Topic 19 – Simple Cells

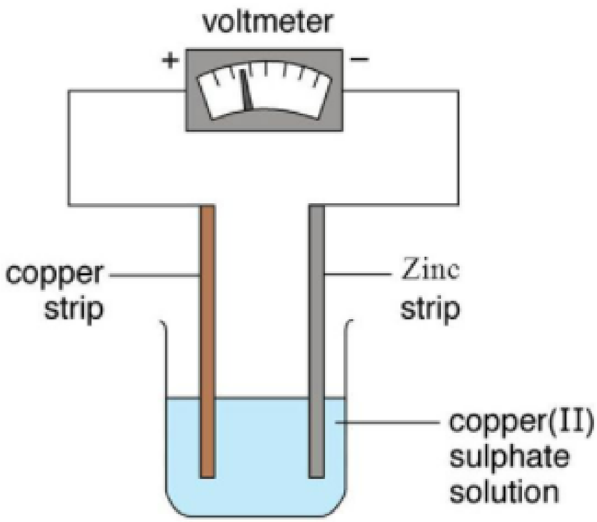
Subject content:

8.5 Simple Electric Cell (a) Describe the production of electrical energy from simple cells (i.e. two electrodes in an electrolyte) linked to the Reactivity Series and redox reactions (in terms of electron transfer)

19.1 Introduction

Simple cell

Device which converts chemical → electrical energy by placing 2 different metals in contact with an electrolyte



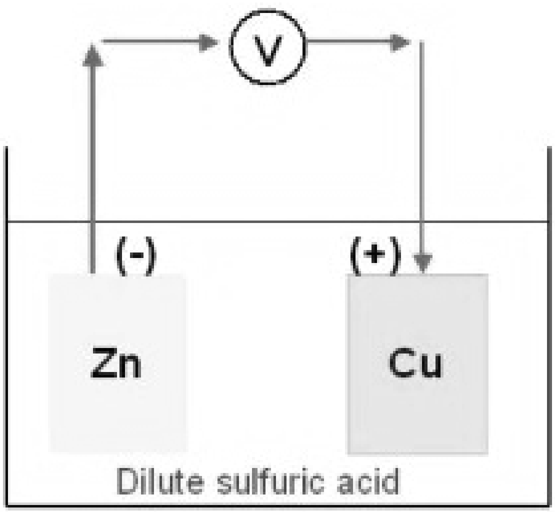
Structure

1. **Electrolyte**	2. **Electrode**
Ionic compound (molten / aqueous) conducts electric current + decomposed in the process	**Electric conductor (metal)** where current enters / leaves electrolyte → connected to external circuit with wires ● **Anode** → oxidation (–ve) ● **Cathode** → reduction (+ve)

Electrodes

Anode	**Cathode**
negative electrode (–)	**positive** electrode (+)
oxidation	**reduction**

19.2 Functioning of Simple Cell



Zinc is more reactive than copper → <u>loses e⁻</u> more readily

- Negative electrode: more reactive metal (zinc)
- Positive electrode: less reactive metal (copper)

Stage	Explanation	Equation
Zinc electrode	oxidation → negative terminal (**<u>anode</u>**)	$Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2\ e^-$
Electron flow	• e^- leave zinc electrode, pass through voltmeter, enter copper electrode • e^- attract Cu^{2+} ions in solution	
Copper electrode	reduction → positive terminal (**<u>cathode</u>**)	$Cu^{2+}\ (aq) + 2\ e^- \rightarrow Cu\ (s)$
Overall	e^- flow in external circuit → **<u>electrical current</u>** 1. Zinc electrode becomes smaller, copper electrode becomes bigger 2. Blue electrolyte fades → colourless ($CuSO_4 \rightarrow ZnSO_4$)	$\mathbf{Zn\ (s) + Cu^{2+}\ (aq) \rightarrow Zn^{2+}\ (aq) + Cu\ (s)}$

19.3 Voltage of Simple Cell

Voltage / potential difference:

The further apart the two metals (electrodes) are in reactivity series, the greater the voltage produced in the cell

