

ELECTROCHEMISTRY**Only concept Assessment Exercises (Solved)****CONCEPT ASSESSMENT EXERCISE 2.1**

Step 1: Assign oxidation states for all atoms

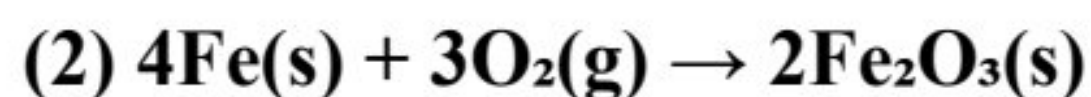
- **Cl₂:** Cl = 0 (elemental form)
- **NaOH:**
Na = +1, O = -2, H = +1
- **NaCl:**
Na = +1, Cl = -1
- **NaClO:**
Na = +1, Cl = +1, O = -2
- **H₂O:**
H = +1, O = -2

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



Step 1: Assign oxidation states

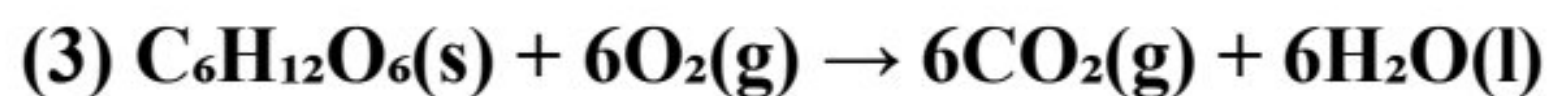
- **Fe (solid):** Fe = 0
- **O₂:** O = 0
- **Fe₂O₃:**
O = -2 (rule for oxygen)
Total O = 3 × (-2) = -6
For neutrality: 2(Fe) + (-6) = 0 → Fe = +3

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.

**Step 1: Assign oxidation states**

- **$\text{C}_6\text{H}_{12}\text{O}_6$:**
 $\text{H} = +1, \text{O} = -2$
 Let average oxidation state of C = x

$$6(x) + 12(+1) + 6(-2) = 0$$

$$6x + 12 - 12 = 0$$

$$6x = 0 \rightarrow x = 0$$

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

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**Step 1: Assign oxidation numbers**

- Cu (elemental) = 0
- In $\text{Cu}(\text{NO}_3)_2$: Cu = +2
- In HNO_3 : N = +5
- In NO : N = +2
- H in H_2O = +1, O = -2

Step 2: Identify changes

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

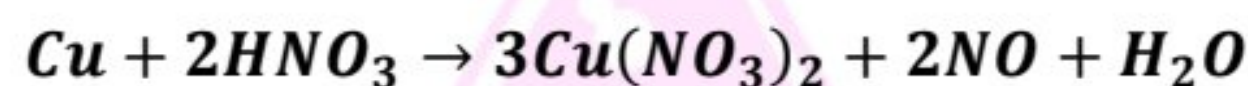
Step 3: Equalize electron exchange

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

LCM of 2 & 3 = 6

Multiply:

- Oxidation (Cu) by 3 $\rightarrow 3 \text{Cu} \rightarrow 3 \text{Cu}^{2+} + 6 \text{e}^-$
- Reduction (N) by 2 $\rightarrow 2 \text{NO}_3^- + 6 \text{e}^- \rightarrow 2 \text{NO}$

Step 4: Build skeleton equation**Step 5: Balance atoms**

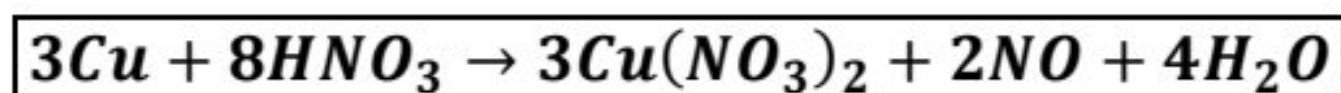
- Start with Cu: $3 \text{Cu} \rightarrow 3\text{Cu}(\text{NO}_3)_2$
- Nitrate: $3 \text{Cu}(\text{NO}_3)_2$ contains 6 NO_3^- , but only 2 NO_3^- are reduced to NO.

Adjust HNO_3 count:

- Total NO_3 needed = 6 (in $\text{Cu}(\text{NO}_3)_2$) + 2 (to make NO) = 8 HNO_3

Step 6: Add water & H balance

- H from 8 $\text{HNO}_3 \rightarrow 8 \text{H} \rightarrow$ forms 4 H_2O .

