# ELECTROCHEMISTRY

## Only concept Assessment Exercises (Solved)

## CONCEPT ASSESSMENT EXERCISE 2.1

### (1) $Cl_2(g) + 2NaOH(aq) \rightarrow NaCl(aq) + NaClO(aq) + H_2O(l)$

**Step 1: Assign oxidation states for all atoms**

- **$Cl_2$:** Cl = **0** (elemental form)
- **NaOH:**
  Na = **+1**, O = **–2**, H = **+1**
- **NaCl:**
  Na = **+1**, Cl = **–1**
- **NaClO:**
  Na = **+1**, Cl = **+1**, O = **–2**
- **$H_2O$:**
  H = **+1**, O = **–2**

**Step 2: Determine changes**

- **Cl:**
  0 → –1 (in NaCl) → **Reduction**
  0 → +1 (in NaClO) → **Oxidation**

**Conclusion:** Chlorine is **both oxidized and reduced** (disproportionation).

- **Oxidation half:** $Cl^0 \rightarrow Cl^{+1}$
- **Reduction half:** $Cl^0 \rightarrow Cl^{-1}$

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### (2) $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

**Step 1: Assign oxidation states**

- **Fe (solid):** Fe = **0**
- **$O_2$:** O = **0**
- **$Fe_2O_3$:**
  O = **–2** (rule for oxygen)
  Total O = 3 × (–2) = –6
  For neutrality: 2(Fe) + (–6) = 0 → Fe = **+3**

**Step 2: Determine changes**

- **Fe:** $0 \rightarrow +3 \rightarrow$ **Oxidation**
- **O:** $0 \rightarrow -2 \rightarrow$ **Reduction**

**Conclusion:**

- Fe is **oxidized** from 0 to +3.
- O is **reduced** from 0 to –2.

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### (3) $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$

**Step 1: Assign oxidation states**

- **$C_6H_{12}O_6$:**
  H = +1, O = –2
  Let average oxidation state of C = x

$6(x) + 12(+1) + 6(-2) = 0$
$6x + 12 - 12 = 0$
$6x = 0 \rightarrow \mathbf{x = 0}$

- **$O_2$:** O = 0
- **$CO_2$:**
  $O = -2 \rightarrow 2(-2) = -4$
  C must be +4 (to balance)
- **$H_2O$:**
  H = +1, O = –2



**Step 2: Determine changes**

- **C:** $0 \rightarrow +4 \rightarrow$ **Oxidation**
- **O (from $O_2$):** $0 \rightarrow -2 \rightarrow$ **Reduction**

**Conclusion:**

- Carbon is **oxidized** from 0 to +4.
- Oxygen is **reduced** from 0 to –2.

## CONCEPT ASSESSMENT EXERCISE 2.2

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### (1) $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$

**Step 1: Assign oxidation numbers**

- Cu (elemental) = **0**
- In **$Cu(NO_3)_2$**: Cu = **+2**
- In **$HNO_3$**: N = **+5**
- In **NO**: N = **+2**
- H in $H_2O$ = +1, O = –2

**Step 2: Identify changes**

- **Cu**: 0 → +2 (**oxidation**, loss of 2 $e^-$)
- **N**: +5 → +2 (**reduction**, gain of 3 $e^-$)

**Step 3: Equalize electron exchange**

- Cu loses **2 $e^-$** each
- N gains **3 $e^-$** each

LCM of 2 & 3 = **6**

Multiply:

- Oxidation (Cu) by 3 → **3 Cu → 3 $Cu^{2+}$ + 6 $e^-$**
- Reduction (N) by 2 → **2 $NO_3^-$ + 6 $e^-$ → 2 NO**

**Step 4: Build skeleton equation**

$$Cu + 2HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + H_2O$$

**Step 5: Balance atoms**

- Start with **Cu**: 3 Cu → **$3Cu(NO_3)_2$**
- Nitrate: 3 $Cu(NO_3)_2$ contains 6 $NO_3^-$, but only 2 $NO_3^-$ are reduced to NO.

**Adjust $HNO_3$ count:**

- Total $NO_3$ needed = **6 (in $Cu(NO_3)_2$) + 2 (to make NO) = 8 $HNO_3$**

**Step 6: Add water & H balance**

- H from 8 $HNO_3$ → 8 H → forms 4 $H_2O$.