*both electrodes made of same metal → no current flow

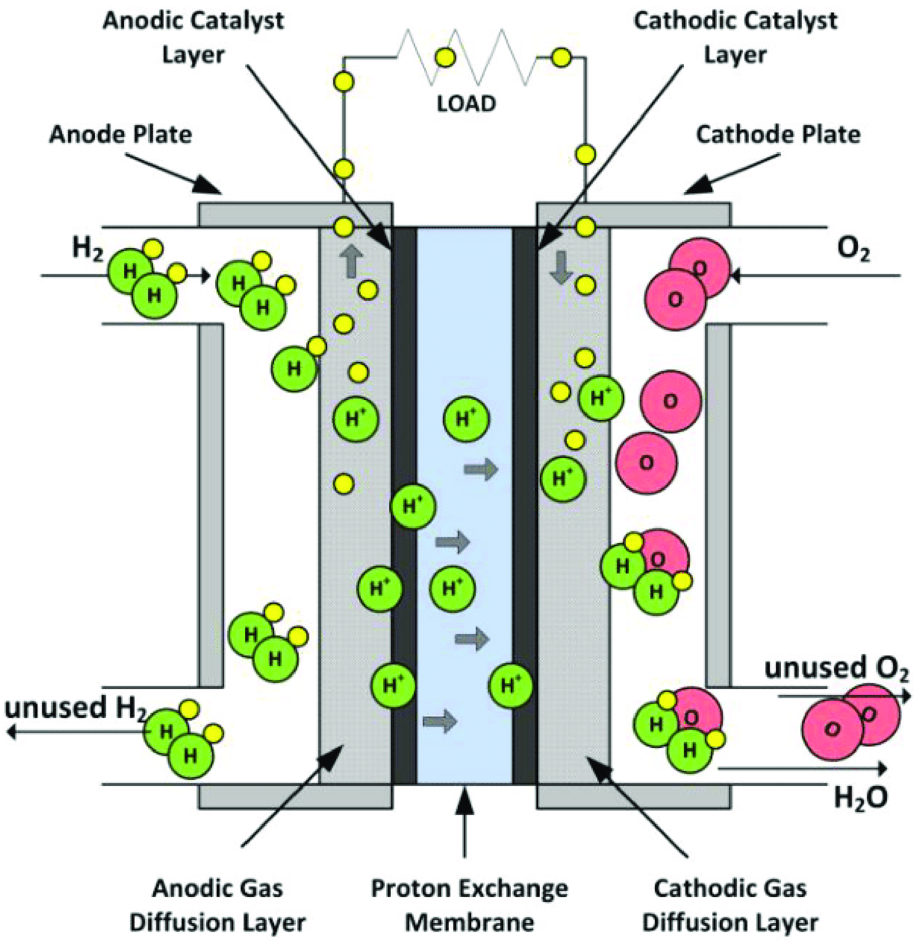
Example:

Metal electrodes		Voltage / V
magnesium	copper	2.7
zinc	copper	1.1
iron	copper	0.8
lead	copper	0.5
copper	copper	0.0

19.4 Hydrogen Fuel Cell

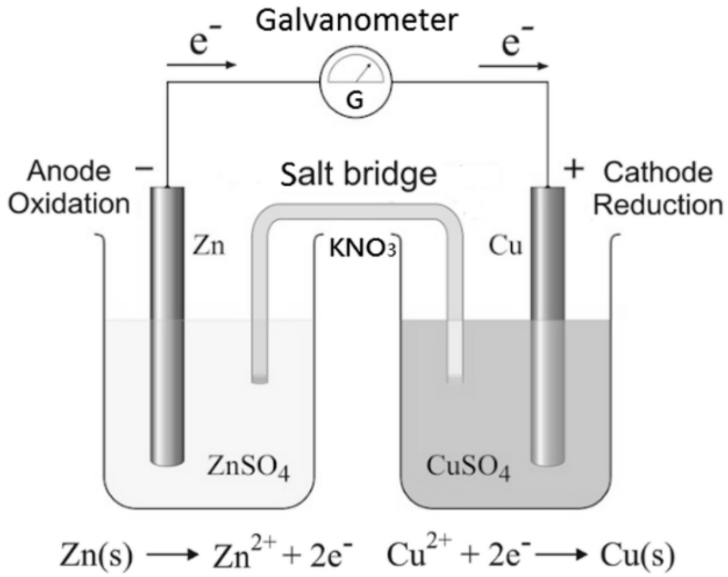
Fuel cells

Simple cells where reactant (fuel + oxygen) are continuously supplied to produce electricity



