24.2 Standard electrode potentials and the Nernst equation

Name: ____________ Class: ________ Date: __________ **Total: 11 marks**

Objective

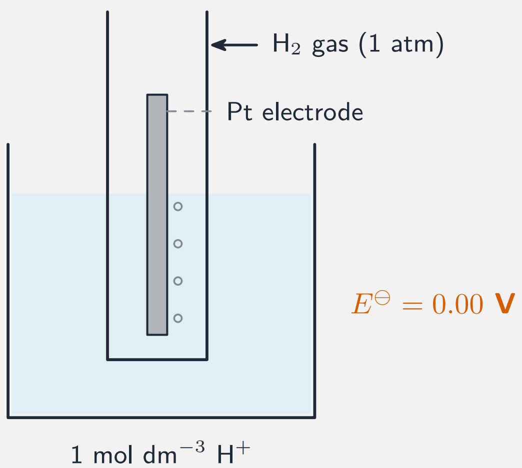
Build the skills to answer exam questions on **standard electrode potentials** 标准电极电势—the **standard hydrogen electrode** 标准氢电极, **cell potentials** 电池电势, **feasibility** 可行性, and the **Nernst equation** 能斯特方程.

You must be able to:

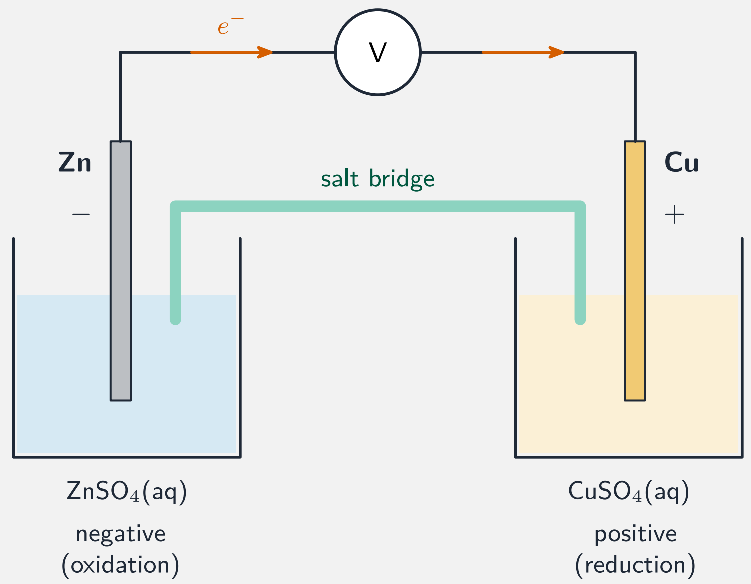
- describe the standard hydrogen electrode and how $E^\ominus$ is measured
- calculate a cell potential and deduce polarity, electron flow and feasibility
- use $E^\ominus$ values to compare oxidising/reducing agents and construct redox equations

1 Worked examples

Electrode and cell potentials



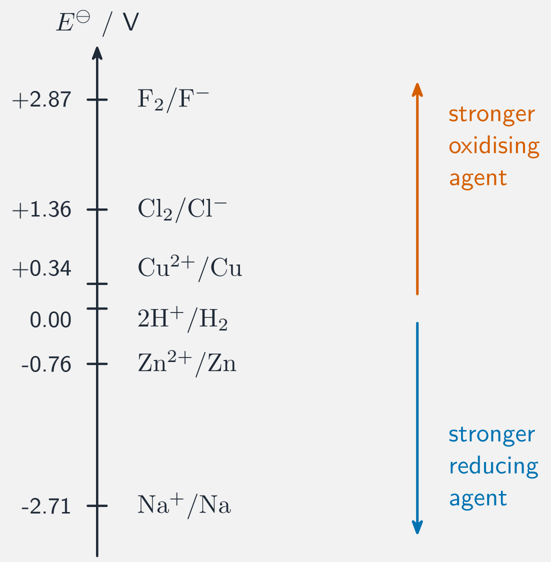
$E^\ominus$ is measured against the standard hydrogen electrode (defined as 0.00 V), under standard conditions, as a reduction



$E^{\ominus}_{cell} = E^{\ominus}$(more positive) − $E^{\ominus}$(less positive). Electrons flow from the more negative electrode

Example. Zn^{2} /Zn (−0.76 V) + Cu^{2} /Cu (+0.34 V): $E^{\ominus}_{\text{cell}} = (+0.34) - (-0.76) =$ +1.10 V.

■ Feasibility and the series



A more positive $E^{\ominus}$ = better oxidising agent; more negative = better reducing agent

A reaction is feasible when $E^{\ominus}_{\text{cell}}$ is **positive**. ($\Delta G^{\ominus} = -nE^{\ominus}_{\text{cell}}F$.) The **Nernst equation** corrects for non-standard concentrations: $E = E^{\ominus} + \dfrac{0.059}{z}\log\dfrac{[\text{ox}]}{[\text{red}]}$.

2 Practice

2.1 Describe the standard hydrogen electrode. [2]

2.2 State the condition on $E^{\ominus}_{\text{cell}}$ for a reaction to be feasible. [1]

3 Exam-style questions

3.1 A more **positive** standard electrode potential indicates a: [1]

- **A** stronger reducing agent
- **B** stronger oxidising agent
- **C** more reactive metal
- **D** higher concentration

3.2 Increasing the concentration of the oxidised species in a half-cell makes its electrode potential: [1]

- **A** more positive
- **B** more negative
- **C** unchanged
- **D** zero

3.3 Use these $E^{\ominus}$ values: $Fe^{2+}/Fe = -0.44$ V; $Ag^{+}/Ag = +0.80$ V.

(a) Calculate the standard cell potential of a cell made from these half-cells. [2]

(b) Deduce which electrode is the negative terminal and the direction of electron flow. [2]

3.4 Using the values above, state whether silver can oxidise iron metal, and explain using $E^{\ominus}_{cell}$. [2]

4 Go further

You are now ready for the real exam questions on this subtopic. Open the **24.2 Standard electrode potentials** past-paper sheet in the Library, or try these in **Practice mode**:

- 9701/41 N25 —Q4 (standard electrode potentials and the nernst equation)
- 9701/41 N25 —Q2 (standard electrode potentials and the nernst equation)
- 9701/42 N25 —Q5 (standard electrode potentials and the nernst equation)

Solutions

2.1 hydrogen gas at 1 atm bubbled over a platinum electrode in 1 mol dm ^{3} H , defined as 0.00 V.

2.2 it must be positive.

3.1 **B**. A more positive $E^{\ominus}$ is a stronger oxidising agent.

3.2 **A**. More oxidised species makes the potential more positive.

3.3 (a) $E^{\ominus}_{\text{cell}} = (+0.80) - (-0.44) = +1.24$ V.
(b) the Fe^{2} /Fe electrode (more negative) is the negative terminal; electrons flow from the iron to the silver electrode in the external circuit.

3.4 yes; $E^{\ominus}_{\text{cell}}$ for Ag oxidising Fe is positive (+1.24 V), so the reaction is feasible.