**Balanced equation:**

$$\boxed{3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O}$$

### (2) $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

**Step 1: Assign oxidation numbers**

- Cu: 0 → +2 (oxidation)
- N: +5 → +4 (in $NO_2$, reduction)

**Step 2: Electron exchange**

- Cu loses 2 $e^-$
- N gains 1 $e^-$

LCM of 2 & 1 = **2**

- 1 Cu → 2 $e^-$
- 2 $NO_3^-$ (from 2 $HNO_3$) → 2 $NO_2$ (gain 2 $e^-$)

**Step 3: Build skeleton**

$$Cu + 2HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + H_2O$$

**Step 4: Count $NO_3$ groups**

- $Cu(NO_3)_2$ has 2 $NO_3^-$ (needs 2 from $HNO_3$)
- 2 $NO_2$ also come from $HNO_3$ → total N from $HNO_3$ = 4

**Step 5: Balance H and O**

- 4 H from $HNO_3$ → 2 $H_2O$

**Balanced equation:**

$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NN_2 + 2H_2O$$

## CONCEPT ASSESSMENT EXERCISE 2.3

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### (i) $Co + Cr^{3+} \rightarrow Co^{2+} + Cr^{2+}$

### Step 1: Assign oxidation numbers

- **Co (s):** 0 → **+2** (oxidation)
- **$Cr^{3+}$:** +3 → **+2** (reduction)

## Step 2: Write half-reactions

**Oxidation (Co):**

$$\mathbf{Co \rightarrow Co^{2+} + 2e^-}$$

**Reduction ($Cr^{3+}$):**

$$Cr^{3+} + e^- \rightarrow Cr^{2+}$$

## Step 3: Equalize electrons

- Oxidation releases **2 $e^-$**
- Reduction consumes **1 $e^-$**

Multiply the reduction half by 2:

$$2Cr^{3+} + 2e^- \rightarrow 2Cr^{2+}$$

## Step 4: Combine

$$Co + 2Cr^{3+} \rightarrow Co^{2+} + 2Cr^{2+}$$

**Balanced**

### (ii) $Au^{2+} + H_2 \rightarrow Au$

## Step 1: Oxidation numbers

- **$Au^{2+}$:** +2 → 0 (reduction)
- **$H_2$:** 0 → +1 (oxidation)

## Step 2: Half-reactions

**Reduction (Au):**

$$Au^{2+} + 2e^- \rightarrow Au$$

**Oxidation ($H_2$):**

$$H_2 \rightarrow 2H^+ + 2e^-$$

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## Step 3: Equalize electrons

Already balanced (2 $e^-$ in each).

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## Step 4: Combine

$$Au^{2+} + H_2 \rightarrow Au + 2H^+$$

**Balanced**

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(iii) $Cl_2 + I^- \longrightarrow I_2 + Cl^-$

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## Step 1: Oxidation numbers

- **$Cl_2$:** 0 → –1 (reduction)
- **$I^-$:** –1 → 0 (oxidation)

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## Step 2: Half-reactions

**Reduction ($Cl_2$):**

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

**Oxidation ($I^-$):**

$$I^- \rightarrow I_2 + 2e^-$$

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## Step 3: Equalize electrons

Already balanced (2 $e^-$ in each).

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## Step 4: Combine

$$Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$$

## CONCEPT ASSESSMENT EXERCISE 2.4

**Construct a redox equation from the following half-equations.**

Question 1: $MnO_4^- \rightarrow Mn^{2+}$ and $H_2O_2 \rightarrow O_2$

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## Step 1: Balance Each Half-Reaction

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### (a) Reduction Half-Reaction

$$+\boldsymbol{MnO_4^- \rightarrow Mn^{2+}}$$

- **Balance Mn atoms:** already balanced.
- **Balance O atoms:** add $\boldsymbol{4H_2O}$ to RHS.

$$\boldsymbol{MnO_4^- \rightarrow Mn^{2+} + 4H_2O}$$

- **Balance H atoms:** add $\boldsymbol{8H^+}$ to LHS.

$$\boldsymbol{MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O}$$

- **Balance charge:**
  LHS: (−1)+8(+1)=+7(-1) + 8(+1) = +7(−1)+8(+1)=+7
  RHS: +2+2+2
  Add **5e⁻** to LHS:

$$\boldsymbol{MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O}$$

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### (b) Oxidation Half-Reaction

$$\boldsymbol{2H_2O_2 \rightarrow O_2}$$

- **Balance O atoms:** already balanced.
- **Balance H atoms:** add $2H^+$ to RHS.

