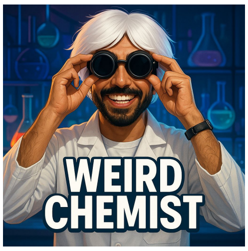

DPP–2 Solution [Electrode Potential and EMF]

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: Standard Hydrogen Electrode & Standard Potentials

Solution 1

Explanation

The standard reduction potential is defined for a reduction reaction under standard conditions:
Standard conditions: Temperature 25°C (298 K), Ion concentration 1 M, Pressure (for gases) 1 bar, State: Aqueous solution for ions, solid for metal.
For zinc, the standard reduction potential is defined by:

$$\mathrm{Zn^{2+}(aq) + 2\,e^- \longrightarrow Zn\,(s)}$$

This represents the reduction of aqueous zinc ions to solid zinc metal under standard conditions.

Approach

Think of standard conditions like a standardized test–everyone takes it under the same conditions so scores can be compared fairly. For electrode potentials, we need a "standard recipe": room temperature (25°C), ions dissolved in water (aq), 1 M concentration. Just like you wouldn't compare test scores if one student took it at home and another in a quiet room, we need standard conditions to compare electrode potentials.

Answer

Option (4): $\mathbf{Zn^{2+}(aq) + 2\,e^- \longrightarrow Zn\,(s)}$

Solution 2

Explanation

The standard hydrogen electrode (SHE) has an electrode potential of exactly 0.00 V by convention, not because of any inherent property of hydrogen. The SHE is chosen as the reference electrode, and all other electrode potentials are measured relative to it. It's an arbitrary choice–we could have chosen any electrode as the reference and assigned it zero potential.

Approach

Think of the standard hydrogen electrode like sea level for measuring altitude. Why is sea level called "zero elevation"? Not because oceans have special properties, but because we chose to call it zero as a reference point. Mount Everest is +8,849 m above sea level, and the Dead Sea is –430 m below sea level–all measured relative to our chosen zero reference. Similarly, SHE is our "zero reference" for electrode potentials.

Answer

Option (2): This electrode potential is assumed to be zero

Solution 3

Explanation

For a standard hydrogen electrode (SHE), the conditions are: H^+ ion concentration 1 M, Temperature 25°C, Pressure of H_2 gas 1 bar.
Option (4) is NOT true. The metallic conductor (typically platinum) must be able to adsorb hydrogen. This is essential because platinum provides a surface for the half-reaction, adsorbed H_2 molecules dissociate into H atoms, and these H atoms can then participate in the electrode reaction. Without hydrogen adsorption, the electrode wouldn't function properly.

Approach

Imagine the platinum electrode as a stage where the chemical reaction performs. The platinum doesn't just stand there–it actively participates by grabbing H_2 molecules from the gas (adsorption), breaking them apart into individual H atoms, and allowing them to react with H^+ ions in solution. If platinum couldn't adsorb hydrogen, it would be like having a stage with no actors–the show can't go on!

Answer

Option (4): It contains a metallic conductor which does not adsorb hydrogen

TYPE–2: EMF & Cell Potential Calculations

Solution 4

Explanation

The EMF of a cell can be calculated using: $E_{\text{cell}} = E^{\circ}_{\text{oxidation}}(\text{anode}) + E^{\circ}_{\text{reduction}}(\text{cathode})$ or $E_{\text{cell}} = E^{\circ}_{\text{oxidation}}(\text{anode}) - E^{\circ}_{\text{oxidation}}(\text{cathode})$. Both (b) and (d) are correct.

Approach

Think of EMF calculation like height difference between two points. Method 1: EMF = How much anode goes down + How much cathode goes up. Method 2: EMF = How much anode goes down – How much cathode goes down. Both give the same answer!

Answer

Option (3): (b) and (d)

Solution 5

Explanation

Given: $E^{\circ}(\text{Ni}^{2+}/\text{Ni}) = -0.25$ V and $E^{\circ}(\text{Ag}^{+}/\text{Ag}) = +0.80$ V.
Cell: Ni | Ni^{2+}(1.0 M) || Ag^{+}(1.0 M) | Ag
$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = 0.80 - (-0.25) = 1.05$ V

Approach

Ag is at +0.80 V (high on a hill), Ni is at –0.25 V (below sea level). Height difference = 0.80 – (–0.25) = 1.05 V. Just like water flowing from a mountain to a valley releases energy!

Answer

Option (3): +1.05 volt

Solution 6

Explanation

Given: $\text{Fe}^{3+} + \text{e}^{-} \longrightarrow \text{Fe}^{2+}$; $E^{\circ} = 0.771$ V and $\text{I}_2 + 2\,\text{e}^{-} \longrightarrow 2\,\text{I}^{-}$; $E^{\circ} = 0.536$ V
Cell reaction: $2\,\text{Fe}^{3+} + 2\,\text{I}^{-} \longrightarrow 2\,\text{Fe}^{2+} + \text{I}_2$
Standard electrode potentials are intensive properties–they don't change with stoichiometric coefficients.
$E^{\circ}_{\text{cell}} = 0.771 - 0.536 = 0.235$ V

Approach

Electrode potential is like temperature–it's intensive. If you have 1 liter of water at 50°C and mix it with another liter at 50°C, you get 2 liters at 50°C, not 100°C! Similarly, even though the equation shows 2 Fe^{3+}, the potential stays 0.771 V.

