scratched, iron corrodes *faster* because Sn is less active. That's why tin cans rust through quickly once dented, but galvanised iron lasts much longer!

Answer

Option (2) — **Zinc plating (galvanisation)** is most suitable. Zn ($E^\circ = -0.76$ V) is more active than Fe ($E^\circ = -0.44$ V) and provides cathodic protection.

TYPE–4 : Electrorefining & Cell Terminals

12. In the electrorefining of metals, impure metal is

Explanation

Electrorefining uses electrolysis to purify metals (e.g. copper refining):

- **Anode:** Impure metal (dissolves = oxidation)

$$Cu(impure) \longrightarrow Cu^{2+} + {}_2e^- \quad \text{(oxidation at anode)}$$

- **Cathode:** Pure metal deposits (reduction)

$$Cu^{2+} + {}_2e^- \longrightarrow Cu(pure) \quad \text{(reduction at cathode)}$$

- Impurities (less active metals like Au, Ag, Pt) fall as **anode mud** below the anode.

Therefore: Impure metal = **Anode** + **Oxidation** takes place there.

Approach & Analogy

Analogy: Electrorefining is like a "metal wash cycle." The impure lump (anode) slowly dissolves into the solution (like dirty clothes releasing dirt). Only the pure Cu^{2+} ions travel across and deposit as shiny clean metal on the cathode (clean clothes come out on the other side). The dirt (impurities) stays behind as anode mud.

Answer

Option (2) — Impure metal is the **anode**, and **oxidation** takes place at it.

13. Select the correct statement about cathode/anode polarity in galvanic and electrolytic cells

Explanation

This is a classic confusion point. Let us be systematic:

Cell Type	Cathode	Anode	Driving force
Galvanic (e.g. Daniel cell)	+ve terminal	–ve terminal	Spontaneous reaction
Electrolytic (e.g. electrolysis)	–ve terminal	+ve terminal	External power supply

Checking Option (3): "Cathode is negative terminal in electrolytic cell and anode is negative terminal in galvanic cell." This matches our table exactly. ✓

- Option (1): Says cathode is negative in *both*—wrong for galvanic (cathode is +ve).
- Option (2): Says anode is negative in *both*—wrong for electrolytic (anode is +ve).
- **Option (3): Correct.**
- Option (4): Reverses the assignment incorrectly.

Approach & Analogy

The one rule that never changes: Cathode = Reduction; Anode = Oxidation. This is true in *both* cell types. The polarity (+ or –) flips because:

In a **galvanic cell**, electrons flow *naturally* from anode to cathode through the wire—anode gives electrons (–), cathode receives them (+).

In an **electrolytic cell**, you *force* electrons in—the external battery's negative terminal pushes electrons to the cathode (making it –), and the positive terminal pulls electrons from the anode (making it +).

Think: Galvanic = natural flow (anode is –); Electrolytic = forced flow (cathode is –).

Answer

Option (3) — Cathode is the negative terminal in an electrolytic cell; Anode is the negative terminal in a galvanic cell.

Key Concepts to Remember — DPP 11

Concept	Key Point
Lead storage battery	Cathode = PbO_2; both electrodes $\rightarrow$ $PbSO_4$ on discharge
Mercury cell	Anode product = ZnO; Cathode product = Hg
Primary vs Secondary	Primary = non-rechargeable; Secondary = rechargeable (lead, Ni-Cd)
Dry cell cathode	MnO_2 is reduced ($Mn^{4+} \rightarrow Mn^{3+}$)
Ni–Cd net reaction	$Cd + {}_2Ni(OH)_3 \longrightarrow CdO + {}_2Ni(OH)_2 + H_2O$
Rusting	Anodic: Fe oxidised; Cathodic: O_2 reduced; then Fe_2O_3 forms
Ship protection	Sacrificial Mg block (cathodic protection)
Galvanisation	Zn plating on Fe; Zn more active so protects Fe even if scratched
Electrorefining	Impure metal = Anode (oxidised); Pure metal deposits at cathode
Cell polarity	Galvanic: Anode(–), Cathode(+); Electrolytic: Anode(+), Cathode(–)

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