Balanced equation:

**Step 1: Assign oxidation numbers**

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

Step 2: Electron exchange

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

LCM of 2 & 1 = 2

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

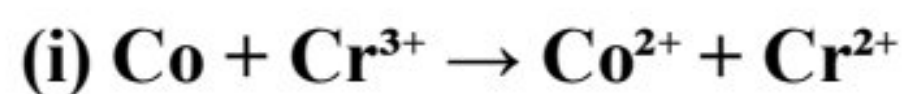
Step 3: Build skeleton**Step 4: Count NO_3 groups**

- $\text{Cu}(\text{NO}_3)_2$ has 2 NO_3^- (needs 2 from HNO_3)
- 2 NO_2 also come from $\text{HNO}_3 \rightarrow$ total N from $\text{HNO}_3 = 4$

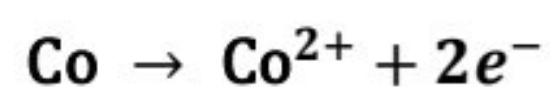
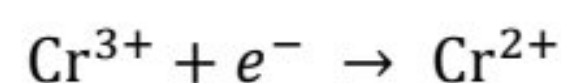
Step 5: Balance H and O

- 4 H from $\text{HNO}_3 \rightarrow$ 2 H_2O

Balanced equation:

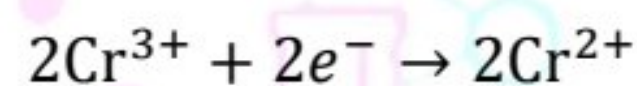
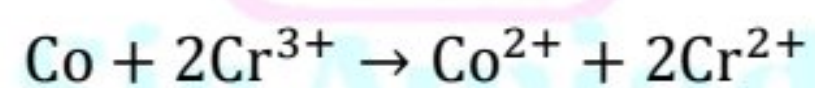
**CONCEPT ASSESSMENT EXERCISE 2.3****Step 1: Assign oxidation numbers**

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

Step 2: Write half-reactions**Oxidation (Co):****Reduction (Cr^{3+}):****Step 3: Equalize electrons**

- Oxidation releases $2e^{-}$
- Reduction consumes $1e^{-}$

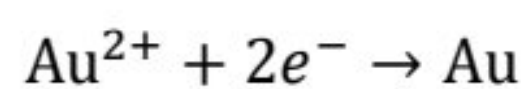
Multiply the reduction half by 2:

**Step 4: Combine****Balanced**

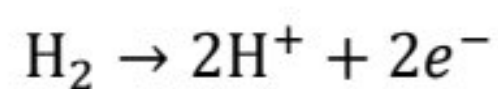
(ii) $\text{Au}^{2+} + \text{H}_2 \rightarrow \text{Au}$ **Step 1: Oxidation numbers**

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

Step 2: Half-reactions**Reduction (Au):**



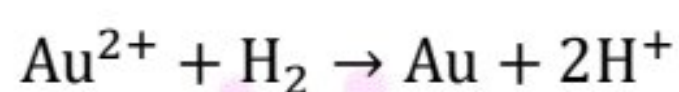
Oxidation (H_2):



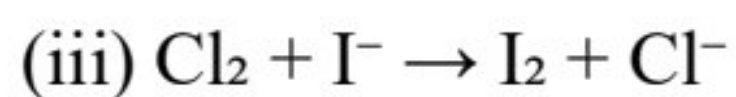
Step 3: Equalize electrons

Already balanced (2 e^{-} in each).

Step 4: Combine



Balanced

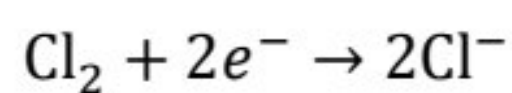


Step 1: Oxidation numbers

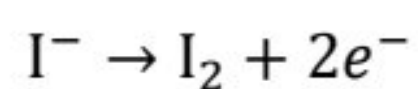
- Cl_2 : $0 \rightarrow -1$ (reduction)
 - I^{-} : $-1 \rightarrow 0$ (oxidation)
-

