

# IB Diploma Programme Chemistry

## Reactivity 3: What are the Mechanisms of Chemical Change?

### Subtopic R3.2

R3.2 - Electron Transfer Reactions (Redox and Electrochemistry)

Comprehensive Revision Notes

Both SL and AHL Content

**MINDLAB**

January 21, 2026

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## Contents

<b>1</b>	<b>R3.2 - Electron Transfer Reactions (Redox and Electrochemistry)</b>	<b>2</b>
1.1	Overview . . . . .	2
1.2	Core Definitions . . . . .	2
1.3	Oxidation States (SL) . . . . .	2
1.4	Half-Equations (SL) . . . . .	2
1.5	Voltaic (Galvanic) Cells (SL) . . . . .	3
1.6	Electrolytic Cells (SL) . . . . .	4
1.7	Organic Redox Reactions (SL) . . . . .	4
1.8	AHL - Standard Electrode Potentials . . . . .	5
1.9	AHL - Standard Cell Potential . . . . .	5
1.10	AHL - Gibbs Energy and Electrochemistry . . . . .	6
1.11	AHL - Electrolysis of Aqueous Solutions . . . . .	6
1.12	AHL - Electroplating . . . . .	7
1.13	Topic Linkages . . . . .	7
<b>2</b>	<b>Diagram Practice</b>	<b>9</b>

## 1 R3.2 - Electron Transfer Reactions (Redox and Electrochemistry)

### 1.1 Overview

**Guiding Question:** *What happens when electrons are transferred?*

**Teaching Hours:** SL: 10 hours — HL: 15 hours (10 SL + 5 HL)

### 1.2 Core Definitions

#### SL Core Content

##### Oxidation and Reduction (OIL RIG):

- **Oxidation:** Loss of electrons, increase in oxidation state
- **Reduction:** Gain of electrons, decrease in oxidation state
- **Oxidizing agent:** Causes oxidation (itself gets reduced)
- **Reducing agent:** Causes reduction (itself gets oxidized)

**Mnemonic:** OIL RIG (Oxidation Is Loss, Reduction Is Gain)

### 1.3 Oxidation States (SL)

#### Rules for Assigning Oxidation States:

1. Elements in pure form: 0 (e.g., O<sub>2</sub>, Fe)
2. Monatomic ions: charge on ion (e.g., Na<sup>+</sup> = +1, Cl<sup>-</sup> = -1)
3. Oxygen: usually -2 (except peroxides -1, OF<sub>2</sub> +2)
4. Hydrogen: usually +1 (except metal hydrides -1)
5. Fluorine: always -1
6. Group 1 metals: always +1
7. Group 2 metals: always +2
8. Sum of oxidation states = overall charge

**Example:** Determine oxidation state of S in H<sub>2</sub>SO<sub>4</sub>.

*Solution:* Let S =  $x$

$$2(+1) + x + 4(-2) = 0$$

$$2 + x - 8 = 0$$

$$x = +6$$

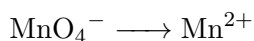
### 1.4 Half-Equations (SL)

#### Balancing Half-Equations in Acidic/Neutral Solution:

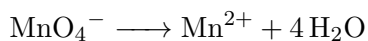
1. Balance atoms other than O and H
2. Balance O by adding H<sub>2</sub>O
3. Balance H by adding H<sup>+</sup>
4. Balance charge by adding e<sup>-</sup>

**Example:**  $\text{MnO}_4^- \longrightarrow \text{Mn}^{2+}$  (acidic)

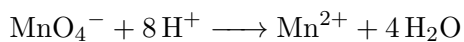
Step 1: Mn already balanced



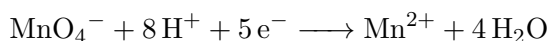
Step 2: Balance O with  $\text{H}_2\text{O}$



Step 3: Balance H with  $\text{H}^+$



Step 4: Balance charge with  $\text{e}^-$



## 1.5 Voltaic (Galvanic) Cells (SL)

### SL Core Content

#### Components:

- **Anode:** Oxidation occurs, negative terminal (-), electron source
- **Cathode:** Reduction occurs, positive terminal (+), electron sink
- **Salt bridge:** Allows ion flow to maintain neutrality (NOT electron flow!)
- **External circuit:** Wire connecting electrodes (electrons flow anode  $\rightarrow$  cathode)

**Mnemonic:** “An Ox” (Anode = Oxidation), “Red Cat” (Reduction = Cathode)

#### Example - Zn/Cu Cell:

- Anode:  $\text{Zn (s)} \longrightarrow \text{Zn}^{2+} \text{ (aq)} + 2 \text{e}^-$  (oxidation)
- Cathode:  $\text{Cu}^{2+} \text{ (aq)} + 2 \text{e}^- \longrightarrow \text{Cu (s)}$  (reduction)
- Overall:  $\text{Zn (s)} + \text{Cu}^{2+} \text{ (aq)} \longrightarrow \text{Zn}^{2+} \text{ (aq)} + \text{Cu (s)}$

