# what is Henderson-Hasselbalch equation.? in chemistry

the **Henderson-Hasselbalch equation** is a formula used to **calculate the pH of a buffer solution**. It relates the pH of a solution to the concentration of an acid and its conjugate base.

# what is weak acid?

A **weak acid** is an acid that **does not completely dissociate** into ions in water.

**⚗️ Definition:**

A **weak acid** only **partially ionizes** in aqueous solution, meaning only a small fraction of its molecules release hydrogen ions (**H⁺** or **protons**) when dissolved.

**🧪 General Example:**

For a weak acid **HA**:

HA ⇌ H⁺ + A⁻

* The **double arrow (⇌)** shows that the reaction is reversible.
* Only some of the HA molecules ionize, while others remain intact.

**📉 Properties of Weak Acids:**

* **Low conductivity** (fewer free ions).
* **Higher pH** than strong acids at the same concentration (typically between pH 4–6).
* **Do not react completely** with bases or metals.
* Have a **small acid dissociation constant (Ka)**.

**🔍 Examples of Weak Acids:**

| **Weak Acid Name** | **Chemical Formula** |
| --- | --- |
| Acetic acid | CH₃COOH |
| Carbonic acid | H₂CO₃ |
| Citric acid | C₆H₈O₇ |
| Phosphoric acid | H₃PO₄ |
| Formic acid | HCOOH |

Let me know if you'd like a **comparison between strong and weak acids**, or if you want to see how weak acids behave in chemical reactions.

Great! Here's a **clear comparison between strong and weak acids**:

**🧪 Strong Acid vs Weak Acid**

| **Feature** | **Strong Acid** | **Weak Acid** |
| --- | --- | --- |
| **Ionization in Water** | Completely ionizes | Partially ionizes |
| **Reaction Arrow** | → (one-way, complete reaction) | ⇌ (reversible, incomplete reaction) |
| **Electrical Conductivity** | High (many free ions) | Low to moderate (fewer free ions) |
| **pH of Solution** | Very low (1–3) | Higher pH (around 4–6) |
| **Acid Dissociation Constant (Ka)** | Large (>>1) | Small (<1) |
| **Strength of Conjugate Base** | Very weak | Relatively stronger |
| **Examples** | HCl, HNO₃, H₂SO₄ | CH₃COOH, H₂CO₃, H₃PO₄ |

**⚗️ Equations Example**

* **Strong Acid (HCl):**  
  HCl → H⁺ + Cl⁻  
  (100% ionized in water)
* **Weak Acid (Acetic acid - CH₃COOH):**  
  CH₃COOH ⇌ H⁺ + CH₃COO⁻  
  (only partially ionized)

**💡 Quick Tip to Remember:**

**Strong acids**: “All in!”  
**Weak acids**: “Some in, some stay.”

Would you like a **visual diagram**, or explanation of how this affects **titration curves** or **buffer solutions**?

# Electrolyte?

In chemistry, an electrolyte is a substance that, when dissolved in a solvent (usually water), dissociates into ions, which are charged particles, and can conduct electricity

# what is weak electrolyte

**⚗️ What is a Weak Electrolyte?**

A **weak electrolyte** is a substance that **partially dissociates** into ions when dissolved in water. Because of this, it conducts **electricity poorly** in an aqueous solution.

**✅ Definition:**

A **weak electrolyte** is a compound that:

* **Dissolves in water**
* **Only a small fraction** of its molecules **ionize** (break into ions)
* Forms a solution that is a **poor conductor** of electricity

**🔬 Examples:**

| **Substance** | **Type** |
| --- | --- |
| Acetic acid (CH₃COOH) | Weak acid & weak electrolyte |
| Ammonia (NH₃) | Weak base & weak electrolyte |
| Carbonic acid (H₂CO₃) | Weak acid & weak electrolyte |

**⚡ Reaction Example:**

For **acetic acid** in water:

CH₃COOH⇌CH₃COO⁻+H⁺\text{CH₃COOH} \rightleftharpoons \text{CH₃COO⁻} + \text{H⁺}

Only a small % of CH₃COOH breaks into ions.

**🧪 Comparison:**

| **Type** | **Ionization** | **Conductivity** |
| --- | --- | --- |
| Strong electrolyte | 100% | High |
| Weak electrolyte | Partial | Low |
| Nonelectrolyte | 0% (no ions) | None |

Let me know if you'd like visuals or a quiz to test your understanding!