Step 2: Half-reactions

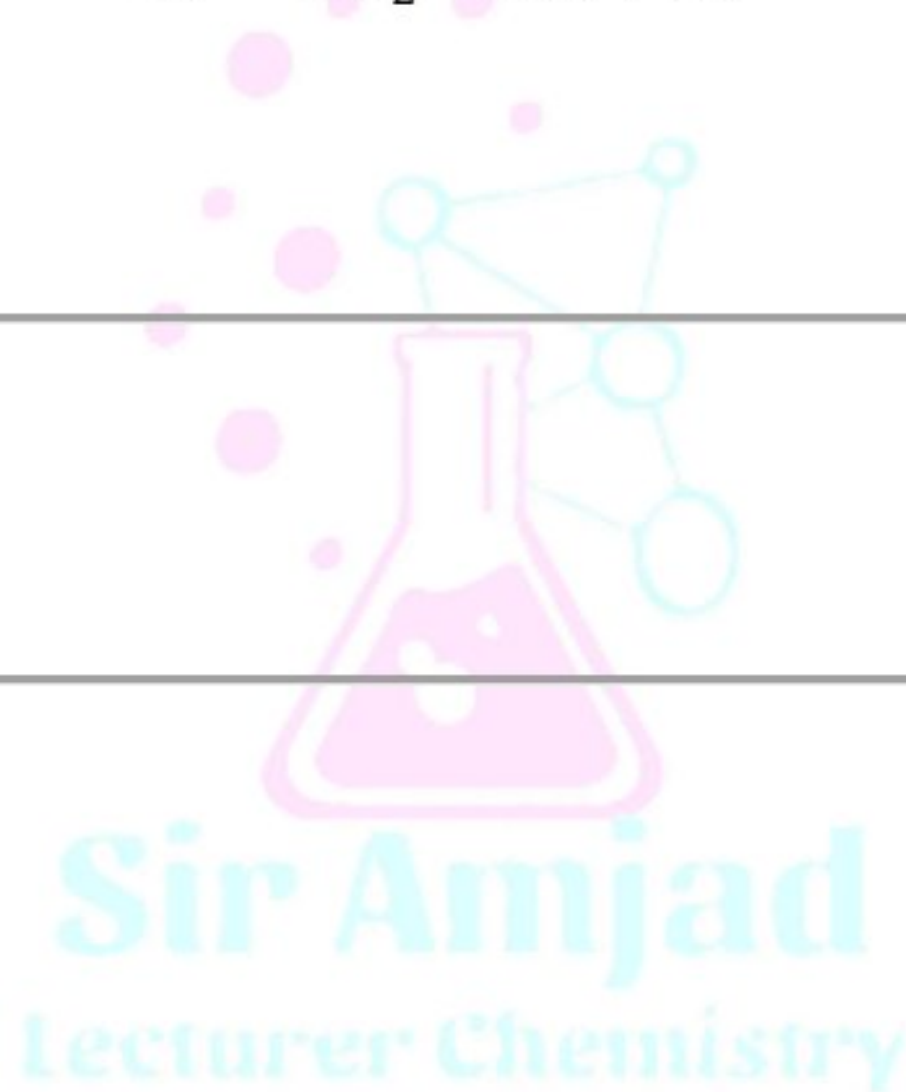
Reduction (Cl_2):



Oxidation (I^{-}):

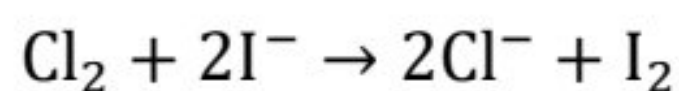


Step 3: Equalize electrons



Already balanced (2 e⁻ in each).

Step 4: Combine



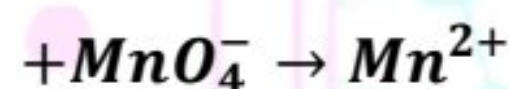
CONCEPT ASSESSMENT EXERCISE 2.4

Construct a redox equation from the following half-equations.

Question 1: $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$ and $\text{H}_2\text{O}_2 \rightarrow \text{O}_2$

Step 1: Balance Each Half-Reaction

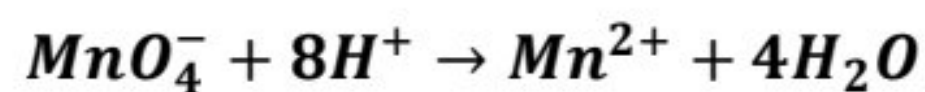
(a) Reduction Half-Reaction



- **Balance Mn atoms:** already balanced.
- **Balance O atoms:** add $4\text{H}_2\text{O}$ to RHS.