Answer

Option (3): 0.771 – 0.536 = 0.235 V

Solution 7

Explanation

Cell: Ni | Ni^{2+}(1.0 M) || Au^{3+}(1.0 M) | Au
$E^\circ(Ni^{2+}/Ni) = -0.25$ V and $E^\circ(Au^{3+}/Au) = 1.50$ V
$E^\circ_{cell} = 1.50 - (-0.25) = 1.75$ V

Approach

Au is on step +1.50 (very high), Ni is on step –0.25 (below ground). Total steps between them = 1.75 V. Electrons fall from Ni to Au, releasing 1.75 V!

Answer

Option (3): 1.75 V

Solution 8

Explanation

Given: $E^\circ(Fe^{2+}/Fe) = -0.441$ V and $E^\circ(Fe^{3+}/Fe^{2+}) = 0.771$ V
Reaction: $Fe + 2\,Fe^{3+} \longrightarrow 3\,Fe^{2+}$
Anode: $Fe \longrightarrow Fe^{2+} + 2\,e^-$; Cathode: $Fe^{3+} + e^- \longrightarrow Fe^{2+}$
$E^\circ_{cell} = 0.771 - (-0.441) = 1.212$ V

Approach

Two staircases meeting at Fe^{2+}: Fe to Fe^{2+} = climb 0.441 V, Fe^{3+} to Fe^{2+} = descend 0.771 V. Total height = 0.441 + 0.771 = 1.212 V.

Answer

Option (3): 1.212 V

Solution 9

Explanation

Cell: $\mathrm{Zn} \mid \mathrm{Zn^{2+}}(1\mathrm{M}) \parallel 2\,\mathrm{H^+}(1\mathrm{M}) \mid \mathrm{H_2}(1\ \mathrm{atm})$
$E^\circ(\mathrm{Zn^{2+}/Zn}) = -0.76$ V, SHE $= 0.00$ V
$E^\circ_{\text{cell}} = 0.00 - (-0.76) = +0.76$ V

Approach

SHE is sea level (0.00 V), Zn is in a valley at –0.76 V. Altitude difference = 0.76 V.

Answer

Option (2): 0.76 V

Solution 10

Explanation

$E^\circ(\mathrm{Pb^{2+}/Pb}) = -0.126$ V and $E^\circ(\mathrm{Zn^{2+}/Zn}) = -0.763$ V
Cell: $\mathrm{Zn} \mid \mathrm{Zn^{2+}}(0.1\ \mathrm{M}) \parallel \mathrm{Pb^{2+}}(1\ \mathrm{M}) \mid \mathrm{Pb}$
$E^\circ_{\text{cell}} = -0.126 - (-0.763) = 0.637$ V

Approach

Pb is at –0.126 V (slightly below sea level), Zn is at –0.763 V (much deeper). Depth difference = 0.637 V.

Answer

Option (1): 0.637 V

Solution 11

Explanation

Cell: $\mathrm{Ag(s)} \mid \mathrm{Ag^+} \parallel \mathrm{Cu^{2+}} \mid \mathrm{Cu(s)}$
Given reduction potentials: $\mathrm{Ag^+ + e^- \longrightarrow Ag}$; $E^\circ = x$ and $\mathrm{Cu^{2+} + 2\,e^- \longrightarrow Cu}$; $E^\circ = y$
If Ag is anode and Cu is cathode: $E^\circ_{\text{cell}} = y - x$

Approach

Cathode (Cu) at height y, Anode (Ag) at height x. Voltage = Top – Bottom = $y - x$.

Answer

Option (3): y − x

Solution 12

Explanation

Cell: Al | Al^{3+}(1 M) || Cu^{2+}(1M) | Cu
Given oxidation potentials: $E°(Al/Al^{3+}) = 1.66$ V and $E°(Cu/Cu^{2+}) = -0.34$ V
Converting to reduction: $E°(Al^{3+}/Al) = -1.66$ V and $E°(Cu^{2+}/Cu) = +0.34$ V
$E_{cell} = 0.34 - (-1.66) = 2.00$ V

Approach

Flip oxidation potentials to reduction: Al: –1.66 V, Cu: +0.34 V. Height difference = 2.00 V.

Answer

Option (1): 2.00 V

Solution 13

Explanation

Cell: Zn | Zn^{2+}(1M) || $2\,H^{+}$(1M) | H_2(1 atm)
$E_{cell} = 0.00 - (-0.76) = +0.76$ V (Same as Solution 9)

Approach

Same logic: Zn at –0.76 V, hydrogen at 0.00 V. Voltage = 0.76 V.

Answer

Option (2): +0.76 V

Solution 14

Explanation

$E°(Fe^{3+}/Fe^{2+}) = +0.77$ V and $E°(Sn^{2+}/Sn) = -0.14$ V
Reaction: $Sn(s) + 2\,Fe^{3+}(aq) \longrightarrow 2\,Fe^{2+}(aq) + Sn^{2+}(aq)$
$E°_{cell} = 0.77 - (-0.14) = 0.91$ V

Approach

Fe^{3+}/Fe^{2+} at +0.77 V (above sea level), Sn at –0.14 V (below). Total = 0.91 V.