### Common Pitfall

**Common Error:** Electrons do NOT flow through salt bridge!

- Electrons: Flow through external wire only (anode  $\rightarrow$  cathode)
- Ions: Flow through salt bridge (anions toward anode, cations toward cathode)

## 1.6 Electrolytic Cells (SL)

### SL Core Content

#### Key Differences from Voltaic Cells:

- Requires external DC power source
- Drives non-spontaneous reactions
- **Anode:** Oxidation, POSITIVE terminal (connected to + of battery)
- **Cathode:** Reduction, NEGATIVE terminal (connected to - of battery)
- Polarity REVERSED compared to voltaic cell!

#### Electrolysis of Molten NaCl:

- Cathode:  $\text{Na}^+ + \text{e}^- \longrightarrow \text{Na}$  (reduction)
- Anode:  $2\text{Cl}^- \longrightarrow \text{Cl}_2 + 2\text{e}^-$  (oxidation)
- Overall:  $2\text{NaCl (l)} \longrightarrow 2\text{Na (l)} + \text{Cl}_2 \text{ (g)}$

## 1.7 Organic Redox Reactions (SL)

### SL Core Content

#### Oxidation of Alcohols:

- **Primary ( $1^\circ$ ):**  $\text{RCH}_2\text{OH} \xrightarrow{\text{oxidize}} \text{RCHO} \xrightarrow{\text{oxidize}} \text{RCOOH}$ 
  - Use distillation to collect aldehyde (prevents further oxidation)
  - Use reflux to form carboxylic acid (prolonged heating)
- **Secondary ( $2^\circ$ ):**  $\text{R}_2\text{CHOH} \xrightarrow{\text{oxidize}} \text{R}_2\text{CO}$  (ketone, stops here)
- **Tertiary ( $3^\circ$ ):** No oxidation under normal conditions

#### Reduction of Carbonyl Compounds:

- $\text{RCHO} \xrightarrow{\text{reduce}} \text{RCH}_2\text{OH}$  (aldehyde  $\rightarrow$  primary alcohol)
- $\text{R}_2\text{CO} \xrightarrow{\text{reduce}} \text{R}_2\text{CHOH}$  (ketone  $\rightarrow$  secondary alcohol)
- $\text{RCOOH} \xrightarrow{\text{reduce}} \text{RCHO} \xrightarrow{\text{reduce}} \text{RCH}_2\text{OH}$  (acid  $\rightarrow$  aldehyde  $\rightarrow$  alcohol)

#### Reduction of Alkenes/Alkynes:

- $\text{C}=\text{C} + \text{H}_2 \longrightarrow \text{C}-\text{C}$  (alkene  $\rightarrow$  alkane)
- $\text{C}\equiv\text{C} + \text{H}_2 \longrightarrow \text{C}=\text{C} + 2\text{H}_2 \longrightarrow \text{C}-\text{C}$  (alkyne  $\rightarrow$  alkene  $\rightarrow$  alkane)

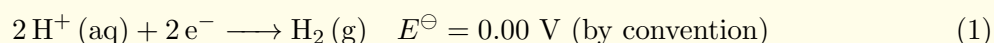
## 1.8 AHL - Standard Electrode Potentials

## AHL Additional Content

Standard Electrode Potential ( $E^\ominus$ ):

- Voltage of half-cell relative to standard hydrogen electrode (SHE)
- Standard conditions: 298 K, 1 bar, 1 mol dm<sup>-3</sup>
- All data booklet values written as REDUCTION potentials
- More positive  $E^\ominus$  = stronger oxidizing agent (easier to reduce)
- More negative  $E^\ominus$  = stronger reducing agent (easier to oxidize)

## Standard Hydrogen Electrode (SHE):



## 1.9 AHL - Standard Cell Potential

## AHL Additional Content

## Key Formula

## Standard Cell Potential:

$$E_{\text{cell}}^\ominus = E_{\text{cathode}}^\ominus - E_{\text{anode}}^\ominus \quad (2)$$

## Spontaneity Criterion:

- If  $E_{\text{cell}}^\ominus > 0$ : Spontaneous (forward reaction favored)
- If  $E_{\text{cell}}^\ominus < 0$ : Non-spontaneous (reverse reaction favored)
- If  $E_{\text{cell}}^\ominus = 0$ : At equilibrium

**Example (HL):** Calculate  $E_{\text{cell}}^\ominus$  for  $\text{Zn}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \longrightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$

Given:  $E^\ominus(\text{Cu}^{2+}/\text{Cu}) = +0.34 \text{ V}$ ,  $E^\ominus(\text{Zn}^{2+}/\text{Zn}) = -0.76 \text{ V}$

*Solution:*

- More positive  $E^\ominus$  = cathode:  $\text{Cu}^{2+}/\text{Cu} = +0.34 \text{ V}$
- More negative  $E^\ominus$  = anode:  $\text{Zn}^{2+}/\text{Zn} = -0.76 \text{ V}$

$$\begin{aligned} E_{\text{cell}}^\ominus &= E_{\text{cathode}}^\ominus - E_{\text{anode}}^\ominus \\ &= (+0.34) - (-0.76) \\ &= +1.10 \text{ V} \end{aligned}$$

Since  $E_{\text{cell}}^\ominus > 0$ , reaction is spontaneous.