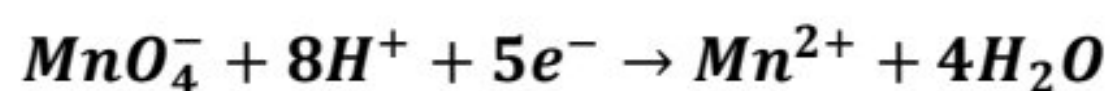
- **Balance H atoms:** add 8H^+ to LHS.



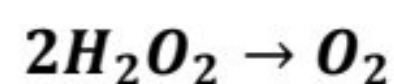
- **Balance charge:**

LHS: $(-1) + 8(+1) = +7$
RHS: $+2 + 2 + 2 = +6$

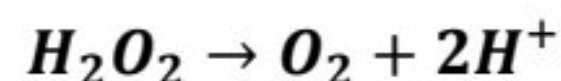
Add 5e^- to LHS:



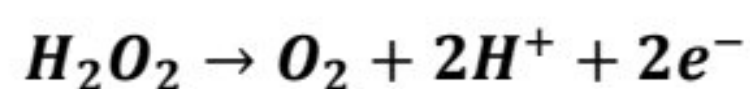
(b) Oxidation Half-Reaction



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



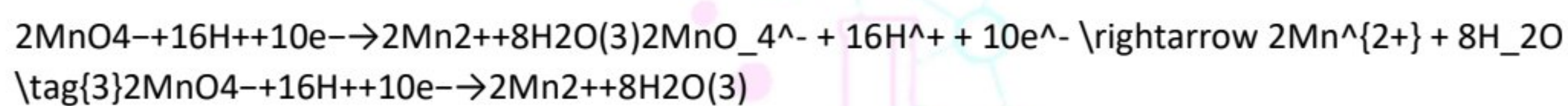
- **Balance charge:**
LHS: 0
RHS: +2+2+2
Add $2e^-$ to RHS:



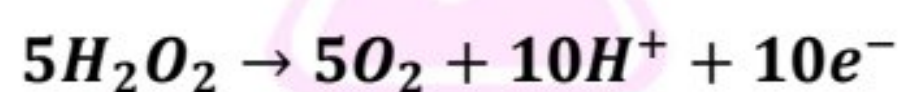
Step 2: Equalize Electrons

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

Multiply Eq. (1) by 2:

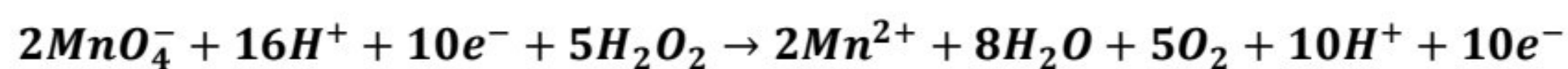


Multiply Eq. (2) by 5:



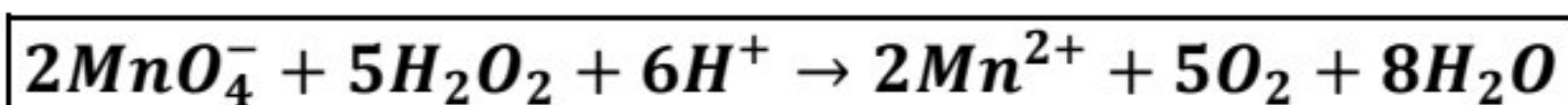
Step 3: Add and Cancel

Add (3) and (4):



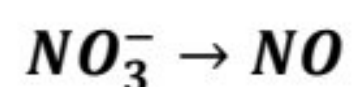
- Cancel $10e^-$.
- Simplify $16H^+ - 10H^+ = 6H^+$.

Final Balanced Equation



Question 2: $\text{NO}_3^- \rightarrow \text{NO}$ and $\text{Br}^- \rightarrow \text{Br}_2$

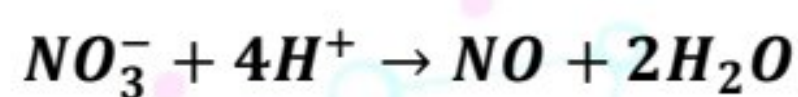
Step 1: Balance Each Half-Reaction

(a) Reduction Half-Reaction


- **Balance N atoms:** already balanced.
- **Balance O atoms:** add $2\text{H}_2\text{O}$ to RHS.