Answer

Option (1): 0.91 V

Solution 15

Explanation

Button cell: $Zn(s) + Ag_2O(s) + H_2O \longrightarrow 2\,Ag(s) + Zn^{2+} + 2\,OH^-$
$E^\circ(Zn^{2+}/Zn) = -0.76$ V and $E^\circ(Ag_2O/Ag) = 0.34$ V
$E^\circ_{cell} = 0.34 - (-0.76) = 1.10$ V

Approach

Silver oxide at +0.34 V, Zinc at –0.76 V. This 1.10 V powers your watch!

Answer

Option (2): 1.10 V

Solution 16

Explanation

Given: (a) Zn–Cu cell: 1.10 V, (b) Cu–Ag cell: 0.46 V, $E^\circ(Cu^{2+}/Cu) = 0.34$ V
From (a): $1.10 = 0.34 - E^\circ_{Zn}$, so $E^\circ_{Zn} = -0.76$ V
From (b): $0.46 = E^\circ_{Ag} - 0.34$, so $E^\circ_{Ag} = 0.80$ V
For (c) Zn–Ag: $E^\circ_{cell} = 0.80 - (-0.76) = 1.56$ V

Approach

Like distances on a map: Zn→Cu = 1.10 V, Cu→Ag = 0.46 V. Total Zn→Ag = 1.10 + 0.46 = 1.56 V.

Answer

Option (4): 1.56 V

Solution 17

Explanation

Al displaces H (more reactive), Ag doesn't (less reactive). Cell EMF = 2.46 V, $E^\circ(Ag^+/Ag) = +0.80$ V
Cell: $Al \mid Al^{3+} \parallel Ag^+ \mid Ag$
$2.46 = 0.80 - E^\circ_{Al}$, so $E^\circ_{Al} = -1.66$ V

Approach

Al is more reactive (deeper underwater), Ag is less reactive (higher up). The 2.46 V gap tells us: Al = 0.80 – 2.46 = –1.66 V.

Answer

Option (3): –1.66 V

Solution 18

Explanation

For maximum cell potential, combine highest reduction potential with lowest reduction potential.
(i) $E_{\text{red}} = -1.5$ V, (ii) $E_{\text{red}} = 2.1$ V, (iii) $E_{\text{red}} = 0.5$ V, (iv) $E_{\text{red}} = 1.5$ V
Maximum: Cathode (ii) at 2.1 V, Anode (i) at –1.5 V
$E = 2.1 - (-1.5) = 3.6$ V

Approach

Biggest waterfall = highest point (2.1 V) to lowest point (–1.5 V). Height = 3.6 V maximum!

Answer

Option (2): (i) and (ii)

Solution 19

Explanation

Convert oxidation potentials to reduction: Zn: –0.76 V, Cu: +0.34 V, Ag: +0.80 V, H: 0 V, Ni: –0.55 V
Check all options: (1) Zn + Cu^{2+}: 0.34 – (–0.76) = 1.10 V (2) Zn + Ag^{+}: 0.80 – (–0.76) = 1.56 V ← Maximum! (3) H + Cu^{2+}: 0.34 V (4) H + Ni^{2+}: –0.55 V (non-spontaneous)

Approach

Pair deepest (Zn at –0.76) with highest (Ag at +0.80). Maximum voltage = 1.56 V!

Answer

Option (2): Zn + 2Ag^{+} -¿ 2Ag + Zn^{2+}

Solution 20

Explanation

$E^\circ(\mathrm{Sn^{4+}/Sn^{2+}}) = +0.15$ V and $E^\circ(\mathrm{Cr^{3+}/Cr}) = -0.74$ V
Cathode: $\mathrm{Sn^{4+}}$ reduction at +0.15 V, Anode: Cr oxidation at –0.74 V
$E^\circ_{\text{cell}} = 0.15 - (-0.74) = 0.89$ V

Approach

Sn at +0.15 V (above sea level), Cr at –0.74 V (below). Height = 0.89 V.

Answer

Option (2): +0.89 V

Solution 21

Explanation

Cell: Cu(s) | $\mathrm{Cu^{2+}}$(1 M) || $\mathrm{Zn^{2+}}$(1 M) | Zn(s)
$E^\circ(\mathrm{Cu^{2+}/Cu}) = +0.34$ V and $E^\circ(\mathrm{Zn^{2+}/Zn}) = -0.76$ V
From cell notation: Anode is Cu (+0.34 V), Cathode is Zn (–0.76 V)
$E_{\text{cell}} = -0.76 - 0.34 = -1.10$ V (negative, non-spontaneous)

Approach

This cell is upside down–trying to make electrons flow uphill! Cu at +0.34 V trying to send electrons to Zn at –0.76 V. That's backwards! EMF is negative (–1.10 V). Like trying to make water flow uphill–you'd need external energy.

Answer

Option (2): Negative