# تغطية تفاعلات الأكسدة والاختزال – الجزء الأول () نهاية الفصل (غير محلول)




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# Mr. Mouad استاذ معاذ

# 0557903129

# EoT1 Chemistry Coverage

# Grade 12 Adv

# 2026-2027



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# Grade 12 Adv

# Part-1 (Q1-Q11)

# 2026-2027




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| | | | |
|---|---|---|---|
| Q1 | Distinguish between oxidation and reduction in terms of loss and gain of electrons, oxygen and hydrogen | figures 1,2+ textbook | 3,4 |
| Q2 | Define oxidation number | textbook | 5 |
| Q3 | Define oxidizing agent and reducing agent in a redox reaction | table 2 | 6 |

# Oxidation and reduction

## Oxidation

**Loss** of electrons

**Gain** of oxygen

**Loss** of hydrogen

## Reduction

**Gain** of electrons

**Loss** of oxygen

**Gain** of hydrogen

Memory aid: **LEO the lion says GER**. **L**oss of **E**lectrons = **O**xidation, **G**ain of **E**lectrons = **R**eduction.

Oxidation and reduction are complementary: one cannot happen without the other.

# Electron transfer in REDOX Reactions

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## Figure 1: $Mg + O_2$

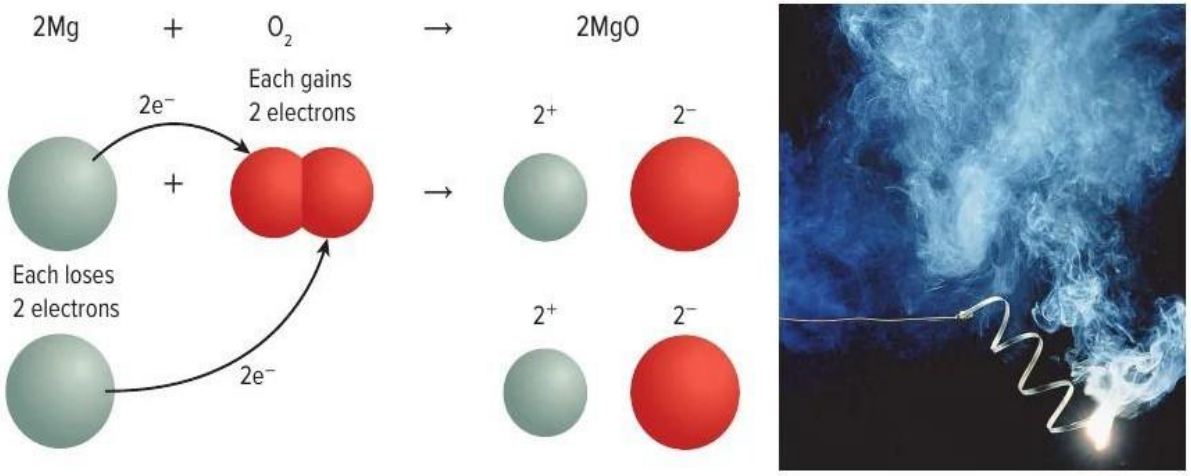


**Figure 1** The reaction of magnesium and oxygen involves a transfer of electrons from magnesium to oxygen. Therefore, this reaction is an oxidation-reduction reaction.

**Classify** *the reaction between magnesium and oxygen.*

## Figure 2: $Br^- + Cl_2$

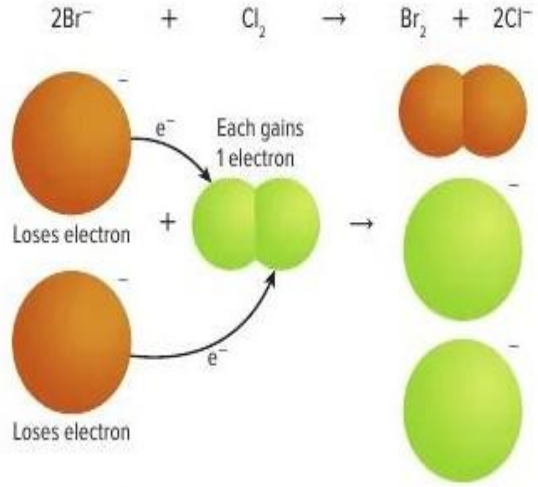


**Figure 2** The reaction between aqueous bromide ions and chlorine gas is a redox reaction. Here, electrons are transferred from bromide ions to chlorine.

**$2Mg + O_2 \rightarrow 2Mg^{2+} + 2O^{2-}$**

- Mg **loses** 2 $e^-$ → oxidized
- $O_2$ **gains** electrons → reduced

**$2Br^- + Cl_2 \rightarrow Br_2 + 2Cl^-$**

- $Br^-$ **loses** an electron → oxidized
- $Cl_2$ **gains** electrons → reduced

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# Oxidation number (n)

- **Oxidation number:** a number assigned to an atom or ion to show its **degree of oxidation or reduction**
- In an ionic compound it equals the number of electrons **lost** (positive) or **gained** (negative) by the atom, which is the ion charge
- Written with the **sign first** (+3, −2); ion charge is written with the sign **after** the number (3+, 2−)
- **Oxidation number increases → oxidation.**
- **Oxidation Number Decreases → reduction**

**Example: $2K + Cl_2 \rightarrow 2KCl$**

K: 0 → **+1** (loses 1 $e^-$)

Cl: 0 → **−1** (gains 1 $e^-$)

Potassium is oxidized; chlorine is reduced.

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Table 3 Rules for Determining Oxidation Numbers

| Rule | Example | $n_{element}$ |
|---|---|---|
| 1. The oxidation number of an atom of an uncombined element is zero. | Na, $O_2$, $Cl_2$, $H_2$ | 0 |
| 2. The oxidation number of a monatomic ion is equal to the charge of the ion. | $Ca^{2+}$ | +2 |
| | $Br^-$ | −1 |
| 3. The oxidation number of the more electronegative atom in a molecule or a complex ion is the same as the charge it would have if it were an ion. | N in $NH_3$ | −3 |
| | O in NO | −2 |
| 4. The oxidation number of the most electronegative element, fluorine, is always −1 when it is bonded to another element. | F in LiF | −1 |
| 5. The oxidation number of oxygen in compounds is always −2 except in peroxides, such as hydrogen peroxide ($H_2O_2$), where it is −1. When it is bonded to fluorine, the only element more electronegative than oxygen, the oxidation number of oxygen is positive. | O in $NO_2$ | −2 |
| | O in $H_2O_2$ | −1 |
| 6. The oxidation number of hydrogen in most of its compounds is +1, except in metal hydrides; then, the oxidation number is −1. | H in NaH | −1 |
| 7. The oxidation numbers of group 1 and 2 metals and aluminum are positive and equal to their number of valence electrons. | K | +1 |
| | Ca | +2 |
| | Al | +3 |
| 8. The sum of the oxidation numbers in a neutral compound is zero. | $CaBr_2$ | (+2) + 2(−1) = 0 |
| 9. The sum of the oxidation numbers of the atoms in a polyatomic ion is equal to the charge of the ion. | $SO_3^{2-}$ | (+4) + 3(−2) = −2 |

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| | Group number | Number of valance electrons | Charge of each group |
|---|---|---|---|
| Metals | 1 | 1 | +1 |
| | 2 | 2 | +2 |
| | 13 | 3 | +3 |
| Non-Metals | 15 | 5 | -3 |
| | 16 | 6 | -2 |
| | 17 | 7 | -1 |

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# Varied Oxidation Numbers of Oxygen (O) حالات تأكسد الأكسجين المختلفة (O)

| Free Element / Ozone | | State | العنصر الحر / الأوزون |
|---|---|---|---|
| $O_2$ Oxygen gas | O.N. = 0 | 0 | O.N. = 0 غاز الأكسجين |
| $O_3$ Ozone | | | الأوزون |

| Peroxides | | State | فوق الأكاسيد |
|---|---|---|---|
| $H_2O_2$ Hydrogen peroxide | O.N. = –1 | -1 | O.N. = –1 فوق أكسيد الهيدروجين |
| $Na_2O_2$ Sodium peroxide | | | عدد التأكسد = 1– فوق أكسيد الصوديوم |

| Fluorine Compounds | | | مع الفلور |
|---|---|---|---|
| $OF_2$ Oxygen difluoride | → State: +2 | +2 | عدد التأكسد = 2+ ثنائي فلوريد الأكسجين |
| $O_2F_2$ Dioxygen difluoride | → State: +1 | +1 | عدد التأكسد = 1+ ثنائي فلوريد ثنائي الأكسجين |

| Standard Oxides | | | الأكاسيد العادية |
|---|---|---|---|
| $H_2O$ Water | | -2 | O.N. = –2 الماء |
| $CO_2$ Carbon dioxide | | | عدد التأكسد = 2– ثاني أكسيد الكربون |

| Superoxides | | | السوبر أكسيد |
|---|---|---|---|
| $KO_2$ Potassium superoxide | | -½ | O.N. = عدد التأكسد = ½– سوبر أكسيد البوتاسيوم |