# weak electrolyte:

chemical substance which does not ionize readily partially and furnishes an less amount of iones either in aqueous medium or in molten state is called weak electrolyte

explain me in simple word .

Sure! Let me explain your sentence in **simple words**:

**✅ Simple Explanation:**

A **weak electrolyte** is a chemical that:

* **Does not break into ions completely** when mixed with water or melted
* It **makes only a few ions**, not many
* Because of that, it **doesn't conduct electricity well**

**🔍 Breaking it down:**

* **"Does not ionize readily"** → It doesn’t break into ions easily
* **"Partially ionizes"** → Only **some** of the molecules turn into ions
* **"Furnishes less amount of ions"** → Makes **few ions**
* **"In aqueous or molten state"** → In water or melted form

**🔌 Simple Example:**

* **Vinegar (acetic acid)** in water  
  → Only a small part breaks into ions  
  → So, electricity flows **weakly**

Would you like a comparison with **strong electrolytes** next?

# what is degree of ionization.

**⚛️ What is Degree of Ionization?**

**Degree of ionization** tells us **how much of a substance turns into ions** when it's dissolved in water (or in molten form).

**✅ Simple Definition:**

The **degree of ionization** is the **percentage or fraction of molecules** that **break into ions** in a solution.

**🔣 Formula:**

Degree of Ionization(α)=Number of ionized moleculesTotal number of molecules\text{Degree of Ionization} (\alpha) = \frac{\text{Number of ionized molecules}}{\text{Total number of molecules}}

**📊 Examples:**

| **Substance** | **Degree of Ionization** | **Type** |
| --- | --- | --- |
| **HCl (strong acid)** | Nearly 100% | Strong electrolyte |
| **CH₃COOH (weak acid)** | Less than 5% | Weak electrolyte |

**💡 In Simple Words:**

If you put 100 molecules of a substance in water:

* If **all 100** become ions → degree = **1** (or 100%)
* If **only 10** become ions → degree = **0.1** (or 10%)

**⚡Why it matters:**

* Higher degree = more ions = better electricity flow
* Lower degree = fewer ions = weak electricity flow

Let me know if you'd like a visual or quiz to test this!

# Repetative question unit-2

Unit 2 of the Chemistry syllabus is dedicated to **Electrochemistry**. Based on the provided sources, several questions from this unit are repetitive across various examination papers, model questions, and assignments:

* **Differentiation between Electrolytic Cell and Galvanic Cell**
  + **Question 13(a)**: "Differentiate between electrolytic and Galvanic cell."
    - **2024** Purabanchal University Exam Paper.
  + **Question 5**: "Differentiate between electrolytic cell and galvanic cell. Give the function of salt bridge."
    - Kantipur City College Unit Test Chemistry.
  + **Question 5**: "Differentiate between electrolytic cell and galvanic cell. Give the function of salt bridge."
    - Assignment (electrochemistry).
* **Nernst Equation (Derivation and/or Numerical Problems)**
  + **Question 12**: "Derive Nernst’s equation. Calculate the emf of the electrode couple of E°Zn/Sn+2 = -1.4V and E° Fe+2/Fe+3. Where the concentration of Sn+2, Fe+2 and Fe+3 is 0.2M, 0.1M, and 1M respectively."
    - **2024** Purabanchal University Exam Paper.
  + **Question 12**: "Derive Nernst’s equation. Calculate the emf of the Zn/Ag electrochemical cell at 39°C, when the concentration of Zn++ and Ag+ are 0.15M and 0.5M respectively. Given E°Zn++/Zn = -0.76V and E°Ag+/Ag =0.80".
    - **2022** Purabanchal University Model Question Paper.
  + **Question 2(b)**: "Calculate the emf of the galvanic cell made by coupling two half cells viz, Cr/Cr+3 (0.5M) and Fe/Fe+2 (0.2M). Given that E°Cr/Cr+3 =0.75V and E°Fe+2/Fe=0.44V also write the cell reaction."
    - **2019** Purabanchal University Exam Paper.
  + **Question 6**: "Derive Nernst’s equation. Also derive the expression for half-cell potential from it."
    - Assignment (electrochemistry).
  + Related numerical problems often appear across various papers, requiring application of the Nernst equation to calculate cell EMF under non-standard conditions.
* **Corrosion (Electrochemical Theory and Prevention/Short Notes)**
  + **Question 18**: "What do you mean by corrosion? Explain the electrochemical theory of rusting iron? Write down the preventive measures of rusting."
    - Assignment (electrochemistry).
  + **Question 4(b)**: "Write short notes on Corrosion of metal".
    - **2018** Purabanchal University Exam Paper.
  + The syllabus specifically covers the "Corrosion of metal (electrochemical theory of rusting of iron), electrochemical theory and its applications and prevention of rusting".
* **Standard Electrode Potential / Standard Hydrogen Electrode (SHE)**
  + **Question 12**: "What is standard electrode potential? How can you measure the standard electrode potential of Copper? Calculate the cell potential at 25°C; Zn (s) / Zn++(aq, 0.3M) // Cu++(aq, 0.1M) / Cu (s). Given E°Zn++/Zn = -0.76 V and E°Cu++/Cu=+0.34V".
    - **2023** Purabanchal University Exam Paper.
  + **Question 4**: "What do you mean by standard and non-standard electrode potential? Describe the method to determine standard electrode potential of copper when it is constructed as anode."
    - Assignment (electrochemistry).
  + **Question 3**: "What is meant by standard hydrogen electrode? Calculate the emf of the following cell at 25oC. Mg/Mg++ (0.1M)// Ag+(1M)/Ag Given, a. E° Mg++/Mg = -2.37V b. E° Ag+/Ag = +0.8V".
    - Assignment (electrochemistry).
* **Function of Salt Bridge**
  + **Question 5**: "Differentiate between electrolytic cell and galvanic cell. Give the function of salt bridge."
    - Kantipur City College Unit Test Chemistry.
  + **Question 5**: "Differentiate between electrolytic cell and galvanic cell. Give the function of salt bridge."
    - Assignment (electrochemistry).
  + **Question 8**: "Define electrode potential. Give the function of salt bridge."
    - Assignment (electrochemistry).
  + The functions of a salt bridge, such as completing the electrical circuit and maintaining electrical neutrality, are detailed in the sources.

## 2.1

Perfect 👌 Let’s make you a **cheat sheet** — short, clear, and exam-friendly.

**⚡ Electrochemistry Cheatsheet**

**Differentiation: Electrolytic Cell vs Galvanic Cell**

| **Feature** | **Electrolytic Cell** | **Galvanic Cell** |
| --- | --- | --- |
| **Definition** | Converts **electrical energy → chemical energy** | Converts **chemical energy → electrical energy** |
| **Spontaneity** | **Non-spontaneous** reaction (needs electricity) | **Spontaneous** reaction (produces electricity) |
| **Energy** | Requires external **battery/power supply** | Produces current (no external power needed) |
| **Electrodes Polarity** | Cathode = (–), Anode = (+) | Cathode = (+), Anode = (–) |
| **Example** | Electrolysis of water, NaCl electrolysis | Daniell cell (Zn–Cu cell), batteries |

**🌉 Function of Salt Bridge**

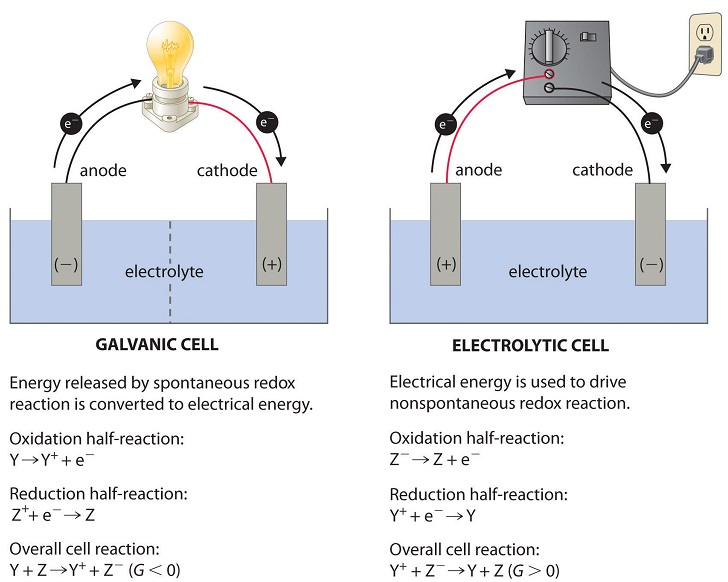
* Maintains **electrical neutrality** in galvanic cell.
* Allows **ion flow** (not electron flow).
* Prevents charge buildup by:
  + Cations (+) moving to cathode compartment.
  + Anions (–) moving to anode compartment.

