# Exercise walk-through

## Standard electrode potentials E , standard cell potentials E cell and the ▪ 24.2

eXercise

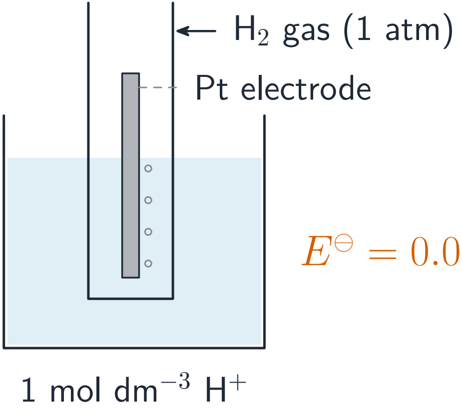
# You must be able to

- 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

# Method — Electrode and cell potentials

*$E^\ominus$ is measured against the standard hydrogen electrode (defined as 0.00 V), under standard conditions, as a reduction*
*$E^\ominus_{cell} = E^\ominus(\text{more positive}) - E^\ominus(\text{less positive})$. Electrons flow from the more negative electrode*
**Example.** $Zn^2$ /Zn (−0.76 V) + $Cu^2$ /Cu (+0.34 V):
$E^\ominus_{\text{cell}} = (+0.34) - (-0.76) = +1.10$ V.

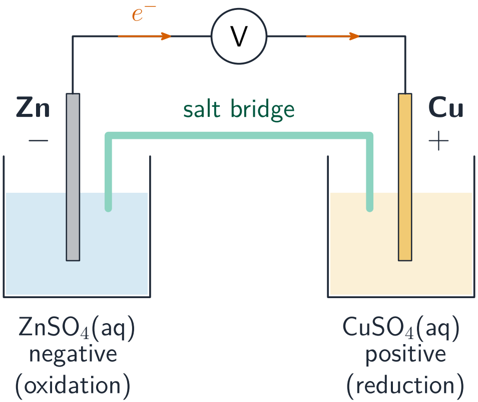
# Method — Electrode and cell potentials




$E^\ominus = 0.00$ **V**

*The standard hydrogen electrode: H at 1 atm over Pt in 1 mol/dm$^3$ H , defined as 0.00 V*

# Method — Electrode and cell potentials

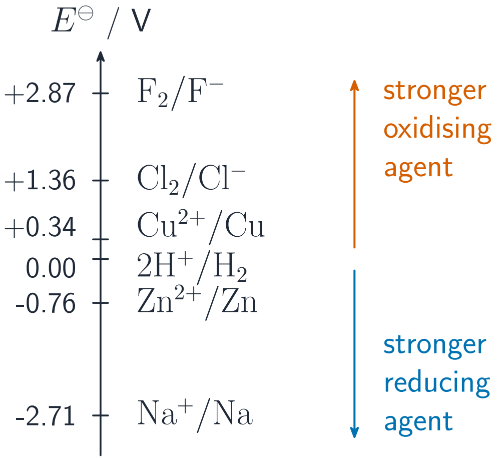



*A Zn and a Cu half-cell joined by a salt bridge and voltmeter; electrons flow Zn → Cu*

# Method — Feasibility and the series

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: $E = E^{\ominus} + \dfrac{0.059}{z} \log \dfrac{[\text{ox}]}{[\text{red}]}$.

# Method — Feasibility and the series



# Practice 2.1 ▪ 2 marks

Describe the standard hydrogen electrode.

# Solution 2.1

Method: Electrode and cell potentials

## Worked steps

hydrogen gas at 1 atm bubbled over a platinum electrode **B1** in 1 mol dm $^{3}$ H , defined as 0.00 V **B1**.

# Practice 2.2 ▪ 1 mark

State the condition on $E^\ominus_{\mathrm{cell}}$ for a reaction to be feasible.

# Solution 2.2

Method: Feasibility and the series

**Worked steps**

it must be positive **B1**.

# Exam-style 3.1 ▪ 1 mark

A more **positive** standard electrode potential indicates a:

**A** stronger reducing agent

**B** stronger oxidising agent

**C** more reactive metal

**D** higher concentration

# Solution 3.1

Method: Feasibility and the series

**A** stronger reducing agent

**B** stronger oxidising agent ✓

**C** more reactive metal

**D** higher concentration

## Worked steps

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

# Exam-style 3.2 ▪ 1 mark

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

**A** more positive

**B** more negative

**C** unchanged

**D** zero

# Solution 3.2

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

**Worked steps**

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

# Exam-style 3.3 ▪ 2 marks

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

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

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

# Solution 3.3

## Worked steps

(a) $E^{\ominus}_{\text{cell}} = (+0.80) - (-0.44) = +1.24$ V **M1** **A1**.

(b) the $Fe^{2}$ /Fe electrode (more negative) is the negative terminal **M1**; electrons flow from the iron to the silver electrode in the external circuit **A1**.

# Exam-style 3.4 ▪ 2 marks

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

# Solution 3.4

Method: Feasibility and the series

**Worked steps**

yes **M1**; $E^{\ominus}_{\text{cell}}$ for Ag oxidising Fe is positive (+1.24 V), so the reaction is feasible **A1**.