Negative electrode (anode)	Positive electrode (cathode)
Hydrogen: oxidised → hydrogen ions	Oxygen: reduced → water
$2\ H_2\ (g) \rightarrow 4\ H^+\ (aq) + 4\ e^-$	$O_2\ (g) + 4\ H^+\ (aq) + 4\ e^- \rightarrow 2\ H_2O\ (l)$
$2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (l)$	

19.5 Daniell Cell



Zinc metal dipped in zinc sulfate solution	Copper metal dipped in copper(II) sulfate solution
<u>Zinc electrode</u> ● Reactivity: zinc > copper ● Lose electrons more readily	<u>Copper electrode</u> ● Reactivity: copper < zinc ● Gain electrons from zinc (through external wire connection)
Electron flow: Zn → Cu	
Working electrical circuit 1. movement of ions (through salt bridge) 2. movement of electrons (connecting wires)	

Salt bridge
- Contains salt / inert ions that <u>do not react with electrolytes</u>
  - sodium nitrate
  - dilute sulfuric acid
- Complete circuit → allow flow of ions through & prevent mixing of 2 electrolytes

Typical questions

Multiple choice questions

1 An example of a non-electrolyte is

A $C_6H_{12}O_6$ (aq)
B NaC*l* (aq)
C HNO_3 (aq)
D $CuSO_4$ (aq)

2 The following shows the lead-acid battery reaction:

$$Pb\ (s) + PbO_2\ (s) + 2\ H_2SO_4\ (aq) \rightarrow 2\ PbSO_4\ (s) + H_2O\ (l)$$

When the lead-acid battery is used, the sulfuric acid is a

A reactant, with decreasing concentration
B reactant, with increasing concentration
C product, with decreasing concentration
D product, with increasing concentration

3 A simple cell setup consists of the electrodes, connecting wires and a salt bridge. When a reaction occurs, electron migration takes place through the

A anode
B salt bridge
C cathode
D connecting wires

4 The following shows the lead-acid battery reaction:

$$Pb\ (s) + PbO_2\ (s) + 2\ H_2SO_4\ (aq) \rightarrow 2\ PbSO_4\ (s) + H_2O\ (l)$$

Which half-equation represents the oxidation that occurs?

A $Pb \rightarrow Pb^{2+} + 2\ e^-$
B $Pb^{2+} + 2\ e^- \rightarrow Pb$
C $Pb^{4+} + 4\ e^- \rightarrow Pb$
D $Pb \rightarrow Pb^{4+} + 4\ e^-$

Structured questions

1 The diagram represents a simple cell made by dipping rods of silver and aluminium into dilute hydrochloric acid.

(a) Explain how this simple cell can produce an electrical current. Include equations at each electrode.

- Aluminium is more reactive than silver, so it loses electrons more readily than silver to form Al^{3+} ions. At aluminium electrode, half equation: $Al\ (s) \rightarrow Al^{3+}\ (aq) + 3\ e^{-}$. The aluminium electrode becomes the negative electrode.
- The electrons from aluminium move to the silver electrode (positive electrode).
- The electrons attract H^{+} ions in the solution (H^{+} ions are preferentially discharged, as Al is more reactive than H, so Al has a greater tendency to remain as cations). H^{+} ions remove electrons from the silver electrode to form hydrogen gas. At the silver electrode, half equation: $2\ H^{+}\ (aq) + 2\ e^{-} \rightarrow H_2\ (g)$.
- The flow of electrons from the aluminium to silver electrode produces an electric current.

(b) State and explain what change to the bulb you would see if the aluminium rod was replaced by a rod of

(i) magnesium;

Brighter.

Magnesium is a more reactive metal than aluminium. The difference in the reactivity between magnesium and silver is greater than that between aluminium and silver. Hence, a greater voltage is produced and the bulb will be brighter.

(ii) silver.

Does not light up.

The electrodes are of the same reactivity, thus the voltage / potential difference of the simple cell is zero. Hence, there is no current flow and the bulb does not light up.