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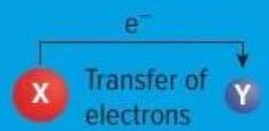
# Oxidizing agent and reducing agent

Table 2 Summary of Redox Reactions

| Process |   |
|---|---|
| **Oxidation**<br>• A reactant loses an electron.<br>• Reducing agent is oxidized.<br>• Oxidation number increases. | • X loses an electron.<br>• X is the reducing agent and becomes oxidized.<br>• The oxidation number of X increases. |
| **Reduction**<br>• Other reactant gains an electron.<br>• Oxidizing agent is reduced.<br>• Oxidation number decreases. | • Y gains an electron.<br>• Y is the oxidizing agent and becomes reduced.<br>• The oxidation number of Y decreases. |

## Oxidizing agent

**Accepts** electrons

Is itself **reduced**

Its oxidation number **decreases**

## Reducing agent

**Supplies** (donates) electrons

Is itself **oxidized**

Its oxidation number **increases**

Example: $2K + Cl_2 \rightarrow 2KCl$. K is the reducing agent; $Cl_2$ is the oxidizing agent.

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| | | | |
|---|---|---|---|
| Q4 | Identify oxidizing agent and reducing agent in a redox reaction | textbook +example problem 1 + practice problems | 5,6,8 |
| Q5 | Identify oxidizing agent and reducing agent in a redox reaction | textbook+ example problem1 + practice problems | 5,6,8 |
| Q6 | Distinguish between oxidation and reduction in terms of change in oxidation number | textbook +example problem1+ practice problems | 5,6,8 |

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## PRACTICE Problems

ADDITIONAL PRACTICE

1. Identify each of the following changes as either oxidation or reduction. Recall that $e^-$ is the symbol for an electron.
   - **a.** $I_2 + 2e^- \rightarrow 2I^-$
   - **b.** $K \rightarrow K^+ + e^-$
   - **c.** $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   - **d.** $Ag^+ + e^- \rightarrow Ag$
2. Identify what is oxidized and what is reduced in the following processes.
   - **a.** $2Br^- + Cl_2 \rightarrow Br_2 + 2Cl^-$
   - **b.** $2Ce + 3Cu^{2+} \rightarrow 3Cu + 2Ce^{3+}$
   - **c.** $2Zn + O_2 \rightarrow 2ZnO$
   - **d.** $2Na + 2H^+ \rightarrow 2Na^+ + H_2$
3. Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer.

   $$Fe(s) + 2Ag^+(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$$
4. **CHALLENGE** Identify the oxidizing agent and the reducing agent in each reaction.
   - **a.** $Mg + I_2 \rightarrow MgI_2$
   - **b.** $H_2S + Cl_2 \rightarrow S + 2HCl$

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# Example Problem 1: identify the agents

**$2Al + 2Fe^{3+} + 3O^{2-} \rightarrow 2Fe + 2Al^{3+} + 3O^{2-}$**

## Oxidation

$Al \rightarrow Al^{3+} + 3e^-$

Al **loses** electrons: it is **oxidized**

→ Al is the **reducing agent**

## Reduction

$Fe^{3+} + 3e^- \rightarrow Fe$

$Fe^{3+}$ **gains** electrons: it is **reduced**

→ $Fe^{3+}$ is the **oxidizing agent**

- $O^{2-}$ is a **spectator ion**: its charge does not change, so oxygen is not involved in the electron transfer.
- Steps: find what is oxidized and what is reduced, then name the agents.

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# Practice problems 3 and 4: agents

| Reaction | Oxidized (reducing agent) | Reduced (oxidizing agent) |
|---|---|---|
| 3. $Fe + 2Ag^+ \rightarrow Fe^{2+} + 2Ag$ | Fe (0 → +2) | $Ag^+$ (+1 → 0) |
| 4a. $Mg + I_2 \rightarrow MgI_2$ | Mg (0 → +2) | $I_2$ (0 → −1) |
| 4b. $H_2S + Cl_2 \rightarrow S + 2HCl$ | $H_2S$ (S: −2 → 0) | $Cl_2$ (0 → −1) |

Tip: the **reducing agent** is the substance that is **oxidized**; the **oxidizing agent** is the substance that is **reduced**.

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# Oxidation and reduction by oxidation number

Oxidation number **increases** → **oxidation**

Oxidation number **decreases** → **reduction**

**Practice problem 1** (electron change)

| Half-reaction | Change | Process |
|---|---|---|
| $I_2 + 2e^- \rightarrow 2I^-$ | 0 → −1 | Reduction |
| $K \rightarrow K^+ + e^-$ | 0 → +1 | Oxidation |
| $Fe^{2+} \rightarrow Fe^{3+} + e^-$ | +2 → +3 | Oxidation |
| $Ag^+ + e^- \rightarrow Ag$ | +1 → 0 | Reduction |

**Practice problem 2** (what is oxidized / reduced)

| Reaction | Oxidized | Reduced |
|---|---|---|
| $2Br^- + Cl_2 \rightarrow Br_2 + 2Cl^-$ | $Br^-$ | $Cl_2$ |
| $2Ce + 3Cu^{2+} \rightarrow 2Ce^{3+} + 3Cu$ | Ce | $Cu^{2+}$ |
| $2Zn + O_2 \rightarrow 2ZnO$ | Zn | $O_2$ |
| $2Na + 2H^+ \rightarrow 2Na^+ + H_2$ | Na | $H^+$ |

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## PRACTICE Problems

ADDITIONAL PRACTICE

1. Identify each of the following changes as either oxidation or reduction. Recall that $e^-$ is the symbol for an electron.
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   - **d.** $2Na + 2H^+ \rightarrow 2Na^+ + H_2$
3. Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer.

$$Fe(s) + 2Ag^+(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$$

4. **CHALLENGE** Identify the oxidizing agent and the reducing agent in each reaction.
   - **a.** $Mg + I_2 \rightarrow MgI_2$
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1. Identify each of the following changes as either oxidation or reduction. Recall that $e^-$ is the symbol for an electron.

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

   **reduction**

   b. $K \rightarrow K^+ + e^-$

   **oxidation**

   c. $Fe^{2+} \rightarrow Fe^{3+} + e^-$

   **oxidation**

   d. $Ag^+ + e^- \rightarrow Ag$

   **reduction**

2. Identify what is oxidized and what is reduced in the following processes.

   a. $2Br^- + Cl_2 \rightarrow Br_2 + 2Cl^-$

   **$Br^-$ is oxidized, Cl is reduced**

   b. $2Ce + 3Cu^{2+} \rightarrow 3Cu + 2Ce^{3+}$

   **Ce is oxidized, $Cu^{2+}$ is reduced**

   c. $2Zn + O_2 \rightarrow 2ZnO$

   **Zn is oxidized, $O_2$ is reduced**

   d. $2Na + 2H^+ \rightarrow 2Na^+ + H_2$

   **$H^+$ is reduced, Na is oxidized**

3. Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer.

