9.3 Electrolytic Cell

At the end of the lesson,the students should be able to:

f) Define Faraday's Law of electrolysis

g) Apply Faraday's Law in calculation

h) State the application of electrolysis in industry

Faraday's Law of Electrolysis

Faraday's First Law

States that the **mass, m** of substance formed at an electrode is **directly proportional** to the quantity of **electricity, Q** (in coulombs, **C**) transferred at the electrode.

$$m \propto Q$$

$$Q = I\ t$$

where I = current in ampere, A

t = time in second, s

Describes the relationship between the amount of electricity passed through an electrolytic cell and the amount of substances produced at electrode.

One Faraday (Faraday Constant / Faraday Number) is the electric charge carried by **one mole of electrons**.

1 *F*

= Avogadro constant, *L* x the charge on an electron

= $(6.02205 \times 10^{23}\ \text{mol}^{-1})(1.60219 \times 10^{-19}\ \text{C})$

≈ $96500\ \text{C mol}^{-1}$

1F ≡ 1 mole electrons

1F = **96 500** C mol^{-1}

Consider the half-reaction below:

Cell 1	Cell 2	Cell 3
Cathode:	Cathode:	Cathode:
$Ag^{+} + e^{-} \rightarrow Ag$	$Al^{3+} + 3e^{-} \rightarrow Al$	$Zn^{2+} + 2e^{-} \rightarrow Zn$

Example 12:

Calculate the mass of platinum deposited on a ring when a current of 0.25 A is passed through the electrolyte for 90.0 s.

Solution:

Example 13:

What volume (in litres at *rtp*) of oxygen gas will be collected at the anode when a current of 1.50 A is passed through a solution of Na_2SO_4 for 5.00 minutes.

Solution:

Example 14:

When a current of 2.50 A is passed through a solution containing cation Cr^{n+} for 50.0 minutes, it is found that 1.35 g of chromium is deposited. Use this information to calculate *n*.

Solution:

