

24.2 Standard electrode potentials and the Nernst equation

Name: _____ Class: _____ Date: _____

Total: 11 marks

KEY WORDS nernst equation 能斯特方程 • electrode potential 电极电势 • standard electrode potential 标准电极电势 • standard cell potential 标准电池电势 • standard hydrogen electrode 标准氢电极

Objective

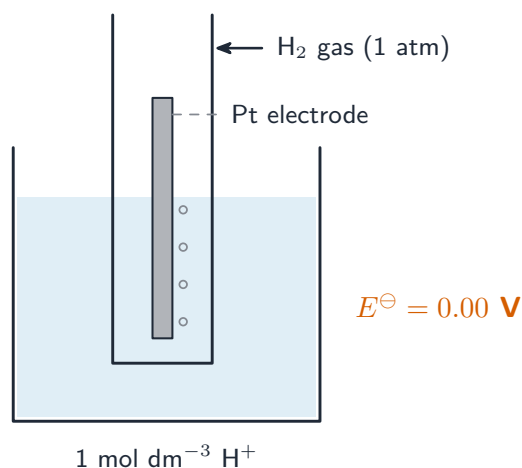
Build the skills to answer exam questions on **standard electrode potentials** 标准电极电势 — the **standard hydrogen electrode** 标准氢电极, **cell potentials** 电池电势, **feasibility** 可行性, and the **Nernst equation** 能斯特方程.

You must be able to:

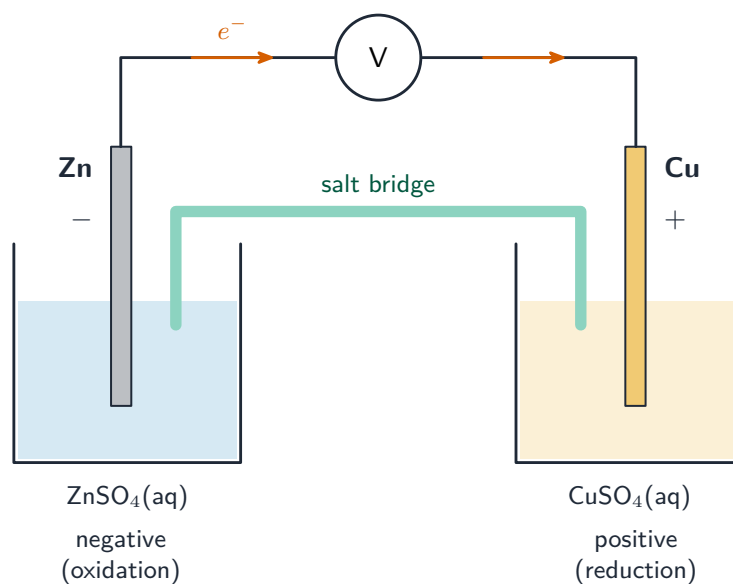
- describe the standard hydrogen electrode and how $E^\ominus$ is measured
- calculate a cell potential and deduce polarity, electron flow and feasibility
- use $E^\ominus$ values to compare oxidising/reducing agents and construct redox equations

1 Worked examples

■ Electrode and cell potentials



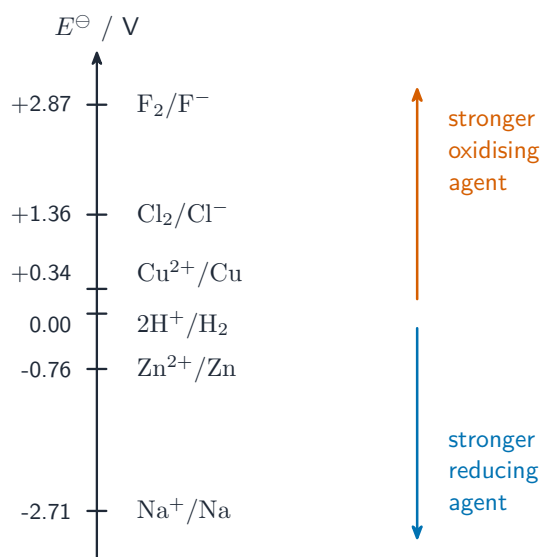
$E^\ominus$ is measured against the standard hydrogen electrode (defined as 0.00 V), under standard conditions, as a reduction



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

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

■ Feasibility and the series



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

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

2 Practice

2.1 Describe the standard hydrogen electrode. [2]

2.2 State the condition on $E_{\text{cell}}^{\ominus}$ 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: $\text{Fe}^{2+}/\text{Fe} = -0.44 \text{ V}$; $\text{Ag}^{+}/\text{Ag} = +0.80 \text{ V}$.

(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 $E_{\text{cell}}^{\ominus}$. [2]

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) $E_{\text{cell}}^\ominus = (+0.80) - (-0.44) = +1.24$ V.

(b) the Fe²⁺/Fe electrode (more negative) is the negative terminal; electrons flow from the iron to the silver electrode in the external circuit.

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