   $Fe(s) + 2Ag^+(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$

   **$Ag^+$ is the oxidizing agent, Fe is the reducing agent; $Ag^+$ is reduced, Fe is oxidized**

4. **Challenge** Identify the oxidizing agent and the reducing agent in each reaction.

   a. $Mg + I_2 \rightarrow MgI_2$

   **$I_2$ is the oxidizing agent, Mg is the reducing agent**

   b. $H_2S + Cl_2 \rightarrow S + 2HCl$

   **$Cl_2$ is the oxidizing agent, $H_2S$ is the reducing agent**

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| | | | |
|---|---|---|---|
| Q7 | Assign oxidation number to atoms, ions and compounds according to a set of rules | textbook +example problem 2 + practice problems+table3 | 9,10 |
| Q8 | Assign oxidation number to atoms, ions and compounds according to a set of rules | textbook+ example problem2 + practice problems+table3 | 9,10 |
| Q9 | Determines the oxidation number of elements and compounds with more complex formulas. | textbook+ example problem2+ practice problems+table3 | 9,10 |

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## Table 3 Rules for Determining Oxidation Numbers

| Rule | Example | $n_{element}$ |
|---|---|---|
| 8. The sum of the oxidation numbers in a neutral compound is zero. | $CaBr_2$ | $(+2) + 2(-1) = 0$ |
| 9. The sum of the oxidation numbers of the atoms in a polyatomic ion is equal to the charge of the ion. | $SO_3^{2-}$ | $(+4) + 3(-2) = -2$ |

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# Varied Oxidation Numbers of Oxygen (O) حالات تأكسد الأكسجين المختلفة (O)

| Free Element / Ozone | | State | العنصر الحر / الأوزون |
|---|---|---|---|
| $O_2$ Oxygen gas | O.N. = 0 | 0 | O.N. = 0 غاز الأكسجين |
| $O_3$ Ozone | | | الأوزون |

| Peroxides | | State | فوق الأكاسيد |
|---|---|---|---|
| $H_2O_2$ Hydrogen peroxide | O.N. = –1 | -1 | O.N. = –1 فوق أكسيد الهيدروجين |
| $Na_2O_2$ Sodium peroxide | | | عدد التأكسد = 1– فوق أكسيد الصوديوم |

| Fluorine Compounds | | | مع الفلور |
|---|---|---|---|
| $OF_2$ Oxygen difluoride | → State: +2 | +2 | عدد التأكسد = 2+ ثنائي فلوريد الأكسجين |
| $O_2F_2$ Dioxygen difluoride | → State: +1 | +1 | عدد التأكسد = 1+ ثنائي فلوريد ثنائي الأكسجين |

| Standard Oxides | | | الأكاسيد العادية |
|---|---|---|---|
| $H_2O$ Water | | -2 | O.N. = –2 الماء |
| $CO_2$ Carbon dioxide | | | عدد التأكسد = 2– ثاني أكسيد الكربون |

| Superoxides | | | السوبر أكسيد |
|---|---|---|---|
| $KO_2$ Potassium superoxide | | -½ | O.N. = عدد التأكسد = ½– سوبر أكسيد البوتاسيوم |

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## EXAMPLE Problem 2

**DETERMINE OXIDATION NUMBERS** Use the rules for determining oxidation numbers to find the oxidation number of each element in potassium chlorate ($KClO_3$) and in a sulfite ion ($SO_3^{2-}$).

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## PRACTICE Problems

ADDITIONAL PRACTICE

5. Determine the oxidation number of the boldface element in the following formulas for compounds.
   a. $NaClO_4$ b. $AlPO_4$ c. $HNO_2$
6. Determine the oxidation number of the boldface element in the following formulas for ions.
   a. $NH_4^+$ b. $AsO_4^{3-}$ c. $CrO_4^{2-}$
7. Determine the oxidation number of nitrogen in each of these molecules.
   a. $NH_3$ b. KCN c. $N_2H_4$
8. **CHALLENGE** Determine the net change of oxidation number of each of the elements in these redox equations.
   a. $C + O_2 \rightarrow CO_2$
   b. $Cl_2 + ZnI_2 \rightarrow ZnCl_2 + I_2$
   c. $CdO + CO \rightarrow Cd + CO_2$

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5. Determine the oxidation number of the boldface element in the following formulas for compounds.

   a. $NaClO_4$

   **+7**

   b. $AlPO_4$

   **+5**

   c. $HNO_2$

   **+3**

6. Determine the oxidation number of the boldface element in the following formulas for ions.

   a. $NH_4^+$

   **–3**

   b. $AsO_4^{3-}$

   **+5**

   c. $CrO_4^{2-}$

   **+6**

7. Determine the oxidation number of nitrogen in each of these molecules or ions.

   a. $NH_3$

   **–3**

   b. KCN

   **–3**

   c. $N_2H_4$

   **–2**

8. **Challenge** Determine the net change of oxidation number of each of the elements in these redox equations.

   a. $C + O_2 \rightarrow CO_2$

   **C, +4; O, –2**

   b. $Cl_2 + ZnI_2 \rightarrow ZnCl_2 + I_2$

   **I, +1; Cl, –1; Zn, no change**

   c. $CdO + CO \rightarrow Cd + CO_2$

   **C, +2; Cd, –2; O, no change**

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| | | | |
|---|---|---|---|
| Q10 | Distinguish between oxidation and reduction in terms of change in oxidation number | check your progress and textbook | 11 |
| Q11 | Distinguish between oxidation and reduction in terms of change in oxidation number | check your progress and textbook | 11 |

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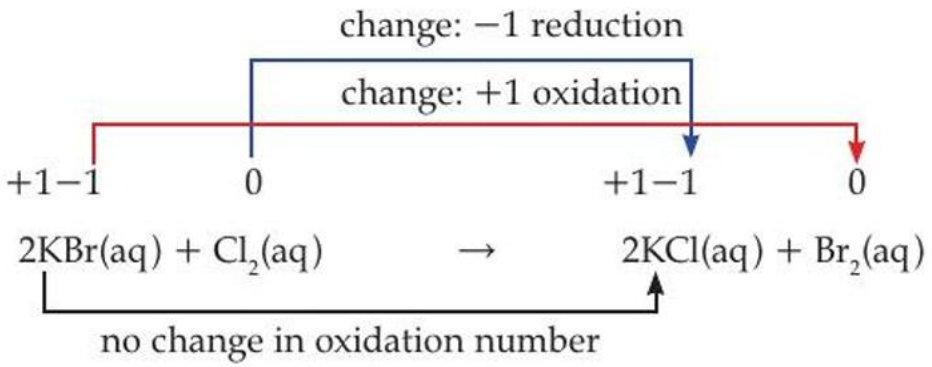
# Oxidation Numbers in Redox Reactions

Recall the replacement reaction of bromine in aqueous potassium bromide (KBr) by chlorine ($Cl_2$). Assign numbers, using **Table 1**, to all elements in the balanced equation. Then review the changes, as shown in the equation below. The oxidation number of bromine changed from −1 to 0, an increase of 1 while the oxidation number of chlorine changed from 0 to −1, a decrease of 1. Therefore, chlorine is reduced, and bromine is oxidized.

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**4. CHALLENGE** Identify the oxidizing agent and the reducing agent in each reaction.

**a.** $Mg + I_2 \rightarrow MgI_2$

**b.** $H_2S + Cl_2 \rightarrow S + 2HCl$

**4. Challenge** Identify the oxidizing agent and the reducing agent in each reaction.

**a.** $Mg + I_2 \rightarrow MgI_2$

**$I_2$ is the oxidizing agent, Mg is the reducing agent**

**b.** $H_2S + Cl_2 \rightarrow S + 2HCl$

**$Cl_2$ is the oxidizing agent, $H_2S$ is the reducing agent**

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# Check Your Progress

## Summary

- Oxidation-reduction reactions involve the transfer of electrons from one atom to another.
- When an atom or ion is reduced, its oxidation number decreases. When an atom or ion is oxidized, its oxidation number increases.
- In oxidation-reduction reactions involving molecular compounds (and polyatomic ions with covalent bonds), the more-electronegative atoms are treated as if they are reduced. The less-electronegative atoms are treated as if they are oxidized.

## Demonstrate Understanding

9. **Explain** why oxidation and reduction must always occur together.
10. **Describe** the roles of oxidizing agents and reducing agents in a redox reaction. How is each changed in the reaction?
11. **Write** the equation for the reaction of iron metal with hydrobromic acid to form aqueous iron(III) bromide and hydrogen gas. Determine the change in oxidation number for the element that is reduced and the element that is oxidized.
12. **Determine** the oxidation number of the boldface element in these compounds.
    a. $H\mathbf{N}O_3$
    b. $Ca_3\mathbf{N}_2$
    c. $\mathbf{Sb}_2O_5$
    d. $Cu\mathbf{W}O_4$
13. **Determine** the oxidation number of the boldface element in these ions.
    a. $\mathbf{I}O_4^-$
    b. $\mathbf{Mn}O_4^-$
    c. $\mathbf{B}_4O_7^{2-}$
    d. $\mathbf{N}H_2^-$
14. **Make and Use Graphs** Alkali metals are strong reducing agents. Make a graph showing how the reducing abilities of the alkali metals increase or decrease as you move down the family from sodium to francium.

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# Check Your Progress

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## Lesson 1 Check Your Progress

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9. **Explain** why oxidation and reduction must always occur together.