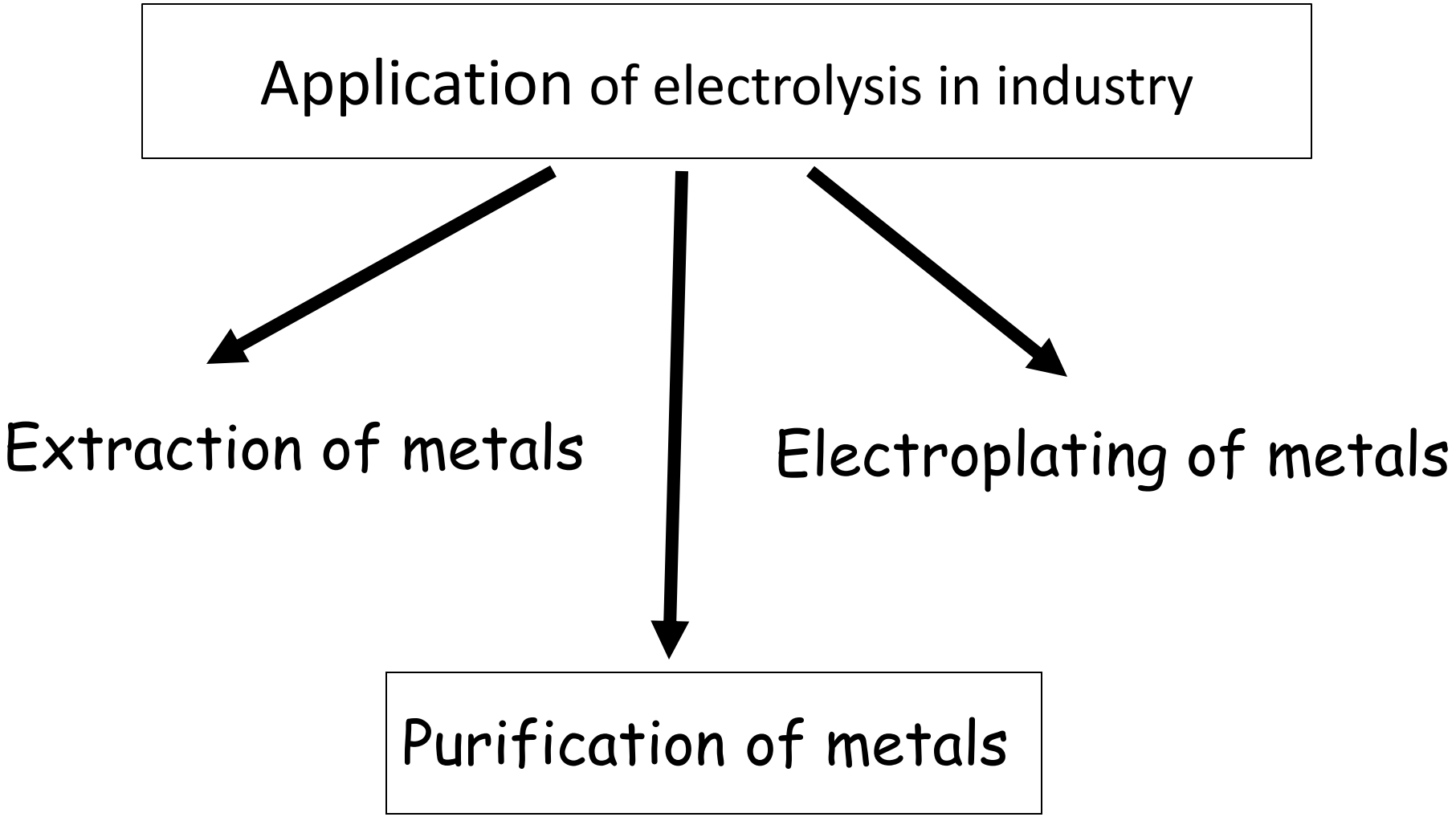
Faraday's Second Law

States that the amount of different substances produced by the same amount of electric charge is Inversely proportional to the charge on the ions.

Consider the half-reaction below.

1F of electricity will discharge ;

Cell 1	Cell 2	Cell 3
Cathode:	Cathode:	Cathode:
$Ag^{+} + e^{-} \rightarrow Ag$	$Al^{3+} + 3e^{-} \rightarrow Al$	$Zn^{2+} + 2e^{-} \rightarrow Zn$


Application of electrolysis in industry
Extraction of metals
Purification of metals
Electroplating of metals