$$H_2O_2 \rightarrow O_2 + 2H^+$$

- **Balance charge:**
  LHS: 0
  RHS: +2+2+2
  Add **2e**⁻ to RHS:

$$H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$$

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## Step 2: Equalize Electrons

- Eq. (1) has **5e**⁻, Eq. (2) has **2e**⁻.
- **LCM = 10 electrons.**

Multiply Eq. (1) by 2:

2MnO4−+16H++10e−→2Mn2++8H2O(3)2MnO_4^- + 16H^+ + 10e^- \rightarrow 2Mn^{2+} + 8H_2O \tag{3}2MnO4−+16H++10e−→2Mn2++8H2O(3)

Multiply Eq. (2) by 5:

$$5H_2O_2 \rightarrow 5O_2 + 10H^+ + 10e^-$$



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## Step 3: Add and Cancel

Add (3) and (4):

$$2MnO_4^- + 16H^+ + 10e^- + 5H_2O_2 \rightarrow 2Mn^{2+} + 8H_2O + 5O_2 + 10H^+ + 10e^-$$

- Cancel **10e**⁻.
- Simplify $16H^+ - 10H^+ = 6H^+$.

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## Final Balanced Equation

$$\boxed{2MnO_4^- + 5H_2O_2 + 6H^+ \rightarrow 2Mn^{2+} + 5O_2 + 8H_2O}$$

**Question 2: $NO_3^-$ → NO and $Br^-$ → $Br_2$**

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## Step 1: Balance Each Half-Reaction

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### (a) Reduction Half-Reaction

$$NO_3^- \rightarrow NO$$

- **Balance N atoms:** already balanced.
- **Balance O atoms:** add $2H_2O$ to RHS.

$$NO_3^- \rightarrow NO + 2H_2O$$

- **Balance H atoms:** add $4H^+$to LHS.

$$NO_3^- + 4H^+ \rightarrow NO + 2H_2O$$

- **Balance charge:**
  LHS: (−1)+4(+1)=+3(-1) + 4(+1) = +3(−1)+4(+1)=+3
  RHS: 0
  Add **3e⁻** to LHS:

$$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$$

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### (b) Oxidation Half-Reaction



$$Br^- \rightarrow Br_2$$

- **Balance Br atoms:** write as

$$2Br^- \rightarrow Br_2$$

- **Balance charge:**
  LHS: $-2$
  RHS: 0
  Add **2e⁻** to RHS:

$$2Br^- \rightarrow Br_2 + 2e^-$$

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## Step 2: Equalize Electrons

- Eq. (1) has **3e⁻**, Eq. (2) has **2e⁻**.
- **LCM = 6 electrons.**

Multiply Eq. (1) by 2:

$$\boldsymbol{2NO_3^- + 8H^+ + 6e^- \rightarrow 2NO + 4H_2O}$$

Multiply Eq. (2) by 3:

$$\boldsymbol{6Br^- \rightarrow 3Br_2 + 6e^-}$$

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## Step 3: Add and Cancel

Add (3) and (4):

$$\boldsymbol{2NO_3^- + 8H^+ + 6e^- + 6Br^- \rightarrow 2NO + 4H_2O + 3Br_2 + 6e^-}$$

Cancel **6e⁻**.

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## Final Balanced Equation

$$\boxed{\boldsymbol{2NO_3^- + 6Br^- + 8H^+ \rightarrow 2NO + 3Br_2 + 4H_2O}}$$

## Question 3: $\boldsymbol{MnO_4^- \rightarrow Mn^{2+} and\ CO_4^{2-} \rightarrow CO_2}$

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## Step 1: Balance Each Half-Reaction

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### (a) Reduction Half-Reaction

$$MnO_4^- \rightarrow Mn^{2+}$$

- **Balance Mn atoms:** already balanced.
- **Balance O atoms:** add $4H_2O$to RHS.

$$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$$

- **Balance H atoms:** add $8H^+$to LHS.

$$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$$

- **Balance charge:**
  LHS: (−1)+8(+1)=+7
  RHS: +2+2+2
  Add **5e⁻** to LHS:

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

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**(b) Oxidation Half-Reaction**

$$C_2O_4^{2-} \rightarrow CO_2$$

- **Balance C atoms:**
  Each oxalate gives **2 $CO_2$**:

$$C_2O_4^{2-} \rightarrow 2CO_2$$

- **Balance charge:**
  Carbon goes from **+3** in oxalate to **+4** in $CO_2$ → each C loses **1e⁻**, so **2e⁻** lost total.