👉 Without salt bridge → circuit breaks, no current flow.

✅ Quick Tip for Exams:

* **Electrolytic → needs battery.**
* **Galvanic → gives battery.**

Would you like me to also make a **one-page printable cheatsheet (PDF)** with a small diagram of both cells + salt bridge for last-minute revision?



## 2.3

Perfect 🚀 Let’s make a **cheatsheet** for **Nernst’s Equation** — short, exam-focused, and with step-by-step examples.

**⚡ Electrochemistry Cheatsheet – Nernst’s Equation**

**🔹 1. Nernst’s Equation (Derivation in short form)**

For a general reaction:

aA+bB  →  cC+dDaA + bB \; \to \; cC + dD

The cell potential at non-standard conditions:

E=E∘−0.0591nlog⁡[C]c[D]d[A]a[B]bE = E^\circ - \frac{0.0591}{n} \log \frac{[C]^c[D]^d}{[A]^a[B]^b}

* E∘E^\circ = standard cell potential
* nn = number of electrons transferred
* [][ ] = concentration of species
* At **T = 25°C (298 K)**, constant becomes **0.0591**

👉 At any temperature:

E=E∘−2.303RTnFlog⁡QE = E^\circ - \frac{2.303RT}{nF} \log Q

where Q=[Products][Reactants]Q = \frac{[Products]}{[Reactants]}

**🔹 2. Steps to Solve Exam Problems**

1. Write half-cell reactions (oxidation & reduction).
2. Find nn (electrons transferred).
3. Calculate **E°cell**:

Ecell∘=Ecathode∘−Eanode∘E^\circ\_\text{cell} = E^\circ\_\text{cathode} - E^\circ\_\text{anode}

1. Write reaction quotient QQ.
2. Apply Nernst’s equation.

**🔹 3. Example Problems**

**(i) Zn/Sn²⁺ vs Fe²⁺/Fe³⁺**

* Given: EZn/Sn2+∘=−1.4VE^\circ\_{Zn/Sn^{2+}} = -1.4V, EFe2+/Fe3+∘=?E^\circ\_{Fe^{2+}/Fe^{3+}} = ?
* [Sn²⁺] = 0.2 M, [Fe²⁺] = 0.1 M, [Fe³⁺] = 1 M.
* Find Ecell∘E^\circ\_\text{cell}, then use Nernst’s equation with n=1n = 1 or 2 depending on reaction.

**(ii) Zn/Ag Cell at 39°C**

* Cell: Zn | Zn²⁺(0.15M) || Ag⁺(0.5M) | Ag
* EZn2+/Zn∘=−0.76VE^\circ\_{Zn^{2+}/Zn} = -0.76V, EAg+/Ag∘=+0.80VE^\circ\_{Ag^+/Ag} = +0.80V
* Ecell∘=0.80−(−0.76)=1.56VE^\circ\_\text{cell} = 0.80 - (-0.76) = 1.56V
* n=2n = 2
* At T=39°C=312KT = 39°C = 312K, use

E=E∘−0.0591nlog⁡QE = E^\circ - \frac{0.0591}{n} \log Q

with Q=[Zn2+][Ag+]2Q = \frac{[Zn^{2+}]}{[Ag^+]^2}.

**(iii) Cr/Cr³⁺ vs Fe/Fe²⁺**

* Given: ECr/Cr3+∘=0.75VE^\circ\_{Cr/Cr^{3+}} = 0.75V, EFe2+/Fe∘=0.44VE^\circ\_{Fe^{2+}/Fe} = 0.44V.
* Concentrations: [Cr³⁺] = 0.5 M, [Fe²⁺] = 0.2 M.
* Ecell∘=0.75−0.44=0.31VE^\circ\_\text{cell} = 0.75 - 0.44 = 0.31V.
* Reaction:

Cr  +  3Fe2+  →  Cr3+  +  3FeCr \;+\; 3Fe^{2+} \;\to\; Cr^{3+} \;+\; 3Fe

* Apply Nernst’s equation with n=3n = 3.