Extraction of metals

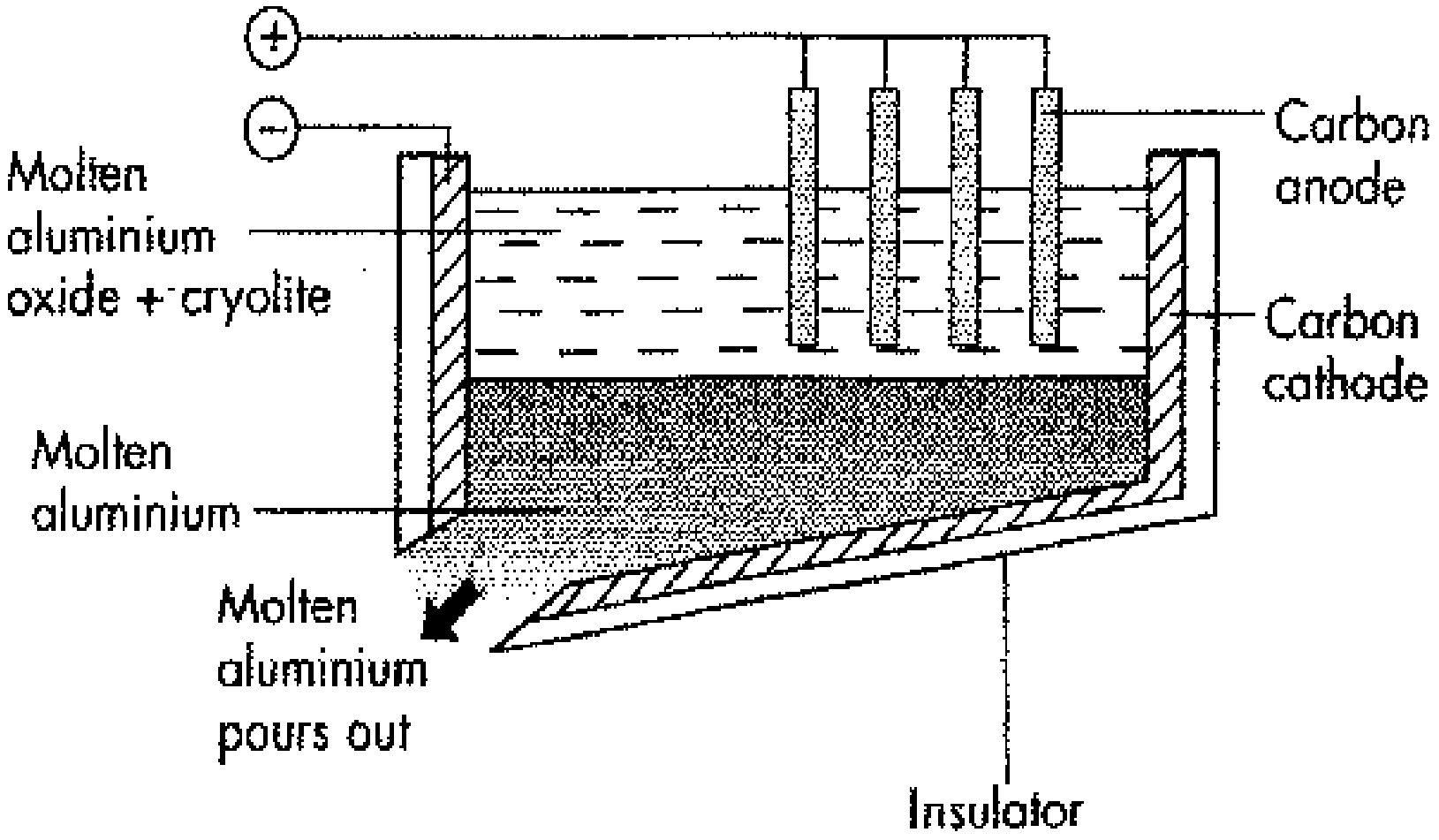


Figure 1: *Extraction of aluminium by electrolysis of bauxite*

- Electrolysis molten Al_2O_3
- Also call Hall process
- Dissolving alumina (Al_2O_3)in cryolite (Na_3AlF_6)
- Solution contains Al^{3+} and O^{2-} from alumina and Na^+ and hexafluoroaluminate , AlF_6^{3-} from cryolite
- Al^{3+} and Na^+ move to cathode
- Al^{3+} selectively reduced form Al atom
- Because : the standard reduction potential of Al is more positive (-1.66 V) than sodium (-2.71 V)

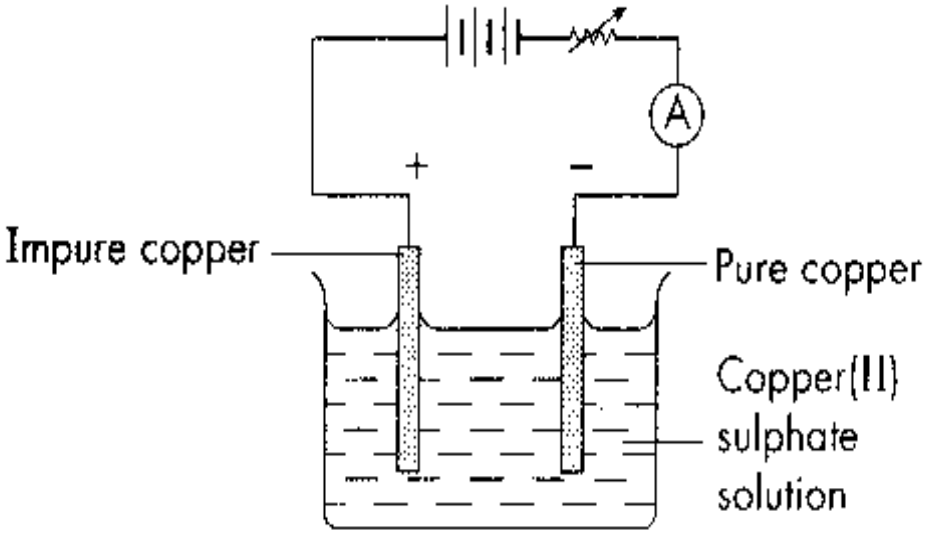
Reactions involved:

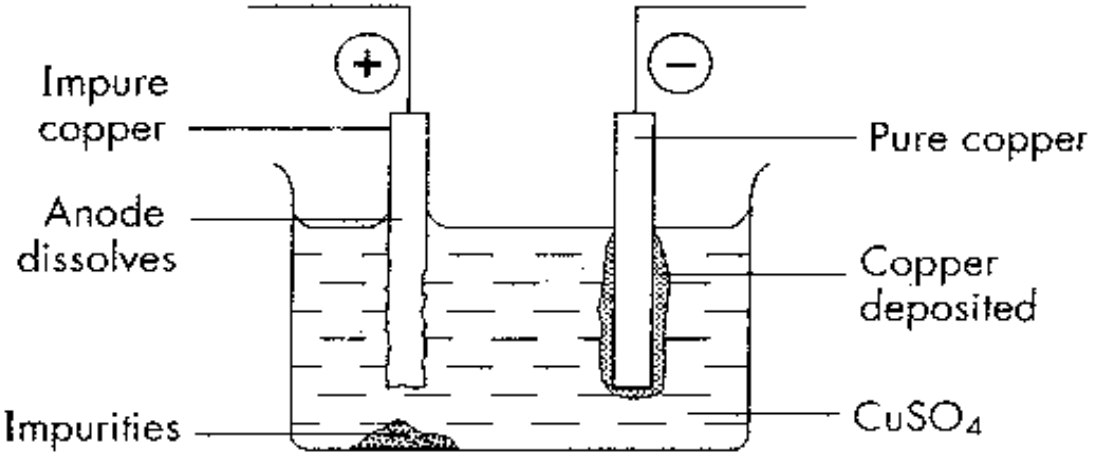
Cathode: $4\ Al^{3+}\ (l) + 12e^- \rightarrow 4Al\ (l)$