## 1.10 AHL - Gibbs Energy and Electrochemistry

## AHL Additional Content

## Key Formula

Relationship Between  $\Delta G^\ominus$  and  $E_{\text{cell}}^\ominus$ :

$$\Delta G^\ominus = -nFE_{\text{cell}}^\ominus \quad (3)$$

Where:

- $n$  = moles of electrons transferred
- $F$  = Faraday constant =  $96,500 \text{ C mol}^{-1}$
- $E_{\text{cell}}^\ominus$  in volts (V)
- $\Delta G^\ominus$  in joules (J) [convert to kJ: divide by 1000]

**Example (HL):** Calculate  $\Delta G^\ominus$  for the Zn/Cu cell ( $E_{\text{cell}}^\ominus = +1.10 \text{ V}$ ,  $n = 2$ ).*Solution:*

$$\begin{aligned} \Delta G^\ominus &= -nFE_{\text{cell}}^\ominus \\ &= -2 \times 96500 \times 1.10 \\ &= -212300 \text{ J} \\ &= -212 \text{ kJ mol}^{-1} \end{aligned}$$

Negative  $\Delta G^\ominus$  confirms spontaneous reaction.

## 1.11 AHL - Electrolysis of Aqueous Solutions

## AHL Additional Content

## Competing Reactions:

At cathode (reduction):

- Metal ion reduction:  $\text{M}^{n+} + n\text{e}^- \longrightarrow \text{M}$
- Water reduction:  $2\text{H}_2\text{O} + 2\text{e}^- \longrightarrow \text{H}_2 + 2\text{OH}^-$  ( $E^\ominus = -0.83 \text{ V}$ )

At anode (oxidation):

- Halide oxidation:  $2\text{X}^- \longrightarrow \text{X}_2 + 2\text{e}^-$
- Water oxidation:  $2\text{H}_2\text{O} \longrightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$  ( $E^\ominus = +1.23 \text{ V}$ )

**Prediction:** Species with more positive  $E^\ominus$  is preferentially reduced (cathode); species with more negative  $E^\ominus$  is preferentially oxidized (anode)**Example - Electrolysis of  $\text{CuSO}_4(\text{aq})$ :**

- Cathode:  $\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$  ( $E^\ominus = +0.34 \text{ V}$ ) preferred over water reduction
- Anode:  $2\text{H}_2\text{O} \longrightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$  ( $\text{SO}_4^{2-}$  cannot be oxidized)

## 1.12 AHL - Electroplating

### AHL Additional Content

#### Electroplating Setup:

- Object to be plated = cathode (reduction deposits metal)
- Metal source = anode (oxidation supplies metal ions)
- Electrolyte contains metal ions

#### Example - Copper Plating:

- Cathode (object):  $\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$  (metal deposited)
- Anode (Cu electrode):  $\text{Cu} \longrightarrow \text{Cu}^{2+} + 2\text{e}^-$  (replenishes  $\text{Cu}^{2+}$ )
- Electrolyte:  $\text{CuSO}_4(\text{aq})$

## 1.13 Topic Linkages

- **R1.4 (Gibbs Energy):**  $\Delta G^\ominus = -nFE_{\text{cell}}^\ominus$  connects electrochemistry to thermodynamics (HL)
- **S3.1 (Periodic Trends):** Reactivity series follows ionization energy trends
- **R2.3 (Equilibrium):** Secondary cells involve reversible reactions



## Key Points to Remember

### Key Tips

- Activity series follows periodic trends:: Group 1 and 2 metals: Very reactive (lose electrons easily)
- Transformed Structure::  $\text{CH}_2\text{OH} \rightarrow \text{-CHO} \rightarrow \text{-COOH}$ , carbon oxidation state  $-1 \rightarrow +1 \rightarrow +3$
- Sum of oxidation states = overall charge: Core principle for all calculations

## 2 Diagram Practice

*Use these spaces to practice drawing essential diagrams for this topic.*

▷ Draw: Oxidation state determination flowchart

▷ Draw: Half-equation balancing steps

▷ Draw: Oxidation and reduction memory aid (OIL RIG)

▷ Draw: Electron transfer in a redox reaction