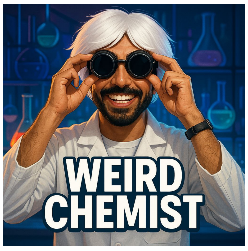

DPP–1 Solution [Working and Representation of Galvanic Cell]

Chapter: Electrochemistry

“If you are here after honestly trying, you are already ahead of most students. Now don’t just check the answer — understand the thinking. – Weird Chemist”

TYPE–1: Galvanic Cell – Electron/Current Flow & Basic Concepts

Solution 1

Explanation

In a galvanic cell, oxidation occurs at the anode and reduction occurs at the cathode. Electrons always flow from the anode (where oxidation happens) to the cathode (where reduction happens) through the external circuit.
For the cell $\mathrm{Cu(s)} \mid \mathrm{Cu^{2+}}\ (1\ \mathrm{M}) \parallel \mathrm{Ag^{+}}\ (1\ \mathrm{M}) \mid \mathrm{Ag(s)}$:
At Cu electrode: $\mathrm{Cu} \longrightarrow \mathrm{Cu^{2+}} + 2\,e^{-}$ (oxidation, so Cu is anode)
At Ag electrode: $\mathrm{Ag^{+}} + e^{-} \longrightarrow \mathrm{Ag}$ (reduction, so Ag is cathode)
Since Cu is the anode and Ag is the cathode, electrons flow from Cu to Ag in the external circuit.

Approach

Think of a galvanic cell like a water slide at a water park. Water (like electrons) naturally flows downhill from high to low. In electrochemistry, electrons flow from where they’re being produced (anode–oxidation) to where they’re being consumed (cathode–reduction).
Cu gives up electrons (becomes $\mathrm{Cu^{2+}}$) while Ag ions grab electrons (becomes Ag metal). So electrons travel from Cu electrode to Ag electrode through the wire, just like water flows from top to bottom of the slide.

Answer

Option (2): from Cu to Ag

Solution 2

Explanation

In a Daniell cell:
Anode (Zn): $\mathrm{Zn} \longrightarrow \mathrm{Zn^{2+}} + 2\,\mathrm{e^-}$
Cathode (Cu): $\mathrm{Cu^{2+}} + 2\,\mathrm{e^-} \longrightarrow \mathrm{Cu}$
Electrons flow from Zn to Cu in the external circuit (outside the cell).
Current, by convention, is the flow of positive charge, which is opposite to electron flow. Therefore, current flows from Cu to Zn in the external circuit.
The question asks about “passage of electricity” when the electrodes are connected, which refers to current flow in the external circuit.

Approach

Imagine electrons as negatively charged particles running from Zn to Cu through the wire (like kids running from one end of a playground to another). Now, “current” is defined as the direction positive charges would move. Since electrons are negative, current moves in the opposite direction–from Cu to Zn.
It’s like watching traffic on a one-way street: if cars (electrons) go left, then we say traffic (current) flows right in conventional terms.

Answer

Option (2): from Cu to Zn outside the cell

Solution 3

Explanation

For a Daniell cell:
Zn electrode (anode): $\mathrm{Zn} \longrightarrow \mathrm{Zn^{2+}} + 2\,\mathrm{e^-}$
Cu electrode (cathode): $\mathrm{Cu^{2+}} + 2\,\mathrm{e^-} \longrightarrow \mathrm{Cu}$
Let’s analyze each statement:
(1) Electrons flow from Cu to Zn – FALSE (electrons flow from Zn to Cu)
(2) Current flows from Zn to Cu – FALSE (current flows from Cu to Zn in external circuit)
(3) Cations move toward Cu electrode – TRUE (cathode attracts cations in the solution)
(4) Cations move toward Zn electrode – FALSE (anode attracts anions, not cations)
In the solution, cations ($\mathrm{Cu^{2+}}$, $\mathrm{Zn^{2+}}$) migrate toward the cathode (Cu) to maintain electrical neutrality as electrons are consumed there.