**🔹 4. Key Exam Hints**

* Always **cathode – anode**.
* Write **oxidation & reduction half-reactions**.
* Identify **n** correctly.
* For QQ: products over reactants (exclude solids).
* If T ≠ 298K, use 2.303RT/F2.303RT/F.

✅ **Memory Aid:**

* **E° = standard** (fixed)
* **E = real** (depends on concentration)
* **Cell runs till Q = K (equilibrium)** → then E = 0

Would you like me to also **fully solve (step-by-step with numbers)** one of these questions (say Zn/Ag at 39°C), so you have a worked-out solution to copy in exams?

## 2.5

help me to make an cheat sheet on below topic. easy to revisoin before exam?

Got it 👍 You want a **concise cheat sheet** to revise quickly before exam. I’ll keep it short, clear, and exam-focused.

**📝 Cheat Sheet: Corrosion (Electrochemical Theory & Prevention)**

**1. Definition of Corrosion**

* **Corrosion**: Gradual destruction of metals due to **chemical/electrochemical reactions** with environment (air, moisture, acids, etc.).
* Example: Rusting of iron → formation of hydrated ferric oxide (Fe2O3⋅xH2O)(Fe₂O₃·xH₂O).

**2. Electrochemical Theory of Rusting (Iron)**

Rusting is an **electrochemical process** (requires water + O₂).

**Steps:**

1. **Anodic reaction (oxidation):**

Fe  →  Fe2++2e−Fe \;→\; Fe^{2+} + 2e^-

Iron loses electrons (corrodes at anodic sites).

1. **Cathodic reaction (reduction of oxygen in water):**

O2+4H++4e−  →  2H2OO₂ + 4H^+ + 4e^- \;→\; 2H₂O

*(In neutral/alkaline water:)*

O2+2H2O+4e−  →  4OH−O₂ + 2H₂O + 4e^- \;→\; 4OH^-

1. **Formation of rust:**
   * Fe2+Fe^{2+} → oxidized to Fe3+Fe^{3+}.
   * Combines with OH−OH^- → Fe(OH)3Fe(OH)₃.
   * Dehydrates → **Rust (Fe₂O₃·xH₂O)**.

**Conditions needed:**

* Presence of **moisture** + **oxygen**.
* Electrolytes (salts, acids) accelerate corrosion.

**3. Prevention of Rusting**

🔑 Methods to protect metal from corrosion:

1. **Barrier Protection** (block O₂ & H₂O contact):
   * Painting, oiling, greasing, plastic coating.
2. **Galvanization** (Zn coating on Fe).
3. **Alloying** (e.g., stainless steel with Cr, Ni).
4. **Cathodic Protection**:
   * **Sacrificial anode method**: Attach Mg, Zn, Al → corrode instead of Fe.
   * **Impressed current method**: External current makes Fe cathodic.
5. **Metallic Coatings**: Electroplating with Ni, Cr.
6. **Environmental control**: Use dry air, inhibitors, corrosion-resistant paints.

**4. Short Notes Points (for 4–5 marks)**

* **Corrosion of metal** = gradual destruction via oxidation/reduction.
* **Electrochemical theory**: Rusting is an anodic-cathodic process.
* **Rust composition**: Hydrated ferric oxide Fe2O3⋅xH2OFe₂O₃·xH₂O.
* **Prevention**: Barrier methods, galvanization, alloying, cathodic protection.

✅ **Quick Memory Aid**:

* **Anode**: Fe → Fe²⁺
* **Cathode**: O₂ reduction → OH⁻
* **Product**: Fe(OH)₃ → Rust
* **Prevention**: "BIG CAP"
  + **B**arrier
  + **I**nhibitors
  + **G**alvanizing
  + **C**athodic protection
  + **A**lloying
  + **P**lating

Do you want me to also **make this into a one-page PDF cheat sheet** (nice formatting with tables & diagrams for last-minute revision)?

## 2.2

help me to make an cheat sheet on below topic. easy to revisoin before exam?

Perfect — this is a classic **electrochemistry exam area**, so let’s make a **cheat sheet** that’s crisp, structured, and easy to revise in 5 minutes.