   **If an atom or ion loses an electron, some other species must gain the electron.**

10. **Describe** the roles of oxidizing agents and reducing agents in a redox reaction. How is each changed in the reaction?

    **An oxidizing agent causes another species to be oxidized by gaining the electrons from it. A reducing agent causes another species to be reduced by losing electrons to that element.**

11. **Write** the equation for the reaction of iron metal with hydrobromic acid to form aqueous iron(III) bromide and hydrogen gas. Determine the change in oxidation number for the element that is reduced and for the element that is oxidized.

    **$2Fe(s) + 6HBr\ (aq) \rightarrow 2FeBr_3\ (aq) + 3H_2(g)$; Fe is oxidized (from 0 to +3), and H is reduced (from +1 to 0).**

12. **Determine** the oxidation number of the boldface element in these compounds.

    a. $HNO_3$

    **+5**

    b. $Ca_3N_2$

    **–3**

    c. $Sb_2O_5$

    **+5**

    d. $CuWO_4$

    **+6**

13. **Determine** the oxidation number of the boldface element in these ions.

    a. $IO_4^-$

    **+7**

    b. $MnO_4^-$

    **+7**

    c. $B_4O_7^{2-}$

    **+3**

    d. $NH_2^-$

    **–3**

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# Previous Exam Practice Problems
# EOT1 Chemistry **Part 1**
# **Q1 to Q11**

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| Which of the following describes oxidation correctly ?<br>أي مما يلي يصف الأكسدة بشكل صحيح؟ |
|---|
| A. A reactant loses electrons, and its oxidation number increases<br>يفقد المتفاعل إلكترونات ويزداد عدد تأكسده |
| B. A reactant loses electrons, and its oxidation number decreases<br>يفقد المتفاعل إلكترونات ويقل عدد تأكسده |
| C. A reactant gains electrons, and its oxidation number inceases<br>يكتسب المتفاعل إلكترونات ويزداد عدد تأكسده |
| D. A reactant gains electrons, and its oxidation number decreases<br>يكتسب المتفاعل إلكترونات ويقل عدد تأكسده |

| Which of the following is correct in relation to the reaction below?<br>أي مما يأتي صحيح فيما يتعلق بالتفاعل أدناه؟<br>$2K_{(s)} + F_{2(g)} \longrightarrow 2KF_{(s)}$ |
|---|
| A. The potassium atom **K** gained an electron, and it is reduced<br>اكتسبت ذرة البوتاسيوم **K** إلكترونا وحدث لها اختزال |
| B. Fluorine $F_2$ lost electrons and was oxidized<br>فقد الفلور $F_2$ إلكترونات وحدث لها أكسدة |
| C. The potassium atom **K** lost an electron, and it was oxidized<br>فقدت ذرة البوتاسيوم **K** إلكترونا وحدث لها أكسدة |
| D. Fluorine $F_2$ gained electrons and was oxidized<br>اكتسبت الفلور $F_2$ إلكترونات وحدث لها أكسدة |

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What is the following is **Not true** about **Y** in the figure shown below?

أي مما يلي **غير صحيح** حول **Y** في الشكل أدناه؟




| | |
|---|---|
| A. **Y** is the reducing agent and becomes oxidized | **Y** عامل مختزل وبتأكسد |
| B. The oxidation number of **Y** is decrease | عدد تأكسد **Y** يقل |
| C. **Y** gains electron | **Y** يكتسب إلكترون |
| D. **Y** is the oxidizing agent and becomes reduced | **Y** عامل مؤكسد ويختزل |

What is the following is **Not true** about **X** in the figure shown below?

أي مما يلي **غير صحيح** حول **X** في الشكل أدناه؟




| | |
|---|---|
| A. **X** is the reducing agent and becomes oxidized | **X** عامل مختزل وبتأكسد |
| B. The oxidation number of **X** is increase | عدد تأكسد **X** يزداد |
| C. **X** loses electron | **X** يفقد إلكترون |
| D. **X** is the oxidizing agent and becomes reduced | **X** عامل مؤكسد ويختزل |

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| |
|---|
| Regarding the reaction below .which of the following is **correct** ?<br>فيما يتعلق بالتفاعل أدناه. أي مما يأتي **صحيح** ؟<br>$2KBr(aq) + Cl_2(aq) \rightarrow 2KCl(aq) + Br_2(aq)$ |
| A. The oxidation number of bromine changed from -1 to 0<br>تغير عدد تأكسد البروم من -1 إلى 0 |
| B. The oxidation number of bromine changed from 0 to -1<br>تغير عدد تأكسد البروم من 0 إلى -1 |
| C. The oxidation number of Chlorine changed from +1 to 0<br>تغير عدد تأكسد الكلور من +1 إلى 0 |
| D. The oxidation number of Chlorine changed from -2 to 0<br>تغير عدد تأكسد الكلور من -2 إلى 0 |

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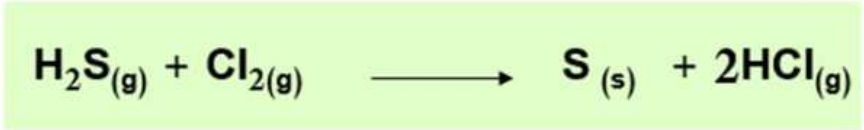
Regarding the reaction below .which of the following is **correct** ?

فيما يتعلق بالتفاعل أدناه. أي مما يأتي **صحيح** ؟



$$H_2S_{(g)} + Cl_{2(g)} \longrightarrow S_{(s)} + 2HCl_{(g)}$$

A. The oxidation number of Chlorine changed from -1 to 0

تغير عدد تأكسد الكلور من -1 إلى 0

B. The oxidation number of Chlorine changed from -2 to 0

تغير عدد تأكسد الكلور من -2 إلى 0

C. The oxidation number of Sulfur changed from -1 to 0

تغير عدد تأكسد الكبريت من -1 إلى 0

D. The oxidation number of Sulfur changed from -2 to 0

تغير عدد تأكسد الكبريت من -2 إلى 0

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Which of the following is correct in relation to the reaction below?

أي مما يأتي صحيح فيما يتعلق بالتفاعل أدناه؟

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$$

- ○ Oxygen $O_2$ gained electrons and was oxidized — اكتسب الأكسجين $O_2$ إلكترونات وحدث له أكسدة
- ○ The magnesium **Mg** atom lost electrons and was oxidized — فقدت ذرة المغنيسيوم **Mg** إلكترونات وحدث لها أكسدة
- ○ Oxygen $O_2$ lost electrons and was oxidized — فقد الأكسجين $O_2$ إلكترونات وحدث له أكسدة
- ○ The magnesium **Mg** atom gained electrons and it is Reduced — اكتسبت ذرة المغنيسيوم **Mg** إلكترونات وحدث لها اختزال

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In the reaction below, if you know that the reactant Fe is the reducing agent. Which if the following describes it **correctly** ?

في التفاعل ادناه ، إذا علمت أن المتفاعل Fe هو العامل المختزل. أي مما يأتي يصفه بشكل **صحيح** ؟

$$Fe(s) + 2Ag^{+}(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$$

| | |
|---|---|
| يكتسب إلكترونات – يزيد عدد تأكسده – تحدث له أكسدة<br>Gains electrons – its oxidation number increases – it is oxidized | 1 |
| يفقد إلكترونات – يزيد عدد تأكسده – تحدث له أكسدة<br>loses electrons – its oxidation number increases – it is oxidized | 2 |
| يكتسب إلكترونات – يقل عدد تأكسده – يحدث له اختزال<br>Gains electrons – its oxidation number decreases – it is reduced | 3 |
| يفقد إلكترونات – يقل عدد تأكسده – يحدث له اختزال<br>loses electrons – its oxidation number decreases – it is reduced | 4 |

A. 1

B. 2

C. 3

D. 4

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| What is the **oxidizing agent** in the following reaction ? ما **المعامل المؤكسد** في التفاعل التالي ؟ |
|---|
| $H_2S_{(g)} + Cl_{2(g)} \longrightarrow S_{(s)} + 2HCl_{(g)}$ |
| A. $H_2S$ |
| B. $Cl_2$ |
| C. S |
| D. HCl |

| What is the **reducing agent** in the following reaction ? ما **المعامل المختزل** في التفاعل التالي ؟ |
|---|
| $H_2S_{(g)} + Cl_{2(g)} \longrightarrow S_{(s)} + 2HCl_{(g)}$ |
| A. $H_2S$ |
| B. $Cl_2$ |
| C. S |
| D. HCl |

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| Which of the following has the highest oxidation number of Nitrogen (**N**) ? | أي مما لديه أعلى عدد تأكسد للنيتروجين (N)؟ |
|---|---|
| A.H**N**$O_3$ | |
| B.**N**$O_2$ | |
| C.**N**$H_3$ | |
| D.H**N**$O_2$ | |

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What is the correct ascending order of the following formulas according to the oxidation number of **chlorine** in each of them?

