



**CHEMISTRY • CLASS 12**

**CHAPTER 2**

# **Electrochemistry**

**NCERT SOLUTIONS**

## ORDER OF METAL DISPLACEMENT

2.1 Arrange Al, Cu, Fe, Mg and Zn in the order in which they displace one another from solutions of their salts.

A more reactive metal displaces a less reactive metal from its salt solution.



The order follows the tendency to lose electrons. More negative standard reduction potentials indicate stronger reducing metals.

| Metal | $E^\circ / \text{V}$ |
|-------|----------------------|
| Mg    | -2.36                |
| Al    | -1.66                |
| Zn    | -0.76                |
| Fe    | -0.44                |
| Cu    | +0.34                |

Each metal can displace metals to its right in the stated order.

## INCREASING REDUCING POWER

2.2 Given the standard electrode potentials below, arrange the metals in increasing order of reducing power.

| Metal | Given $E^\circ / \text{V}$ |
|-------|----------------------------|
| K     | -2.93                      |
| Ag    | +0.80                      |
| Hg    | +0.79                      |
| Mg    | -2.37                      |
| Cr    | -0.74                      |

### Solution:

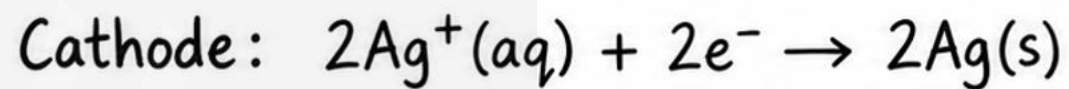
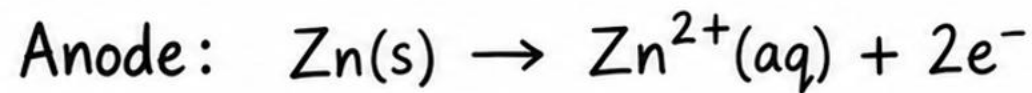
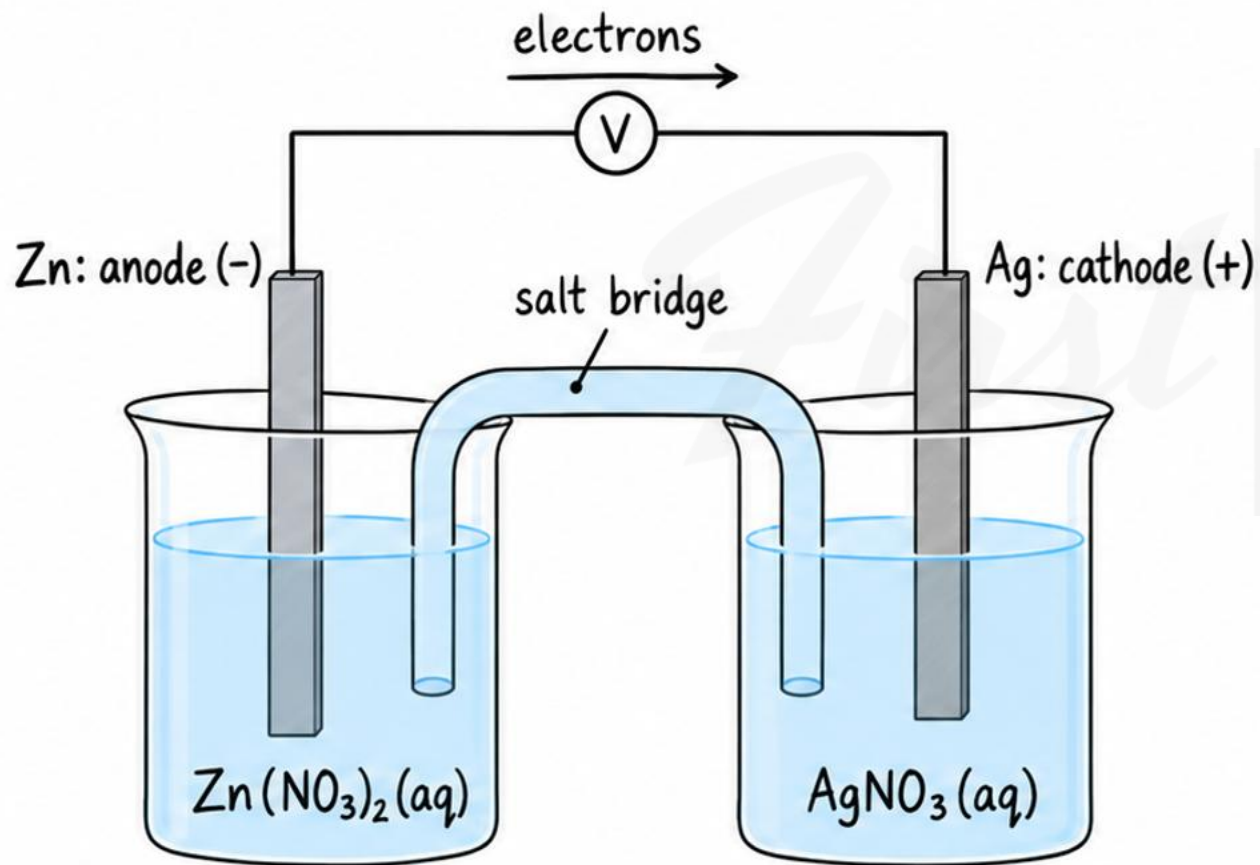
A reducing metal loses electrons. The more negative its standard reduction potential, the greater its reducing power.



Silver is the weakest reducing metal here; potassium is the strongest.

# ZINC-SILVER GALVANIC CELL

2.3 Depict the galvanic cell for  $\text{Zn(s)} + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag(s)}$ . Identify the negative electrode, current carriers and both electrode reactions.



Zinc is the negative electrode.

Electrons carry current through the external wire. Ions carry current through the solutions and salt bridge.

Bridge anions move towards the anode; cations move towards the cathode.



## CELL POTENTIAL, GIBBS ENERGY & K

2.4(i) For  $2\text{Cr(s)} + 3\text{Cd}^{2+}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 3\text{Cd(s)}$ , calculate  $E^\circ_{\text{cell}}$ ,  $\Delta_r G^\circ$  and the equilibrium constant at 298 K.

Chromium is oxidised and cadmium ions are reduced.

$$E^\circ(\text{Cr}^{3+}/\text{Cr}) = -0.74 \text{ V}; \quad E^\circ(\text{Cd}^{2+}/\text{Cd}) \approx -0.40 \text{ V}$$

The cadmium value is an additional rounded reference datum.

Solution:

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$E^\circ_{\text{cell}} = -0.40 - (-0.74) = +0.34 \text{ V}$$

$$n = 6; \quad F = 96500 \text{ C mol}^{-1}$$

$$\Delta_r G^\circ = -nFE^\circ_{\text{cell}} = -6 \times 96500 \times 0.34$$

$$\Delta_r G^\circ = -196860 \text{ J mol}^{-1} = -196.9 \text{ kJ mol}^{-1}$$

$$\log_{10} K = nE^\circ_{\text{cell}} / 0.0591 = (6 \times 0.34) / 0.0591 = 34.5178$$

$$K \approx 3.29 \times 10^{34}$$

The logarithm coefficient is for room temperature; do not multiply electrode potentials by stoichiometric coefficients.



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