Approach

Think of the cell like a party with two rooms. In one room (Zn anode), people (electrons) are leaving, so positive ions (cations) from outside need to come in to keep things balanced. In the other room (Cu cathode), electrons are arriving, so positive ions (cations) also move there to balance the negative charge.
But which room do cations prefer? The cathode (Cu)! Because that's where reduction happens and positive ions are needed to balance the electrons being consumed.

Answer

Option (3): Cations move toward copper electrode

TYPE–2: Anode vs Cathode – Oxidation & Reduction

Solution 4

Explanation

An anodic reaction is an oxidation reaction (loss of electrons). Let's check each option:
(1) $Ag^+ \longrightarrow Ag - e^-$ – This is WRONG notation. It should be either $Ag^+ + e^- \longrightarrow Ag$ (reduction) or $Ag \longrightarrow Ag^+ + e^-$ (oxidation). As written, this suggests Ag^+ loses an electron, which is not an anodic reaction in the conventional sense.
(2) $Cu \longrightarrow Cu^{2+} + 2\,e^-$ – Oxidation (anodic)
(3) $Fe^{2+} \longrightarrow Fe^{3+} + e^-$ – Oxidation (anodic)
(4) $4\,OH^- \longrightarrow 2\,H_2O + O_2 + 4\,e^-$ – Oxidation (anodic)
Option (1) is not an anodic reaction because it represents reduction ($Ag^+ + e^- \longrightarrow Ag$), though written incorrectly.

Approach

Remember the phrase: "An Ox" and "Red Cat"
Anode = **Ox**idation
Reduction = **Cat**hode
Oxidation means losing electrons (think of it as "losing your wallet"–you become more positive). Look for reactions where electrons appear on the right side (products).
Option (1) shows Ag^+ gaining an electron to become Ag–that's like finding a wallet, making you richer (or more negative). That's reduction, not oxidation!

Answer

Option (1): $Ag^+ \longrightarrow Ag - e^-$

Solution 5

Explanation

The fundamental rules for electrochemical cells are:
Galvanic cell (spontaneous):
Anode: Oxidation occurs
Cathode: Reduction occurs
Electrolytic cell (non-spontaneous, driven by external voltage):
Anode: Oxidation occurs
Cathode: Reduction occurs
The key point: **In BOTH types of cells, oxidation occurs at the anode and reduction occurs at the cathode.** The difference is which electrode is positive/negative:
Galvanic: Anode is negative (−), Cathode is positive (+)
Electrolytic: Anode is positive (+), Cathode is negative (−)
But the chemistry (oxidation at anode, reduction at cathode) remains the same.

Approach

Think of "anode" and "cathode" not as positive or negative terminals, but as job descriptions:
Anode's job: Always host oxidation (like a factory that produces electrons)
Cathode's job: Always host reduction (like a factory that consumes electrons)
Whether it's a battery powering your phone (galvanic) or charging your phone (electrolytic), the anode is still the oxidation site and the cathode is still the reduction site. The roles don't switch–only the signs of the electrodes switch!

Answer

Option (1): Oxidation occur at anode in both galvanic and electrolytic cell

Solution 6

Explanation

For the cell: Pt | H_2(g) 1 atm | H^+(1M) || Cu^{2+} (1M) | Cu(s)
The half-reactions are:
At H_2 electrode: $H_2 \longrightarrow 2\,H^+ + 2\,e^-$ (oxidation, anode)
At Cu electrode: $Cu^{2+} + 2\,e^- \longrightarrow Cu$ (reduction, cathode)
Let's check each statement:
(1) H_2 is cathode, Cu is anode – FALSE (opposite is true)
(2) Oxidation occurs at Cu electrode – FALSE (reduction occurs at Cu)
(3) Reduction occurs at H_2 electrode – FALSE (oxidation occurs at H_2)
(4) H_2 is anode, Cu is cathode – TRUE

Approach

To find which is anode and which is cathode, look at what's happening:
H_2 gas is losing electrons (giving them away) to become H^+ – this is like donating blood, you're losing something. That's oxidation = anode.
Cu^{2+} ions are gaining electrons to become Cu metal – this is like receiving blood, you're gaining something. That's reduction = cathode.
The donor (H_2) is the anode, the receiver (Cu) is the cathode.