ما الترتيب التصاعدي الصحيح للصيغ التالية حسب عدد تأكسد **الكلور** في كل منها؟

$Cl_2$ – $NaCl$ – $KClO_4$ – $KClO$

$NaCl \longrightarrow Cl_2 \longrightarrow KClO \longrightarrow KClO_4$

$Cl_2 \longrightarrow NaCl \longrightarrow KClO_4 \longrightarrow KClO$

$KClO_4 \longrightarrow NaCl \longrightarrow Cl_2 \longrightarrow KClO$

$KClO_4 \longrightarrow KClO \longrightarrow Cl_2 \longrightarrow NaCl$

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In which of the following formulas is the oxidation number of manganese **greatest**?

في أي الصيغ التالية، يكون عدد تأكسد المنجنيز **أكبر ما يُمكن**؟

| | | | | | | |
|---|---|---|---|---|---|---|
| $KMnO_4$ | – | $MnO_2$ | – | $MnCl_2$ | – | $K_2MnO_4$ |

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In the redox reaction below. **What represents the oxidation half-reaction and the reduction half-reaction?**

في تفاعل الأكسدة والاختزال أدناه. **ما الذي يُمثل نصفي تفاعل الأكسدة والاختزال؟**

$$2Br^- + Cl_2 \rightarrow Br_2 + 2Cl^-$$

| نصف تفاعل الاختزال<br>Reduction half-reaction | نصف تفاعل الأكسدة<br>Oxidation Half-Reaction | |
|---|---|---|
| $2Br^- + 2e^- \longrightarrow Br_2$ | $Cl_2 \longrightarrow 2Cl^- + 2e^-$ | 1 |
| $Cl_2 \longrightarrow 2Cl^- + 2e^-$ | $2Br^- + 2e^- \longrightarrow Br_2$ | 2 |
| $2Br^- \longrightarrow Br_2 + 2e^-$ | $Cl_2 + 2e^- \longrightarrow 2Cl^-$ | 3 |
| $Cl_2 + 2e^- \longrightarrow 2Cl^-$ | $2Br^- \longrightarrow Br_2 + 2e^-$ | 4 |

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What is the correct **ascending order** of the following formulas according to the oxidation number of hydrogen in each of them?

ما الترتيب التصاعدي الصحيح للصيغ التالية تبعا لعدد تأكسد الهيدروجين في كل منها ؟

$H_2$ – HCl – LiH

A. The smallest is $H_2$ → LiH → HCl ( the greatest)

الأصغر هو $H_2$ ← LiH ← HCl ( الأكبر )

B. The smallest is $H_2$ → HCl → LiH ( the greatest)

الأصغر هو $H_2$ ← HCl ← LiH ( الأكبر )

C. The smallest is LiH → HCl → $H_2$ ( the greatest)

الأصغر هو LiH ← HCl ← $H_2$ ( الأكبر )

D. The smallest is LiH → $H_2$ → HCl ( the greatest)

الأصغر هو LiH ← $H_2$ ← HCl ( الأكبر )

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What is the correct **ascending** order of the following formulas according to the oxidation number of oxygen in each of them?

ما الترتيب التصاعدي الصحيح للصيغ التالية تبعا لعدد تأكسد الأكسجين في كل منها ؟

$CO_2$ - $O_2$ - $H_2O_2$

A. The smallest is $H_2O_2 \rightarrow CO_2 \rightarrow O_2$ ( the greatest)

الأصغر هو $H_2O_2 \leftarrow CO_2 \leftarrow O_2$ ( الأكبر )

B. The smallest is $O_2 \rightarrow CO_2 \rightarrow H_2O_2$ ( the greatest)

الأصغر هو $O_2 \leftarrow CO_2 \leftarrow H_2O_2$ ( الأكبر )

C. The smallest is $CO_2 \rightarrow H_2O_2 \rightarrow O_2$ ( the greatest)

الأصغر هو $CO_2 \leftarrow H_2O_2 \leftarrow O_2$ ( الأكبر )

D. The smallest is $H_2O_2 \rightarrow O_2 \rightarrow CO_2$ ( the greatest)

الأصغر هو $H_2O_2 \leftarrow O_2 \leftarrow CO_2$ ( الأكبر )

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What is the correct **descending** order of the following formulas according to the oxidation number of **oxygen** in each of them?

ما الترتيب **التنازلي** الصحيح للصيغ التالية حسب عدد تأكسد **الأكسجين** في كل منها؟

$H_2O_2$ – $H_2O$ – $O_2$ – $OF_2$

- ○ $H_2O \longrightarrow OF_2 \longrightarrow H_2O_2 \longrightarrow O_2$
- ○ $O_2 \longrightarrow H_2O_2 \longrightarrow OF_2 \longrightarrow H_2O$
- ○ $OF_2 \longrightarrow O_2 \longrightarrow H_2O_2 \longrightarrow H_2O$
- ○ $H_2O \longrightarrow H_2O_2 \longrightarrow O_2 \longrightarrow OF_2$

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What is the **oxidizing** agent in the following reaction?

ما العامل **المؤكسد** في التفاعل التالي؟

$$Na_2SO_{4\,(s)} + 4C_{(s)} \longrightarrow Na_2S_{(s)} + 4CO_{(g)}$$

C

CO

$Na_2SO_4$

$Na_2S$

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When is the oxidation number of oxygen positive?

متى يكون عدد تأكسد الأكسجين موجبًا؟

a. When it is bonded to itself — عندما يرتبط مع نفسه

b. In peroxides compounds — في مركبات فوق الأكاسيد

c. In oxides compounds — في مركبات الأكاسيد

d. When it is bonded to fluorine — عندما يرتبط مع الفلور

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# End Of Part 1 ☺

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# EoT1 Chemistry Coverage
# Grade 12 Adv
# Part-2 (Q12-Q30)
# 2026-2027

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| | | | |
|---|---|---|---|
| Q12 | Describe the steps for balancing redox reactions, in acidic medium, by the half-reaction method | table 6 | 17 |
| Q13 | Write oxidation-half reaction and reduction-half reaction for a redox reaction | figure 8+ textbook+ table 5 | 16,17 |

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# Writing the two half-reactions

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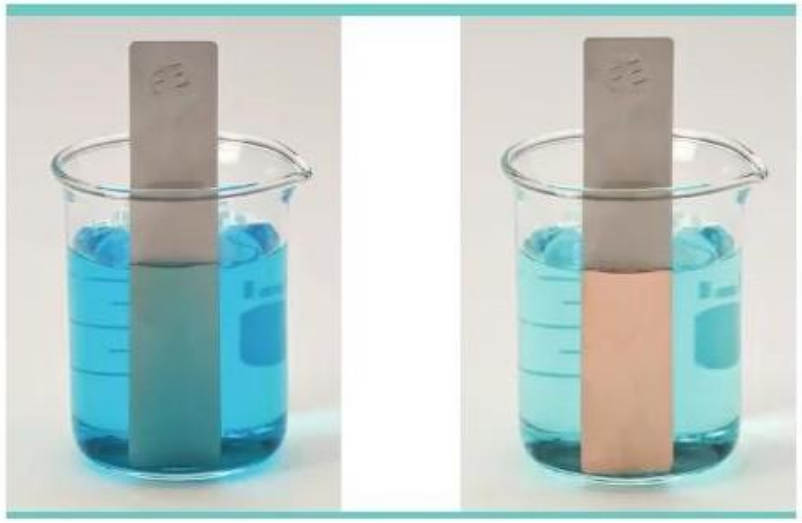

**Figure 8** As a result of this redox reaction between iron and copper(II) sulfate solution, solid copper metal is deposited on the iron. To balance the chemical equation for this reaction, you could use half-reactions.

**Figure 8:** $Fe + CuSO_4 \rightarrow Cu + Fe_2(SO_4)_3$

**Net ionic equation: $Fe + Cu^{2+} \rightarrow Cu + Fe^{3+}$**

**$Fe \rightarrow Fe^{3+} + 3e^-$ (oxidation)**

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

# Writing the two half-reactions

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**Table 5:** iron is always oxidized as $Fe \rightarrow Fe^{3+} + 3e^-$; the reduction half-reaction depends on what reacts with iron.