**⚡ Cheat Sheet: Standard Electrode Potential & SHE**

**1. Key Definitions**

* **Electrode potential (E):** Potential difference between a metal electrode and its ion solution.
* **Standard Electrode Potential (E°):** Electrode potential measured under *standard conditions*:
  + Ion concentration = 1 M
  + Gas pressure = 1 atm
  + Temperature = 25 °C (298 K)
* **Standard Hydrogen Electrode (SHE):**
  + Reference electrode with assigned potential = **0.00 V**.
  + Consists of **Pt electrode** in contact with 1 M H⁺ solution & H₂ gas at 1 atm.
  + Used to measure other electrode potentials.

**2. Measurement of Standard Electrode Potential (Example: Cu)**

1. Construct a cell:  
   **Pt | H₂ (1 atm) | H⁺ (1 M) || Cu²⁺ (1 M) | Cu**
2. Connect SHE as reference.
3. Measure emf (cell potential).
   * If Cu is **cathode** → reduction occurs: Cu²⁺ + 2e⁻ → Cu.
   * If Cu is **anode** → oxidation occurs: Cu → Cu²⁺ + 2e⁻.
4. The measured emf = **E°(Cu²⁺/Cu)** (since SHE = 0 V).

**3. Standard vs. Non-Standard Electrode Potential**

* **Standard (E°):** Measured at standard conditions.
* **Non-standard (E):** When conditions deviate (concentration ≠ 1 M, pressure ≠ 1 atm).
  + Calculated using **Nernst Equation**:

E=E°−0.0591nlog⁡QE = E° - \frac{0.0591}{n} \log Q

where *Q* = reaction quotient, *n* = number of electrons.

**4. Cell Potential (EMF)**

* For a cell:

Ecell=Ecathode∘−Eanode∘E\_\text{cell} = E^\circ\_\text{cathode} - E^\circ\_\text{anode}

* **Cathode** = reduction (more positive E°).
* **Anode** = oxidation (less positive E°).

**5. Worked Examples (Exam Style)**

**(a) Zn/Cu Cell (non-standard conc.)**

Cell: **Zn(s) | Zn²⁺(0.3 M) || Cu²⁺(0.1 M) | Cu(s)**

* E°Zn²⁺/Zn = –0.76 V
* E°Cu²⁺/Cu = +0.34 V

1. Standard EMF:

Ecell∘=0.34−(−0.76)=+1.10 VE^\circ\_\text{cell} = 0.34 - ( -0.76 ) = +1.10 \text{ V}

1. Reaction: Zn(s) + Cu²⁺(aq) → Zn²⁺(aq) + Cu(s)  
   n = 2
2. Nernst equation:

E=Ecell∘−0.05912log⁡[Zn2+][Cu2+]E = E^\circ\_\text{cell} - \frac{0.0591}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}

Plug in:

E=1.10−0.05912log⁡0.30.1E = 1.10 - \frac{0.0591}{2} \log \frac{0.3}{0.1} E≈1.10−0.02955log⁡3≈1.10−0.014≈1.086 VE ≈ 1.10 - 0.02955 \log 3 ≈ 1.10 - 0.014 ≈ 1.086 \, V

**(b) Mg/Ag Cell**

Cell: **Mg | Mg²⁺(0.1 M) || Ag⁺(1 M) | Ag**

* E°Mg²⁺/Mg = –2.37 V
* E°Ag⁺/Ag = +0.80 V

1. Standard EMF:

Ecell∘=0.80−(−2.37)=3.17 VE^\circ\_\text{cell} = 0.80 - ( -2.37 ) = 3.17 \, V

1. Overall reaction: Mg(s) + 2Ag⁺ → Mg²⁺ + 2Ag(s), n = 2
2. Nernst equation:

E=3.17−0.05912log⁡[Mg2+][Ag+]2E = 3.17 - \frac{0.0591}{2} \log \frac{[Mg^{2+}]}{[Ag^+]^2} E=3.17−0.02955log⁡0.112E = 3.17 - 0.02955 \log \frac{0.1}{1^2} E=3.17−0.02955(−1)=3.17+0.0296≈3.20 VE = 3.17 - 0.02955 (-1) = 3.17 + 0.0296 \approx 3.20 \, V

**6. Quick Mnemonics**

* **SHE conditions:** “**1–1–25**” → 1 atm, 1 M, 25°C.
* **Cell EMF formula:** *Cathode – Anode*.
* **AnOx – RedCat:** **Anode = Oxidation, Cathode = Reduction**.

✅ This is exam-ready.  
Do you want me to also **make this into a one-page PDF cheat sheet** (with table + formulas highlighted) for quick last-night revision?