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

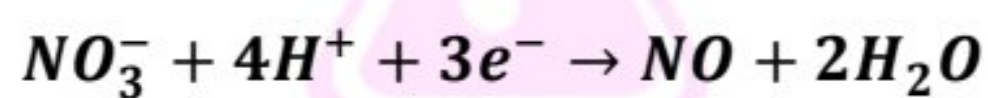
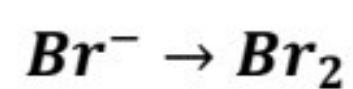


- **Balance charge:**

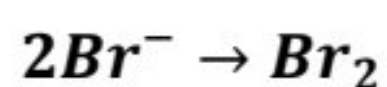
$$\text{LHS: } (-1) + 4(+1) = +3 \quad \text{RHS: } (-1) + 4(+1) = +3$$

$$\text{RHS: } 0$$

Add 3e^- to LHS:


(b) Oxidation Half-Reaction


- **Balance Br atoms:** write as



- **Balance charge:**

$$\text{LHS: } -2$$

$$\text{RHS: } 0$$

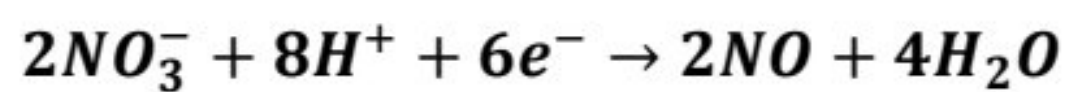
Add 2e^- to RHS:



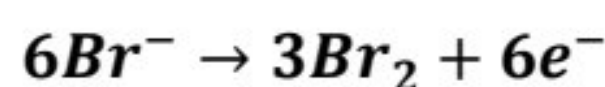
Step 2: Equalize Electrons

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

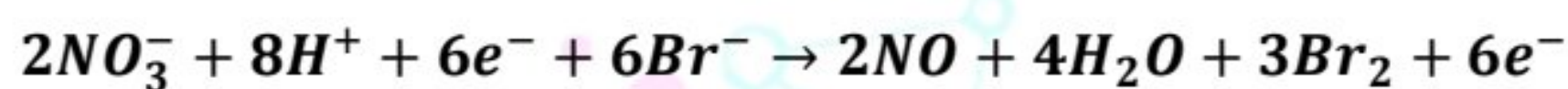
Multiply Eq. (1) by 2:



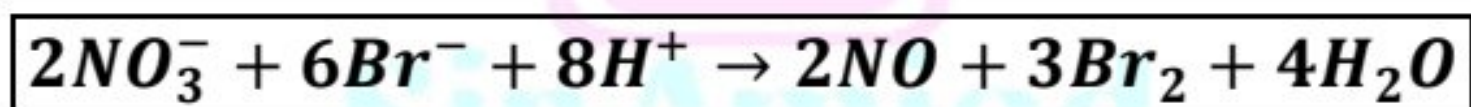
Multiply Eq. (2) by 3:

**Step 3: Add and Cancel**

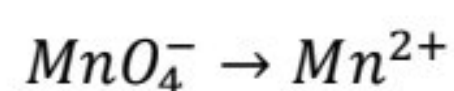
Add (3) and (4):



Cancel $6e^-$.

Final Balanced Equation

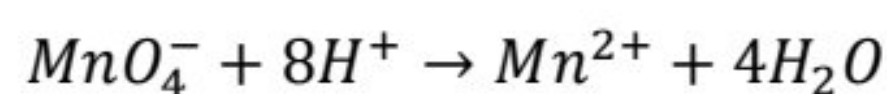
Question 3: $MnO_4^- \rightarrow Mn^{2+}$ and $CO_3^{2-} \rightarrow CO_2$

Step 1: Balance Each Half-Reaction**(a) Reduction Half-Reaction**

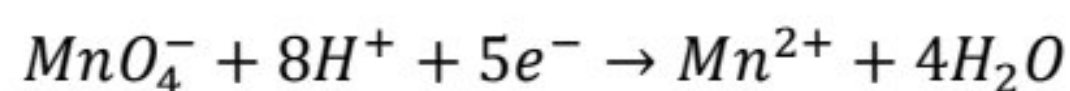
- Balance Mn atoms:** already balanced.
- Balance O atoms:** add $4H_2O$ to RHS.



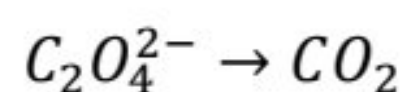
- Balance H atoms:** add $8H^+$ to LHS.