Anode: $6O^{2-}\ (l) \rightarrow 3O_2\ (g) + 12e^-$

Cell reaction: $4\ Al^{3+}\ (l) + 6O^{2-}\ (l) \rightarrow 4Al\ (l) + 3O_2\ (g)$

Purification of metals





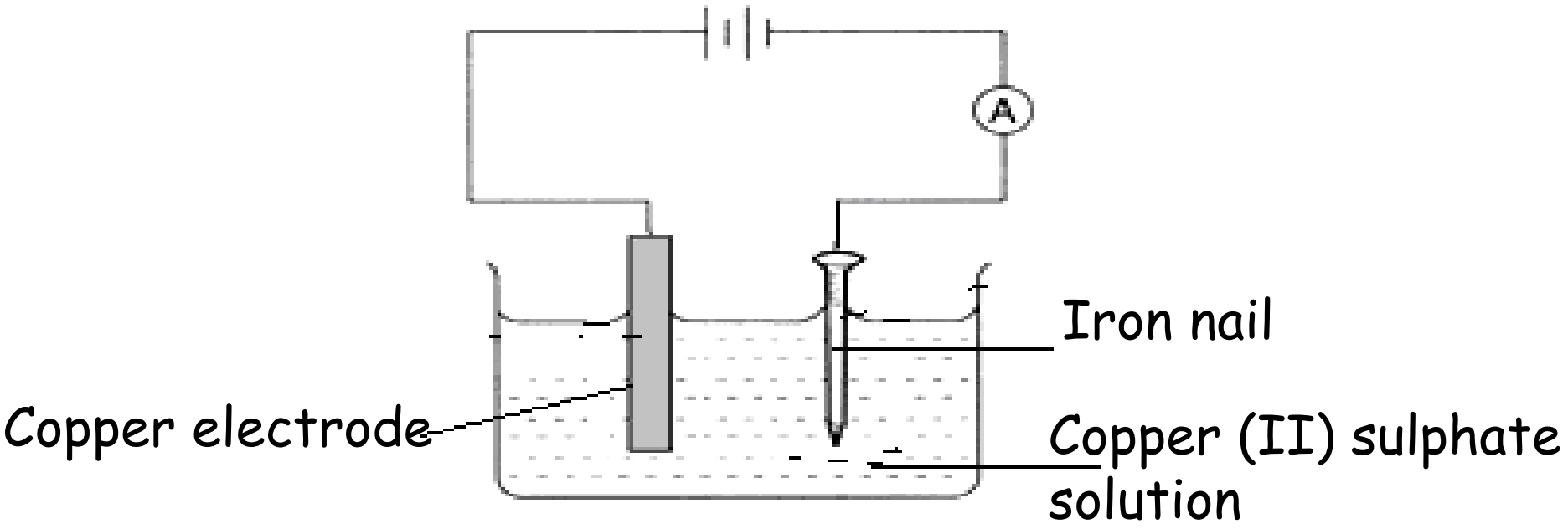
Reactions involved:

Anode: $Cu\ (s) \rightarrow Cu^{2+}\ (aq) + 2e^-$

Cathode: $Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ \ (s)$

Overall equation: $Cu\ (s)\ \ Cu^{2+}\ (aq) \rightarrow Cu^{2+}\ (aq) + Cu\ (s)$

Electroplating of metals



Function of electroplating:

1. To protect them from corrosion
2. Give them attractive appearance

Reactions involved:

Anode: $Cu\ (s) \rightarrow Cu^{2+}\ (aq) + 2e^-$

Cathode: $Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ (s)$

Overall equation: $Cu(s) + Cu^{2+}\ (aq) \rightarrow Cu^{2+}\ (aq) + Cu\ (s)$