Add **2e⁻** to RHS:

$$C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$$

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## Step 2: Equalize Electrons

- Eq. (1) has **5e⁻**, Eq. (2) has **2e⁻**.
- **LCM = 10 electrons.**

Multiply Eq. (1) by 2:

$$2MnO_4^- + 16H^+ + 10e^- \rightarrow 2Mn^{2+} + 8H_2O$$

Multiply Eq. (2) by 5:

$$5C_2O_4^{2-} \rightarrow 10CO_2 + 10e^-$$

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## Step 3: Add and Cancel

Add (3) and (4):

$$2MnO_4^- + 16H^+ + 10e^- + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2 + 10e^-$$

- Cancel **$10e^-$**.

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### Final Balanced Equation

$$\boxed{2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O}$$

## CONCEPT ASSESSMENT EXERCISE 2.5

### Reaction

$$\mathrm{CuO} + \mathrm{H_2} \rightarrow \mathrm{Cu} + H_2O$$

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### Step 1: Assign oxidation numbers

- **Cu** in CuO = **+2** (since O = –2)
- **H** in $H_2$ = **0**
- **Cu** in Cu (solid) = **0**
- **H** in $H_2O$ = **+1**

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### Step 2: Identify changes

- **Cu:** +2 → 0 (**reduction**)
- **H:** 0 → +1 (**oxidation**)

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### Step 3: Identify agents

- The substance that **is reduced** is the **oxidizing agent**.
  → **CuO** ($Cu^{2+}$ is reduced to $Cu^0$) is the **oxidizing agent**.
- The substance that **is oxidized** is the **reducing agent**.
  → **$H_2$** ($H^0$ is oxidized to $H^+$) is the **reducing agent**.

## CONCEPT ASSESSMENT EXERCISE 2.6

### Given Data

- $Cu^{2+} + 2e^- \rightarrow Cu(s) \quad E^\circ = +0.34\ V$
- $F_2(g) + 2e^- \rightarrow 2F^- \quad E^\circ = +2.87\ V$

We want to:

1. $Estimate\boldsymbol{E^\circ_{cell}}$
2. **Write the cell reaction**
3. **Choose the cathode**
4. **Show the direction of electron flow**

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## Step 1: Identify which half-reaction is reduction and which is oxidation

- The species with the **higher** reduction potential is **more easily reduced**.
  $F_2$ has **+2.87 V**, much higher than $Cu^{2+}$ (+0.34 V).

**At cathode (reduction):**

$$F_2(g) + 2e^- \rightarrow 2F^-(aq)$$

**At anode (oxidation):** Reverse the copper half-reaction:

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

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## Step 2: Write the overall cell reaction

$$Cu(s) + F_2(g) \rightarrow Cu^{2+}(aq) + 2F^-(aq)$$

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## Step 3: Calculate Ecell∘E^\circ_{\text{cell}}Ecell∘

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

- $E^\circ_{cathode} = +2.87\ V$ F_2 reduction
- $E^\circ_{anode} = +0.34\ V \quad (Cu^{2+}/Cu\ reduction\ potential)$

$$E^\circ_{cell} = 2.87 - 0.34 = \boxed{2.53\ V}$$

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## Step 4: Identify cathode & anode

- **Cathode (reduction): $F_2$ gas electrode** (because $F_2$ is reduced)
- **Anode (oxidation): Copper metal electrode**

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## Step 5: Direction of electron flow

- **Electrons flow** from **anode → cathode**, i.e. **from Cu (copper) to $F_2$.**

## CONCEPT ASSESSMENT EXERCISE 2.7

## 1. Using emf data, argue on the following:

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**(i) Can Mg displace Cu from a solution of Copper(II) sulphate?**

**Reaction to check:**

$$\boldsymbol{Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)}$$

**Step 1: Write the half-reactions with standard reduction potentials (E°).**

$$\boldsymbol{Cu^{2+} + 2e^- \rightarrow Cu \quad E^\circ = +0.34\ \mathrm{V}}$$

$$\boldsymbol{Mg^{2+} + 2e^- \rightarrow Mg \quad E^\circ = -2.37\ \mathrm{V}}$$

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**Step 2: Determine which species is oxidized and which is reduced.**