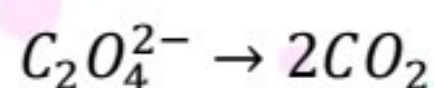
- **Balance charge:**
LHS: $(-1) + 8(+1) = +7$
RHS: $+2 + 2 + 2$
Add 5e^- to LHS:



(b) Oxidation Half-Reaction

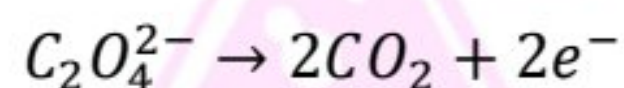


- **Balance C atoms:**
Each oxalate gives 2CO_2 :



- **Balance charge:**
Carbon goes from $+3$ in oxalate to $+4$ in $\text{CO}_2 \rightarrow$ each C loses 1e^- , so 2e^- lost total.

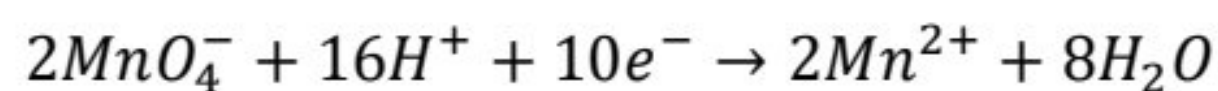
Add 2e^- to RHS:



Step 2: Equalize Electrons

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

Multiply Eq. (1) by 2:

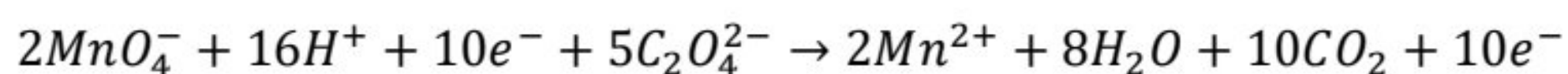


Multiply Eq. (2) by 5:



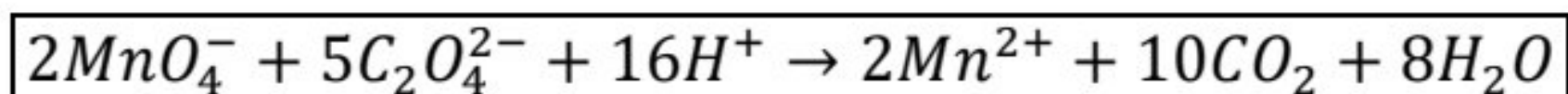
Step 3: Add and Cancel

Add (3) and (4):



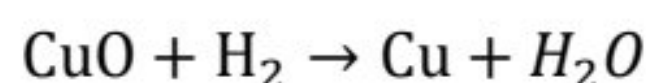
- Cancel 10e^- .

Final Balanced Equation



CONCEPT ASSESSMENT EXERCISE 2.5

Reaction



Step 1: Assign oxidation numbers

- **Cu** in $\text{CuO} = +2$ (since $\text{O} = -2$)
- **H** in $\text{H}_2 = 0$
- **Cu** in Cu (solid) = **0**
- **H** in $\text{H}_2\text{O} = +1$

Step 2: Identify changes

- **Cu:** $+2 \rightarrow 0$ (**reduction**)
- **H:** $0 \rightarrow +1$ (**oxidation**)

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**.
→ **H₂** (H^0 is oxidized to H^+) is the **reducing agent**.

CONCEPT ASSESSMENT EXERCISE 2.6

Given Data

- $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}(s) \quad E^\circ = +0.34 \text{ V}$
- $\text{F}_2(g) + 2e^- \rightarrow 2\text{F}^- \quad E^\circ = +2.87 \text{ V}$

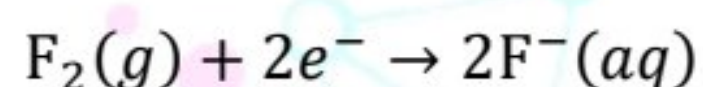
We want to:

1. Estimate E°_{cell}
2. Write the cell reaction
3. Choose the cathode
4. Show the direction of electron flow

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).