Table 5 Redox Reactions that Oxidize Iron

| Overall Reaction (Unbalanced) | Oxidation Half-Reaction | Reduction Half-Reaction |
|---|---|---|
| $Fe + O_2 \rightarrow Fe_2O_3$ | $Fe \rightarrow Fe^{3+} + 3e^-$ | $O_2 + 4e^- \rightarrow 2O^{2-}$ |
| $Fe + F_2 \rightarrow FeF_3$ | | $F_2 + 2e^- \rightarrow 2F^-$ |
| $Fe + HBr \rightarrow H_2 + FeBr_3$ | | $2H^+ + 2e^- \rightarrow H_2$ |
| $Fe + AgNO_3 \rightarrow Ag + Fe(NO_3)_3$ | | $Ag^+ + e^- \rightarrow Ag$ |
| $Fe + CuSO_4 \rightarrow Cu + Fe_2(SO_4)_3$ | | $Cu^{2+} + 2e^- \rightarrow Cu$ |

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Table 6 The Half-Reaction Method

1. Write the unbalanced, net ionic equation for the reaction, omitting spectator ions.

$$Fe + Cu^{2+} + SO_4^{2-} \rightarrow Cu + 2Fe^{3+} + 3SO_4^{2-}$$

$$Fe + Cu^{2+} \rightarrow Cu + 2Fe^{3+}$$

2. Write separate, incomplete equations for the oxidation and reduction half-reactions, including oxidation numbers.

$$\overset{0}{Fe} \rightarrow \overset{+3}{2Fe^{3+}} \text{ (oxidation)} \qquad \overset{+2}{Cu^{2+}} \rightarrow \overset{0}{Cu} \text{ (reduction)}$$

3. Balance the atoms in the half-reactions. Balance the charges in each half-reaction by adding electrons as reactants or products.

$$2Fe \rightarrow 2Fe^{3+} + 6e^- \qquad Cu^{2+} + 2e^- \rightarrow Cu$$

4. Adjust the coefficients so that the number of electrons lost in oxidation equals the number of electrons gained in reduction.

$$2Fe \rightarrow 2Fe^{3+} + 6e^- \qquad 3Cu^{2+} + 6e^- \rightarrow 3Cu$$

5. Add the half-reactions and cancel or reduce like terms on both sides of the equation.

$$2Fe + 3Cu^{2+} \rightarrow 3Cu + 2Fe^{3+}$$

6. Return spectator ions, if desired. Restore state descriptions.

$$2Fe(s) + 3CuSO_4(aq) \rightarrow 3Cu(s) + Fe_2(SO_4)_3(aq)$$

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| | | | |
|---|---|---|---|
| Q14 | Balance redox reaction in acidic medium using half-reaction method | textbook +example problem 5 + practice problems and check your progress | 16,17,18, 19 |
| Q15 | Balance redox reaction in acidic medium using half-reaction method | textbook +example problem 5 + practice problems | 16,17,18 |

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# خطوات وزن تفاعلات الأكسدة والاختزال في الوسط الحمضي
# Steps for Balancing Redox Reactions in Acidic Solution

| | | |
|---|---|---|
| 1. Separate into half-reactions |  | 1. تقسيم التفاعل إلى نصفين<br>تجزئة التفاعل إلى أكسدة واختزال |
| 2. Balance all atoms except H and O |  | 2. وزن جميع الذرات عدا الهيدروجين والأكسجين<br>موازنة الذرات غير الهيدروجين والأكسجين في الطرفين |
| 3. Balance Oxygen atoms by adding $H_2O$ to the O-deficient side |  | 3. وزن ذرات الأكسجين<br>إضافة جزيئات الماء $H_2O$ للطرف الناقص في الأكسجين |
| 4. Balance Hydrogen atoms by adding $H^+$ to the H-deficient side |  | 4. وزن ذرات الهيدروجين<br>إضافة أيونات الهيدروجين $H^+$ للطرف الناقص في الهيدروجين |
| 5. Balance charges with electrons ($e^-$) on the more positive side |  | 5. وزن الشحنات باستخدام الإلكترونات ($e^-$)<br>إضافة الإلكترونات لجعل الشحنات متساوية في الطرفين |
| 6. Equalize $e^-$ and combine half-reactions (cancel species on both sides) |  | 6. مساواة الإلكترونات وجمع نصفين<br>ضرب أنصاف التفاعل لتساوي الإلكترونات ثم جمعها واختصارها |

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## PRACTICE Problems

ADDITIONAL PRACTICE

**Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.**

**23.** $Cr_2O_7^{2-}(aq) + I^-(aq) \rightarrow Cr^{3+}(aq) + I_2(s)$ (in acidic solution)

**24.** $Mn^{2+}(aq) + BiO_3^-(aq) \rightarrow MnO_4^-(aq) + Bi^{2+}(aq)$ (in acidic solution)

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## Practice Problems

*page 615*

**23.** $Cr_2O_7^{2-}(aq) + I^-(aq) \rightarrow Cr^{3+}(aq) + I_2(s)$ (in acid solution)

**$2I^-(aq) \rightarrow I_2(s) + 2e^-$(oxidation)**

**$14H^+(aq) + 6e^- + Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$ (reduction)**

**Multiply oxidation half-reaction by 3 and add to reduction half-reaction**

**$14H^+(aq) + 6e^- + Cr_2O_7^{2-}(aq) + 6I^-(aq) \rightarrow 3I_2(s) + 2Cr^{3+}(aq) + 7H_2O(l) + 6e^-$**

**$14H^+(aq) + Cr_2O_7^{2-}(aq) + 6I^-(aq) \rightarrow 3I_2(s) + 2Cr^{3+}(aq) + 7H_2O(l)$**

**24.** $Mn^{2+}(aq) + BiO_3^-(aq) \rightarrow MnO_4(aq) + Bi^{2+}(aq)$ (in acid solution)

**$Mn^{2+}(aq) + 4H_2O(l) \rightarrow MnO_4^-(aq) + 5e^- + 8H^+(aq)$ (oxidation)**

**$BiO_3^-(aq) + 3e^- + 6H^+(aq) \rightarrow Bi^{2+}(aq) + 3H_2O(l)$ (reduction)**

**Multiply oxidation half-reaction by 3. Multiply reduction half-reaction by 5 and add to oxidation half-reaction.**

**$3Mn^{2+}(aq) + 12H_2O(l) + 5BiO_3^-(aq) + 15e^- + 30H^+(aq) \rightarrow 3MnO_4^-(aq) + 15e^- + 24H^+(aq) + 5Bi^{2+}(aq) + 15H_2O(l)$**

**$3Mn^{2+}(aq) + 5BiO_3^-(aq) + 6H^+(aq) \rightarrow 3MnO_4^-(aq) + 5Bi^{2+}(aq) + 3H_2O(l)$**

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| | | | |
|---|---|---|---|
| Q16 | Describe the steps for balancing redox reactions, in basic medium, by the half-reaction method | textbook +example problem 5 + practice problems and check your progress | 16,17,18, 19 |
| Q17 | Balance redox reaction in basic medium using half-reaction method | textbook +example problem 5 + practice problems and check your progress | 16,17,18, 19 |

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# خطوات وزن تفاعلات الأكسدة والاختزال في الوسط القاعدي
# Steps for Balancing Redox Reactions in Basic Solution