- Magnesium will **oxidize** (lose electrons):

$$Mg \rightarrow Mg^{2+} + 2e^- \quad E^\circ = +2.37\ \mathrm{V}\ \text{(reverse sign)}$$

- Copper will **reduce** (gain electrons):

$$Cu^{2+} + 2e^- \rightarrow Cu \quad E^\circ = +0.34\ \mathrm{V}$$

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**Step 3: Calculate the EMF.**

$$E^\circ{}_{cell} = E^\circ{}_{cathode} - E^\circ{}_{anode}$$

$$= (+0.34) - (-2.37) = +2.71\ V$$

**Conclusion:**
Since **E°cell is positive**, the reaction is **spontaneous**.
**Magnesium can displace copper** from copper(II) sulphate solution.

**(ii) Can iodine displace chlorine from an aqueous solution of potassium chloride?**

**Reaction to check:**

$$I_2(s) + 2Cl^-(aq) \longrightarrow 2I^-(aq) + Cl_2(g)$$

**Step 1: Write the relevant reduction potentials.**

$$Cl_2 + 2e^- \rightarrow 2Cl^- \quad E° = +1.36\ VC$$

$$I_2 + 2e^- \rightarrow 2I^- \quad E° = +0.54\ \mathrm{V}$$

**Step 2: Decide redox direction.**

- To displace chlorine, iodine would need to **oxidize $Cl^-$ to $Cl_2$**.
- This means $Cl^- \rightarrow Cl_2$ must be easier than $I^- \rightarrow I_2$.

**Step 3: Compare reduction potentials.**

- Chlorine has a **higher E° (+1.36 V)** than iodine **(+0.54 V)**.
- The halogen with **higher E°** is **stronger oxidizing agent**.

Thus, **$Cl_2$ can oxidize $I^-$**, but **$I_2$ cannot oxidize $Cl^-$**.

**Conclusion:**
**Iodine cannot displace chlorine** from potassium chloride solution.

## 2. Is the following reaction feasible?

$$Al + Mn^{2+} \rightarrow Al^{3+} + Mn$$

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**Step 1: Write the standard reduction potentials.**

$$Al^{3+} + 3e^{-} \rightarrow Al \quad E^{\circ} = -1.66\ \mathrm{V}A$$

$$Mn^{2+} + 2e^{-} \rightarrow Mn \quad E^{\circ} = -1.18\ \mathrm{V}$$

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**Step 2: Set up half-reactions.**

- **Oxidation ($Al \rightarrow Al^{3+}$):**

$$Al \rightarrow Al^{3+} + 3e^{-} \quad E^{\circ} = +1.66\ \mathrm{V}$$

- **Reduction ($Mn^{2+} \rightarrow Mn$):**

$$Mn^{2+} + 2e^{-} \rightarrow Mn \quad E^{\circ} = -1.18\ \mathrm{V}$$

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**Step 3: Calculate EMF.**

$$E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$$

$$= (-1.18) - (-1.66) = +0.48\ \mathrm{V}$$

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**Conclusion:**
Since **E°cell is positive**, the reaction is **feasible**.

Aluminium **will displace manganese** from $Mn^{2+}$ solution.

## CONCEPT ASSESSMENT EXERCISE 2.8

## Given Data:

- Standard electrode potential:

$$E \circ (Fe3+/Fe2+) = +0.77\ V$$

- Concentration of $Fe^{3+} = 1.0$ M

- Concentration of $Fe^{2+} = 0.2$ M
- Temperature: **298 K (25 °C)**

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### Step 1: Write the half-reaction

$$Fe^{3+} + e^{-} \longrightarrow Fe^{2+}$$

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### Step 2: Write the Nernst equation

$$E = E^{\circ} - \frac{0.0591}{n} log \frac{[\mathrm{Fe}^{2+}]}{[\mathrm{Fe}^{3+}]}$$

Here:

- $n = 1$ (1 electron involved)
- $[Fe^{2+}] = 0.2$ M
- $[Fe^{3+}] = 1.0$ M

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### Step 3: Substitute values

$$E = 0.77 - \frac{0.0591}{1} \log \frac{0.2}{1.0}$$

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### Step 4: Simplify the log term

$$\log \frac{0.2}{1.0} = \log(0.2) = -0.699$$

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### Step 5: Calculate

$$E = 0.77 - (0.0591)(-0.699)$$

$$E = 0.77 + 0.0413$$

$$E = 0.811\, V$$

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**Final Answer:**

$$\boxed{E = 0.811\ \text{V}}$$

# CONCEPT ASSESSMENT EXERCISE 2.9

## a) Magnesium ribbon in a solution of silver nitrate

**Reaction to consider:**

$$Mg(s) + 2AgNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2Ag(s)$$

**Reasoning:**
Magnesium is **more reactive** (higher in the activity series) than silver.
It can lose electrons easily and displace silver ions from solution.