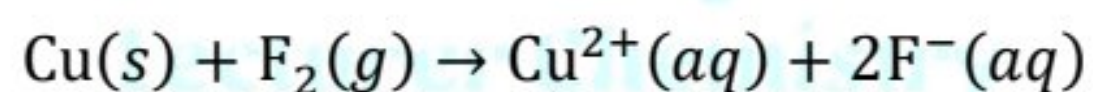
At cathode (reduction):



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



Step 2: Write the overall cell reaction



Step 3: Calculate E°_{cell}

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

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

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

Step 4: Identify cathode & anode

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

Step 5: Direction of electron flow

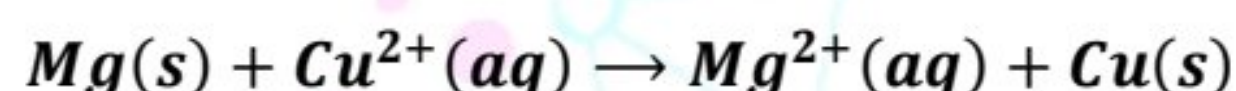
- Electrons flow from anode $\rightarrow$ cathode, i.e. from Cu (copper) to F_2 .

CONCEPT ASSESSMENT EXERCISE 2.7

1. Using emf data, argue on the following:

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

Reaction to check:

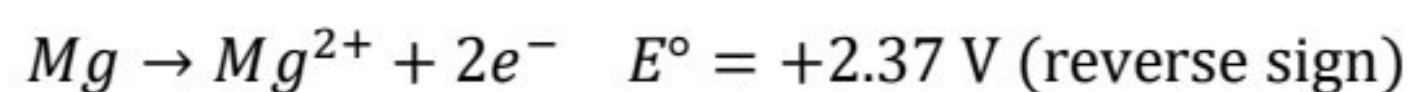


Step 1: Write the half-reactions with standard reduction potentials (E°).



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

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



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



Step 3: Calculate the EMF.

$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} \\ &= (+0.34) - (-2.37) = +2.71 \text{ V} \end{aligned}$$

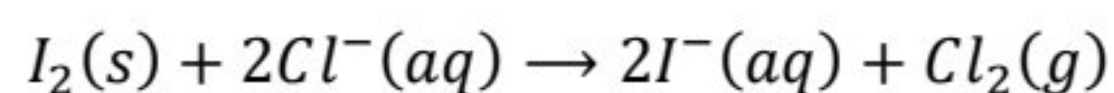
Conclusion:

Since E°_{cell} is **positive**, the reaction is **spontaneous**.

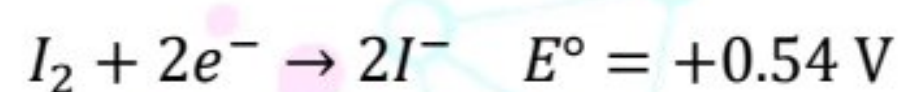
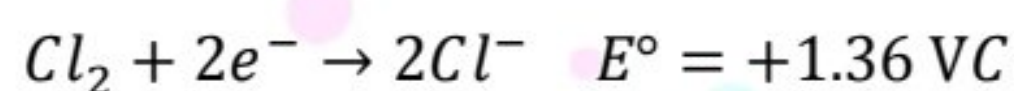
Magnesium can displace copper from copper(II) sulphate solution.

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

Reaction to check:



Step 1: Write the relevant reduction potentials.



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 $I^- \rightarrow I_2$.

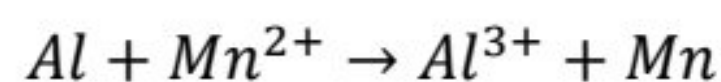
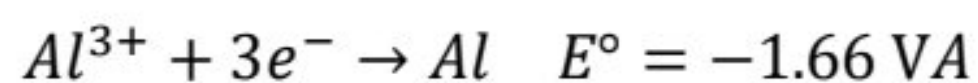
Step 3: Compare reduction potentials.

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

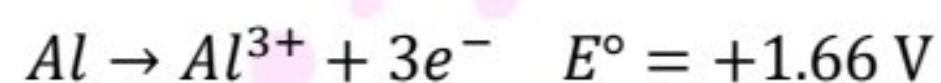
Thus, **Cl_2 can oxidize I^-** , but **I_2 cannot oxidize Cl^-** .

Conclusion:

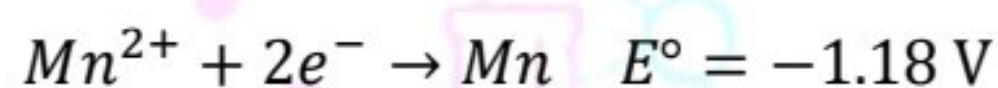
Iodine cannot displace chlorine from potassium chloride solution.