Steps (based on acidic steps with 2 final crucial basic steps)

| | | |
|---|---|---|
| 1. **Follow all steps for Acidic Solutions**<br>Balance atoms, O (using $H_2O$), H (using $H^+$), and charge (using $e^-$). | $O_2$ | 1. **اتبع خطوات الوسط الحمضي بالكامل**<br>موازنة الذرات، الأكسجين (بالماء $H_2O$)، الهيدروجين (أيونات $H^+$)، والشحنات (بالإلكترونات $e^-$) |
| 2. **Add $OH^-$ to both sides equal to $H^+$**<br>Check the count of $H^+$ and add that exact number of hydroxide ions ($OH^-$) to both sides. | $OH^-$ | 2. **إضافة أيونات $OH^-$ للطرفين مساوية لأيونات $H^+$**<br>انظر لعدد أيونات $H^+$ وأضف نفس العدد من الهيدروكسيد $OH^-$ لطرفي المعادلة |
| 3. **Combine $H^+$ and $OH^-$ to form $H_2O$**<br>Whenever $^+H$ meets $^-OH$ on the same side, combine them into water ($H_2O$). | $OH$ | 3. **دمج أيونات $H^+$ مع $OH^-$ لتكوين جزيئات ماء $H_2O$**<br>حيثما يلتقي $H^+$ مع $OH^-$ في نفس الطرف، يدمجان معاً لتكوين ماء $H_2O$ |
| 4. **Simplify and cancel water ($H_2O$)**<br>Subtract overlapping $H_2O$ molecules that appear on both sides of the equation. | | 4. **اختصار وتبسيط جزيئات الماء $H_2O$**<br>حذف جزيئات الماء $H_2O$ المتكررة في الطرفين للحصول على المعادلة النهائية |
| 5. **Verify balance of atoms and charge**<br>Double-check that both sides have equal atoms and the same total charge. | | 5. **التحقق النهائي من وزن الذرات والشحنات**<br>التأكد من تساوي الذرات والشحنات الكلية في الطرفين |

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## PRACTICE Problems

ADDITIONAL PRACTICE

**Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.**

**25. CHALLENGE** $N_2O(g) + ClO^-(aq) \rightarrow NO_2^-(aq) + Cl^-(aq)$ (in basic solution)

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25. **Challenge** $N_2O(g) + ClO^-(aq) \rightarrow NO_2^-(aq) + Cl^-(aq)$ (in basic solution).

**$6OH^-(aq) + N_2O(g) \rightarrow 2NO_2^-(aq) + 4e^- + 3H_2O(l)$(oxidation)**

**$ClO^-(aq) + 2e^- + H_2O(l) \rightarrow Cl^-(aq) + 2OH^-(aq)$(reduction)**

**Multiply reduction half-reaction by 2 and add to oxidation half-reaction.**

**$6OH^-(aq) + N_2O(g) + 2ClO^-(aq) + 4e^- + 2H_2O(l) \rightarrow 2NO_2^-(aq) + 4e^- + 3H_2O(l) + 2Cl^-(aq) + 4OH^-(aq)$**

**$N_2O(g) + 2ClO^-(aq) + 2OH^-(aq) \rightarrow 2NO_2^-(aq) + 2Cl^-(aq) + H_2O(l)$**

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# Check Your Progress

## Summary

- The oxidation-number method is based on the number of electrons transferred from atoms or ions equaling the number of electrons accepted by other atoms or ions.
- To balance equations for reactions in an acidic (or basic) solution, add hydrogen (or hydroxide) ions and water molecules.
- A half-reaction is one of the two parts of a redox reaction.

## Demonstrate Understanding

26. **Characterize** redox reactions by explaining how changes in oxidation numbers are related to the electrons transferred in a redox reaction. How are the changes related to the processes of oxidation and reduction?
27. **Describe** why it is important to know the conditions under which an aqueous oxidation-reduction reaction takes place in order to balance the ionic equation for the reaction.
28. **Explain** the steps of the oxidation-number method of balancing equations.
29. **State** what an oxidation half-reaction shows. What does a reduction half-reaction show?
30. **Write** the oxidation and reduction half-reactions for the redox equation $Pd^{2+} + Pb \rightarrow Pd + Pb^{2+}$.

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## Check Your Progress

30. **Write** the oxidation and reduction half-reactions for the redox equation $Pd^{2+} + Pb \rightarrow Pd + Pb^{2+}$.

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26. **Characterize** redox reactions by explaining how changes in oxidation number are related to the electrons transferred in a redox reaction. How are the changes related to the processes of oxidation and reduction?

**Because the nucleus (specifically, number of protons) never changes during this type of reaction, whenever there is a transfer of electrons to or from a chemical species, there is a change in the net charge of that species. Oxidation increases the oxidation number; reduction reduces it.**

27. **Describe** why it is important to know the conditions under which an aqueous redox reaction takes place in order to balance the ionic equation for the reaction.

**It is important to know whether $H_2O$ and either $H^+$ or $OH^-$ ions are available to balance the equation.**

28. **Explain** the steps of the oxidation-number method of balancing equations.

**Assign an oxidation number of each element in the equation. Identify the atoms or ions that are oxidized and reduced. Adjust the coefficients so that the total increase in oxidation number is equal to the total decrease in oxidation number. Use the conventional method to balance the remainder of the equation.**

29. **State** what an oxidation half-reaction shows. What does a reduction half-reaction show?

**An oxidation half-reaction shows the loss of electrons from an atom or ion that undergoes an increase in oxidation number. A reduction half-reaction shows the combining of electrons with an atom or ion that undergoes a decrease in oxidation number.**

30. **Write** the oxidation and reduction half-reactions for this redox equation:

$Pb(s) + Pd(NO_3)_2(aq) \rightarrow Pb(NO_3)_2(aq) + Pd(s)$

**oxidation:** $\mathbf{Pb \rightarrow Pb^{2+} + 2e^-}$

**reduction:** $\mathbf{Pd^{2+} + 2e^- \rightarrow Pd}$

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# Extra Exam Practice on Q12 to 17

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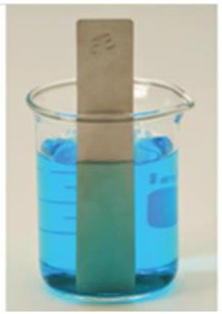
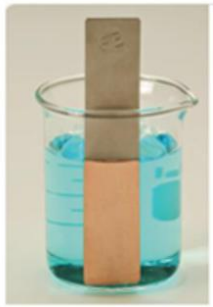
When you put an iron plate into a solution of copper(II) sulfate, as shown in the figure below , the iron plate colour changes. What explains this?

عند وضع صفيحة من الحديد في محلول كبريتات النحاس (II) كما في الشكل أدناه يتغير لون صفيحة الحديد . ما الذي يفسر ذلك؟

$$Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + Fe_2(SO_4)_3(aq)$$





A.

| نصف تفاعل الاختزال<br>Reduction –half reaction | نصف تفاعل الأكسدة<br>Oxidation –half reaction |
|---|---|
| $3Cu^{2+} + 6e^- \longrightarrow 3Cu$ | $2Fe \longrightarrow 2Fe^{3+}+6e^-$ |

B.

| نصف تفاعل الاختزال<br>Reduction –half reaction | نصف تفاعل الأكسدة<br>Oxidation –half reaction |
|---|---|
| $2Fe \longrightarrow 2Fe^{3+}+6e^-$ | $3Cu^{2+} + 6e^- \longrightarrow 3Cu$ |

C. Iron ions are reduced, and copper atoms are oxidized

تم اختزال أيونات الحديد وأكسدة ذرات النحاس

D. Both iron and copper ions are reduced

تم اختزال كلًا من أيونات الحديد وأيونات النحاس

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| What should be the value of X to balance the following half reaction ?<br>ماذا يجب أن تكون قيمة X لموازنة معادلة نصف التفاعل التالي<br>$SO_2 + 2H_2O \rightarrow SO_4^{2-} + Xe^- + 4H^+$ |
|---|
| A. 1 |
| B. 3 |
| C. 5 |
| D. 2 |

| What should be the value of X to balance the following half reaction ?<br>ماذا يجب أن تكون قيمة X لموازنة معادلة نصف التفاعل التالي<br>$MnO_4^- + 8H^+ + Xe^- \rightarrow Mn^{2+} + 4H_2O$ |
|---|
| A. 1 |
| B. 3 |
| C. 5 |
| D. 2 |

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| | |
|---|---|
| When balancing equations for redox reactions in **a basic** solution. What is done in the **last** step in the balancing method? | عند موازنة معادلات تفاعلات الأكسدة والاختزال في محلول **قاعدي**. ماذا يتم في الخطوة **الأخيرة** في طريقة الوزن؟ |
| A. Hydroxide ions ($OH^-$) and water molecules are added to either side of the equation | تتم إضافة ايونات هيدروكسيد ( $OH^-$) وجزيئات الماء لأي من طرفي المعادلة |
| B. Hydrogen ions ($H^+$) and water molecules are added to either side of the equation | تتم إضافة ايونات هيدروجين ( $H^+$) وجزيئات الماء لأي من طرفي المعادلة |
| C. Adjusting the coefficients so that the number of electrons lost in oxidation equals the number of electrons gained in reduction | يتم ضبط المعاملات بحيث يكون عدد الإلكترونات المفقودة في الأكسدة مساويًا عدد الإلكترونات المكتسبة في الاختزال |
| D. The numbers of atoms only balanced in both sides of the equation | تتم موازنة أعداد الذرات فقط في طرفي المعادلة |

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When using half-reactions to balance the redox equation below in acidic solution. Which of the following is the correct balanced equation?