Answer

Option (4): H_2 is anode, Cu is cathode

TYPE–3: Salt Bridge – Function & Selection

Solution 7

Explanation

The cell given is: $Ag/Ag^+(0.1\ M)\ ||\ Ag^+/Ag\ (1M)$
This is a concentration cell with Ag/Ag^+ in both half-cells.
Salt bridge electrolytes should not react with the ions in either half-cell. Since both half-cells contain Ag^+ ions, we must avoid using salts that form precipitates with Ag^+.
$NaNO_3$: $Ag^+ + NO_3^-$ is soluble – CAN be used
KCl: $Ag^+ + Cl^- \rightarrow AgCl(s)$ precipitate – CANNOT be used
KNO_3: $Ag^+ + NO_3^-$ is soluble – CAN be used
KCl cannot be used because Cl^- ions would precipitate with Ag^+ ions.

Approach

Think of the salt bridge like a messenger service between two offices. The messengers (K^+, Cl^-, NO_3^-) need to travel back and forth without causing problems.
If you use KCl, the Cl^- messengers will react with Ag^+ in the offices and form a solid (AgCl)–like messengers getting stuck in traffic! This clogs up communication.
NO_3^- ions are like professional messengers–they don't react with Ag^+, so they can move freely without getting stuck.

Answer

Option (2): KCl

Solution 8

Explanation

The functions of a salt bridge are:
a. It completes the cell circuit – TRUE. Without the salt bridge, the circuit is incomplete and current cannot flow.
b. It allows movement of ions from one solution to other without mixing – TRUE. The salt bridge allows ion migration while preventing direct mixing of the two electrode solutions.
c. It maintains electrical neutrality – TRUE. As oxidation occurs at anode (producing cations), anions from salt bridge migrate to maintain neutrality. As reduction occurs at cathode (consuming cations), cations from salt bridge migrate to maintain neutrality.
All three functions are correct.

Approach

Imagine the salt bridge as a hallway connecting two rooms:
Function a: The hallway (salt bridge) allows people (ions) to walk between rooms, completing the connection. Without it, the rooms are isolated.
Function b: The hallway has a one-way door–people can pass through, but furniture and large items (bulk solutions) cannot mix.
Function c: If one room gets too crowded with men (cations being produced), women (anions) from the hallway enter to balance things out. If another room loses men (cations being consumed), men from the hallway enter to balance. This keeps both rooms balanced.

Answer

Option (2): a, b & c

Solution 9

Explanation

For a salt bridge to function properly, the cation and anion should have similar ionic mobilities (migration velocities). This ensures:
Equal transport numbers
Minimal liquid junction potential
No accumulation of charge on either side
K^+ and $NO_3{}^-$ ions have nearly equal velocities (ionic mobilities), making KNO_3 an ideal choice for salt bridges.
While high solubility helps in preparing saturated solutions, it is not the primary reason for using KNO_3.

Approach

Imagine K^+ and NO_3^- as a pair of runners in a relay race. If one runner is much faster than the other, they'll be out of sync–one side of the bridge will have too many fast ions, the other too many slow ions.
But K^+ and NO_3^- run at nearly the same speed, so they stay balanced. When K^+ moves right, NO_3^- moves left at the same rate, keeping both sides electrically neutral without creating a charge buildup.
It's like having two perfectly synchronized swimmers–they maintain balance in the pool.

Answer

Option (3): Velocity of both K^+ and NO_3^- are nearly the same

Solution 10

Explanation

Let's analyze each statement:
Statement 1: If salt bridge is removed, potential falls to zero – TRUE. Without the salt bridge, the circuit is incomplete, charge accumulates at the electrodes, and the cell voltage drops to zero (or very close to it) because the reaction stops.
Statement 2: KCl cannot be used in salt bridge with silver electrode – TRUE. As explained in Solution 7, Cl^- ions form AgCl precipitate with Ag^+ ions, interfering with cell operation.
Statement 3: Cell reaction is spontaneous if emf is positive – TRUE. A positive cell potential ($E_{\text{cell}} > 0$) indicates a spontaneous reaction ($\Delta G < 0$).
All statements are correct.