**Conclusion:**
**Reaction will occur.** Magnesium will dissolve, and **silver metal** will be deposited.

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## b) Copper plate dipped in iron(II) sulphate solution

**Reaction to consider:**

$$Cu(s) + FeSO_4(aq) \rightarrow CuSO_4(aq) + Fe(s)$$

**Reasoning:**
Copper is **less reactive** than iron in the reactivity series.
It **cannot** displace $Fe^{2+}$ ions from solution.

**Conclusion:**
**No reaction will occur.**

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## c) Lead(II) oxide heated with powdered zinc

**Reaction to consider (redox on heating):**

$$Zn(s) + PbO(s) \xrightarrow{\Delta} ZnO(s) + Pb(s)$$

**Reasoning:**
Zinc is **more reactive** than lead and will **reduce PbO to metallic Pb**, forming zinc oxide.

**Conclusion:**
**Reaction will occur**; **lead metal** will be produced.

**d) Nickel wire in a solution of silver nitrate**

**Reaction to consider:**

$$Ni(s) + 2AgNO_3(aq) \rightarrow Ni(NO_3)_2(aq) + 2Ag(s)$$

- **Reasoning:**
  Nickel is **more reactive** than silver.
  It can displace $Ag^+$ from solution.

**Conclusion:**
**Reaction will occur.** Nickel will go into solution, and **silver crystals** will deposit.

## CONCEPT ASSESSMENT EXERCISE 2.10

### Question No 1: Mass of substance liberated in 2 hours

**Given:**

- Time = 2 hours = $2 \times 3600 = 7200\ s$
- Current = 0.5 A
- Molar mass of substance = 107.9 g/mol
- Assume the substance is **monovalent** (like $Ag^+$) → **n = 1**

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**Step 1: Calculate charge (Q)**

$$Q = I \times t = 0.5 \times 7200 = 3600\ C$$

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**Step 2: Moles of electrons**

$$Moles\ of\ e^- = \frac{Q}{F} = \frac{3600}{96500} = 0.0373 \text{ mol e}^-$$

$$(Faraday's\ constant F = 96500\ C/mol)$$

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**Step 3: Moles of substance deposited**

For a **monovalent ion**:

$$Moles\ of\ substance = Moles\ of\ e^- = 0.0373\ mol$$

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**Step 4: Mass deposited**

$$Mass = Moles \times Molar\ mass =$$

$$0.0373 \times 107.9 = 4.02\ g$$

**Answer:**

$$\boxed{4.02 \text{ g of the substance}}$$

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## Question No 2: Charge when 3 moles of electrons flow

**Step 1: Use Faraday's law**

$$Q = n \times F$$

Where:

- $n = 3\ mol\ e^-$
- $F = 96500\ C/mol$

$$Q = 3 \times 96500 = 289{,}500\ C$$

**Answer:**



$$\boxed{2.895 \times 10^5 \text{ C}}$$

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## Question No 3: Mass of silver deposited from $AgNO_3$ solution

**Given:**

- Current = 0.1 A
- Time = 20 minutes = $20 \times 60 = 1200\ s$
- Molar mass of silver = 107.9 g/mol
- $Ag^+$ is **monovalent → n = 1**

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**Step 1: Calculate charge**

$$Q = I \times t = 0.1 \times 1200 = 120\ C$$

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**Step 2: Moles of electrons**

$$\text{Moles of e}^- = \frac{Q}{F} = \frac{120}{96500} = 0.00124\ \text{mol}$$

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**Step 3: Moles of Ag deposited**

For $Ag^+$ (n=1), moles of Ag = moles of $e^-$ = 0.00124 mol.

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**Step 4: Mass of Ag**

$$\text{Mass} = 0.00124 \times 107.9 = 0.134\ \text{g}$$

**Answer:**

$$\boxed{0.134\ \text{g of Ag deposited}}$$