2. Is the following reaction feasible?**Step 1: Write the standard reduction potentials.****Step 2: Set up half-reactions.**

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



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

**Step 3: Calculate EMF.**

$$\begin{aligned} E^{\circ}_{\text{cell}} &= E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} \\ &= (-1.18) - (-1.66) = +0.48 \text{ V} \end{aligned}$$

Conclusion:

Since E°_{cell} is positive, the reaction is **feasible**.

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

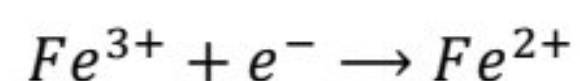
CONCEPT ASSESSMENT EXERCISE 2.8**Given Data:**

- Standard electrode potential:

$$E^{\circ} (Fe^{3+}/Fe^{2+}) = +0.77 \text{ V}$$

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

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

Step 1: Write the half-reaction**Step 2: Write the Nernst equation**

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

Here:

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

Step 3: Substitute values

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

Step 4: Simplify the log term

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

Step 5: Calculate

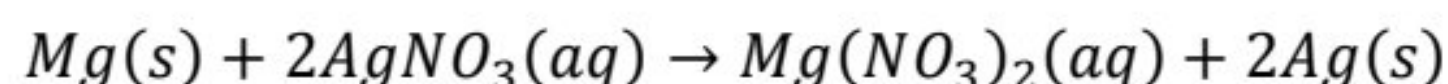
$$E = 0.77 - (0.0591)(-0.699)$$

$$E = 0.77 + 0.0413$$

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

Final Answer:

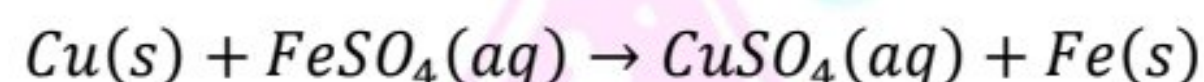
$$E = 0.811 \text{ V}$$

CONCEPT ASSESSMENT EXERCISE 2.9**a) Magnesium ribbon in a solution of silver nitrate****Reaction to consider:****Reasoning:**

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

Conclusion:

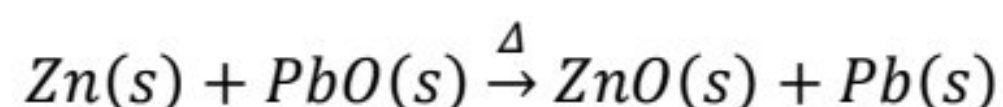
Reaction will occur. Magnesium will dissolve, and **silver metal** will be deposited.

b) Copper plate dipped in iron(II) sulphate solution**Reaction to consider:****Reasoning:**

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

Conclusion:

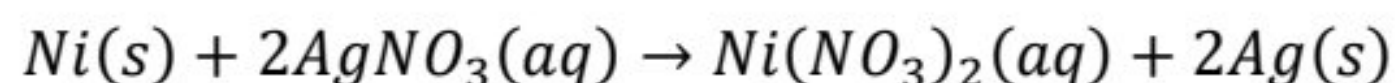
No reaction will occur.

c) Lead(II) oxide heated with powdered zinc**Reaction to consider (redox on heating):****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:**

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

Step 1: Calculate charge (Q)

$$Q = I \times t = 0.5 \times 7200 = 3600 \text{ C}$$

Step 2: Moles of electrons

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

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

Step 3: Moles of substance depositedFor a **monovalent ion**:

$$\text{Moles of substance} = \text{Moles of } e^- = 0.0373 \text{ mol}$$

Step 4: Mass deposited

$$\begin{aligned}\text{Mass} &= \text{Moles} \times \text{Molar mass} = \\ 0.0373 \times 107.9 &= 4.02 \text{ g}\end{aligned}$$

Answer:

4.02 g of the substance

Question No 2: Charge when 3 moles of electrons flow**Step 1: Use Faraday's law**

$$Q = n \times F$$

Where:

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

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

Answer:

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

Question No 3: Mass of silver deposited from AgNO_3 solution**Given:**

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

Step 1: Calculate charge

$$Q = I \times t = 0.1 \times 1200 = 120 \text{ C}$$

Step 2: Moles of electrons

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

Step 3: Moles of Ag deposited

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

Step 4: Mass of Ag

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

Answer:

0.134 g of Ag deposited

For More Information:

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