### "What is standard electrode potential? How can you measure the standard electrode potential of Copper? Calculate the cell potential at 25°C; Zn (s) / Zn++(aq, 0.3M) // Cu++(aq, 0.1M) / Cu (s). Given E°Zn++/Zn = -0.76 V and E°Cu++/Cu=+0.34V".

Perfect — let’s make a compact **cheat sheet** for this exam-style question. I’ll break it into 3 parts: **definition, measurement, and calculation**.

**🔑 CHEAT SHEET: Standard Electrode Potential & Cell Potential**

**1. Standard Electrode Potential (E°)**

* **Definition:** The potential of an electrode measured against the **Standard Hydrogen Electrode (SHE)** under standard conditions:
  + [ion] = 1 M
  + Pressure = 1 atm
  + Temperature = 25 °C
* It shows how easily an element can be oxidized or reduced.
  + More **positive E°** → stronger oxidizing agent (likes to gain electrons).
  + More **negative E°** → stronger reducing agent (likes to lose electrons).

**2. Measurement of E° (Copper example)**

* Set up an electrochemical cell:
  + **Half-cell 1:** Copper electrode dipped in 1 M CuSO₄.
  + **Half-cell 2:** Standard Hydrogen Electrode (Pt electrode with H₂ gas at 1 atm in 1 M H⁺ solution).
* Connect with a salt bridge and voltmeter.
* The measured potential = **E°(Cu²⁺/Cu)**.
* For Cu: **+0.34 V** (positive → copper tends to get reduced).

**3. Cell Potential Calculation (Nernst Equation)**

Cell:

Zn(s)∣Zn2+(0.3M)∣∣Cu2+(0.1M)∣Cu(s)\text{Zn(s)}|\text{Zn}^{2+}(0.3M)||\text{Cu}^{2+}(0.1M)|\text{Cu(s)}

* Step 1: **Identify electrodes**
  + Zn → anode (oxidation)
  + Cu → cathode (reduction)
* Step 2: **Standard EMF (E°cell):**

E°cell=E°cathode−E°anodeE°\_\text{cell} = E°\_\text{cathode} - E°\_\text{anode} =(+0.34)−(−0.76)=+1.10 V= (+0.34) - (-0.76) = +1.10 \, \text{V}

* Step 3: **Nernst equation:**

Ecell=E°cell−0.0591nlog⁡QE\_\text{cell} = E°\_\text{cell} - \frac{0.0591}{n} \log Q

where

* n=2n = 2 (electrons transferred)
* Q=[Zn2+][Cu2+]=0.30.1=3Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = \frac{0.3}{0.1} = 3

Ecell=1.10−0.05912log⁡(3)E\_\text{cell} = 1.10 - \frac{0.0591}{2} \log(3) Ecell=1.10−(0.02955)(0.4771)E\_\text{cell} = 1.10 - (0.02955)(0.4771) Ecell≈1.10−0.014=1.086 VE\_\text{cell} \approx 1.10 - 0.014 = 1.086 \, \text{V}

✅ **Final Answer:**

* **Standard Electrode Potential of Cu = +0.34 V**
* **Cell potential at 25 °C = ~1.09 V**

Would you like me to format this into a **super-short memory version** (like 5 bullet points only) for a *last-minute glance* before the exam?

#### In simple term

Alright, let’s make it super simple.

Think of it like this:

* Every metal (like **copper**, **zinc**, **magnesium**) has its own “push or pull” on electrons.
* But we need a **starting point** to compare them.
* Scientists chose **hydrogen electrode** as the **zero point (0.00 V)**. This is called the **Standard Hydrogen Electrode (SHE)**.

👉 So when we say:

* **Copper electrode potential = +0.34 V** → Copper is **better than hydrogen at grabbing electrons** (it gets reduced easily).
* **Zinc electrode potential = –0.76 V** → Zinc is **worse than hydrogen at grabbing electrons** (so it loses electrons instead, i.e., oxidizes).

✨ Imagine SHE as a **ruler**:

* Zero mark = hydrogen.
* If a metal shows a **positive value**, it’s stronger than hydrogen in gaining electrons.
* If a metal shows a **negative value**, it prefers to lose electrons (weaker than hydrogen).

Does this “hydrogen = ruler, metals compared to it” idea make sense to you?