Approach

Think of these statements as three safety rules for electrochemical cells:
Rule 1: Remove the salt bridge = cell dies. It's like unplugging a device–no current flows, voltage drops to zero.
Rule 2: Don't use KCl with silver = recipe for disaster. It's like mixing oil and water–they react and create a mess (precipitate).
Rule 3: Positive EMF = spontaneous reaction. It's like a ball rolling downhill–it happens naturally without pushing (positive EMF means energy is released).

Answer

Option (4): All of these

TYPE–4: Cell Notation & Cell Reactions

Solution 11

Explanation

The given reaction is:

$$2\,\mathrm{AgCl(s)} + \mathrm{H_2(g)} \longrightarrow 2\,\mathrm{HCl(aq)} + 2\,\mathrm{Ag(s)}$$

Breaking this into half-reactions:
Oxidation (anode): $\mathrm{H_2} \longrightarrow 2\,\mathrm{H^+} + 2\,\mathrm{e^-}$
Reduction (cathode): $2\,\mathrm{AgCl} + 2\,\mathrm{e^-} \longrightarrow 2\,\mathrm{Ag} + 2\,\mathrm{Cl^-}$
In cell notation:
Anode on left, cathode on right
Single line (|) separates phases
Double line (||) represents salt bridge
Pt is used as inert electrode for H_2 gas
The complete notation:

$$\mathrm{Pt(s)|H_2(g), 1\ bar|1\ M\ HCl(aq)|AgCl(s)|Ag(s)}$$

Note: HCl provides H^+ for the anode reaction and the system naturally forms the product.

Approach

Cell notation is like writing an address–you follow a specific format from left to right:
Left side (Anode): Where oxidation happens
Start with the electrode material: Pt (since H_2 needs a surface)
Then the reactant gas: $H_2(g)$
Then the solution: HCl (provides H^+)
Right side (Cathode): Where reduction happens
The reactant: AgCl
The product: Ag
Think of | as "and" and || as "connected via salt bridge". You're building a sentence: Pt and H_2 and HCl || AgCl and Ag.

Answer

Option (3): Pt(s)|$\mathbf{H_2}$(g), 1bar|1M HCl(aq)|AgCl(s)|Ag(s)

Solution 12

Explanation

The cell notation is: A | A^-(x M) || B^+(y M) | B
The measured EMF is +0.20 V (positive).
In cell notation:
Left electrode (A) is the anode (oxidation)
Right electrode (B) is the cathode (reduction)
Since EMF is positive, the reaction proceeds as written. Looking at standard conventions and the fact that we have metal A on the left and metal B on the right:
At anode: $A \longrightarrow A^+ + e^-$ (A loses electrons–oxidation)
At cathode: $B^+ + e^- \longrightarrow B$ (B^+ gains electrons–reduction)
Overall reaction: $A + B^+ \longrightarrow A^+ + B$
Note: There appears to be a typo in the cell notation where A^- is shown, but based on standard electrochemistry and option (4), it should represent A^+ or the notation represents metal A in contact with solution containing ions.

Approach

Reading a cell is like reading a comic strip–left to right tells the story:
Left side (Anode): A loses electrons (gets oxidized)
Right side (Cathode): B^+ gains electrons (gets reduced)
Since the EMF is positive (+0.20 V), this reaction is spontaneous–it happens naturally, like a ball rolling downhill.
The overall story: A gives electrons to B^+, so A becomes A^+ (loses electron) and B^+ becomes B (gains electron).
Think of it as A donating to B: A becomes positive (donor), B becomes neutral (receiver).

Answer

Option (4): $\mathbf{A + B^+ \longrightarrow A^+ + B}$
(Note: There appears to be a typo in the cell notation where A^- should likely be A^+ to be consistent with standard cell notation and the answer.)