عند استخدم طريقة نصف التفاعل لوزن معادلة الأكسدة والاختزال أدناه في المحلول الحمضي. أي مما يأتي هي المعادلة الموزونة الصحيحة؟

$$ClO_3^- + SO_2 \rightarrow Cl^- + SO_4^{2-}$$

A. $3SO_2 + ClO_3^- + 3H_2O \longrightarrow 3SO_4^{2-} + Cl^- + 6H^+$

B. $3SO_2 + ClO_3^- + 6H^+ \longrightarrow 3SO_4^{2-} + Cl^- + 3H_2O$

C. $2SO_2 + 3ClO_3^- + 6H^+ \longrightarrow 2SO_4^{2-} + 3Cl^- + 3H_2O$

D. $2SO_2 + 3ClO_3^- + 3H_2O \longrightarrow 2SO_4^{2-} + 3Cl^- + 6H^+$

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When the following redox equation is balanced in an **acidic** solution, which of the following is **correct**?

عند وزن معادلة الأكسدة والاختزال التالية في محلول **حمضي**. أي مما يأتي **صحيح**؟

$$Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \rightarrow Cr^{3+}_{(aq)} + Fe^{3+}_{(aq)}$$

| | عدد أيونات الهيدروجين $H^+$ Number of $H^+$ ions | موقع $H^+$ في المعادلة الموزونة $H^+$ position in the balanced equation | عدد جزيئات الماء $H_2O$ Number of water molecules | موقع $H_2O$ في المعادلة الموزونة $H_2O$ position in the balanced equation |
|---|---|---|---|---|
| A | 14 | جهة النواتج (إلى اليمين) products side (right) | 7 | جهة المتفاعلات (إلى اليسار) reactants side (left) |
| B | 14 | جهة المتفاعلات (إلى اليسار) reactants side (left) | 7 | جهة النواتج (إلى اليمين) products side (right) |
| C | 7 | جهة المتفاعلات (إلى اليسار) reactants side (left) | 14 | جهة النواتج (إلى اليمين) products side (right) |
| D | 7 | جهة النواتج (إلى اليمين) products side (right) | 14 | جهة المتفاعلات (إلى اليسار) reactants side (left) |

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| Which of the following is the correct balanced redox equation for the following reaction using the half-reaction method ? | أي مما يلي هي المعادلة الموزونة الصحيحة لتفاعل الأكسدة والاختزال التالي باستخدام طريقة نصف التفاعل؟ |
|---|---|
| $NO_3^- + Cl^- \rightarrow ClO^- + NO_{(\text{in acidic } solution)}$ | $NO_3^- + Cl^- \rightarrow ClO^- + NO$ (في محلول حمضي) |

A- $2NO_3^- + 3Cl^- \rightarrow 2NO + 3ClO^-$

B- $2NO_3^- + 3Cl^- \rightarrow 2NO + 3ClO^- + H_2O$

C- $2NO_3^- + 3Cl^- + 2H^+ \rightarrow 2NO + 3ClO^- + H_2O$

D- $2NO_3^- + 3Cl^- + H^+ \rightarrow 2NO + 3ClO^- + H_2O$

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The reaction between nickel and copper (II) chloride shown below. What are the half-reactions for this redox reaction?

التفاعل بين النيكل وكلوريد النحاس (II) الموضح أدناه. ما نصفا تفاعلات الأكسدة والاختزال؟

$$Ni_{(s)} + CuCl_{2(aq)} \longrightarrow Cu_{(s)} + NiCl_{2(aq)}$$

| | نصف تفاعل الأكسدة<br>Oxidation Half Reaction | نصف تفاعل الاختزال<br>Reduction Half Reaction |
|---|---|---|
| A | $Ni \longrightarrow Ni^{2+} + 2e^-$ | $Cl_2 \longrightarrow 2Cl^- + 2e^-$ |
| B | $Ni \longrightarrow Ni^{2+} + e^-$ | $Cu^+ + e^- \longrightarrow Cu$ |
| C | $Ni \longrightarrow Ni^{2+} + 2e^-$ | $Cu^{2+} + 2e^- \longrightarrow Cu$ |
| D | $Cu^+ + e^- \longrightarrow Cu$ | $Ni \longrightarrow Ni^{2+} + 2e^-$ |

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Using the half-reaction method to balance the oxidation–reduction reactions, what is the correct balanced equation for the following reaction in **acidic solution**?

باستخدام طريقة نصف التفاعل لوزن تفاعلات الأكسدة والاختزال، ما المعادلة الموزونة الصحيحة للتفاعل التالي في **المحلول الحمضي**؟

$$I_{2\,(s)} + S_2O_3^{2-}{}_{(aq)} \longrightarrow S_2O_4^{2-}{}_{(aq)} + I^-_{(aq)}$$

$$S_2O_3^{2-} + H_2O + I_2 \longrightarrow S_2O_4^{2-} + 2H^+ + 2I^-$$

$$S_2O_3^{2-} + I_2 + 2H^+ \longrightarrow S_2O_4^{2-} + H_2O + 2I^-$$

$$2S_2O_3^{2-} + 2H_2O + I_2 \longrightarrow 2S_2O_4^{2-} + 4H^+ + 2I^-$$

$$S_2O_3^{2-} + H_2O + I_2 \longrightarrow S_2O_4^{2-} + 2H^+ + I^-$$

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Using the half-reaction method. Which of the following is the balanced equation for the reaction below in an **acidic** solution?

باستخدام طريقة نصف التفاعل. أي مما يأتي هي المعادلة الموزونة للتفاعل أدناه في محلول **حمضي**؟

$$Cl^-(aq) + NO_3^-(aq) \rightarrow ClO^-(aq) + NO(g)$$

$$3Cl^-_{(aq)} + 2NO_3^-{}_{(aq)} + 2H^+_{(aq)} \longrightarrow 3ClO^-_{(aq)} + 2NO_{(g)} + H_2O_{(l)}$$

$$2Cl^-_{(aq)} + 3NO_3^-{}_{(aq)} + 2H^+_{(aq)} \longrightarrow 2ClO^-_{(aq)} + 3NO_{(g)} + H_2O_{(l)}$$

$$3Cl^-_{(aq)} + 2NO_3^-{}_{(aq)} + H_2O_{(l)} \longrightarrow 3ClO^-_{(aq)} + 2NO_{(g)} + 2H^+_{(aq)}$$

$$2Cl^-_{(aq)} + 3NO_3^-{}_{(aq)} + H_2O_{(l)} \longrightarrow 2ClO^-_{(aq)} + 3NO_{(g)} + 2H^+_{(aq)}$$

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If given the following two redox half-reactions. What minimum numbers of $Ce^{4+}$ ions and $Sn^{2+}$ ions would have to react in order to have zero electrons left over?

إذا أُعطيت نصفي تفاعل الأكسدة والاختزال التاليين. ما هو الحد الأدنى لأيونات $Ce^{4+}$ وأيونات $Sn^{2+}$ التي يُمكنها التفاعل لكيلا يبقى إلكترونات؟

| | |
|---|---|
| $Ce^{4+} + e^- \longrightarrow Ce^{3+}$ | نصف تفاعل الاختزال<br>Reduction Half-Reaction |
| $Sn^{2+} \longrightarrow Sn^{4+} + 2e^-$ | نصف تفاعل الأكسدة<br>Oxidation Half-Reaction |

| $Sn^{2+}$ | $Ce^{4+}$ | |
|---|---|---|
| 3 | 2 | A |
| 1 | 2 | B |
| 2 | 1 | C |
| 2 | 3 | D |

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| | | | |
|---|---|---|---|
| Q18 | Describe an electrochemical cell while specifying its types | figures 1,2+ textbook | 25,26 |
| Q19 | Identify components of a voltaic or galvanic cell (anode, cathode, salt bridge or porous barrier, wires, electrolyte compartments); while explaining the role of each component, when does the reaction start and determining the direction of electron and current flow | figures 1,2+ textbook | 25,26 |

# Redox in Electrochemistry

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- **Electrochemistry** is the study of the redox processes by which **chemical energy** is converted to **electrical energy** and vice versa.

- Redox reactions involve a transfer of electrons from the species that is oxidized to the species that is reduced.

# Redox in Electrochemistry

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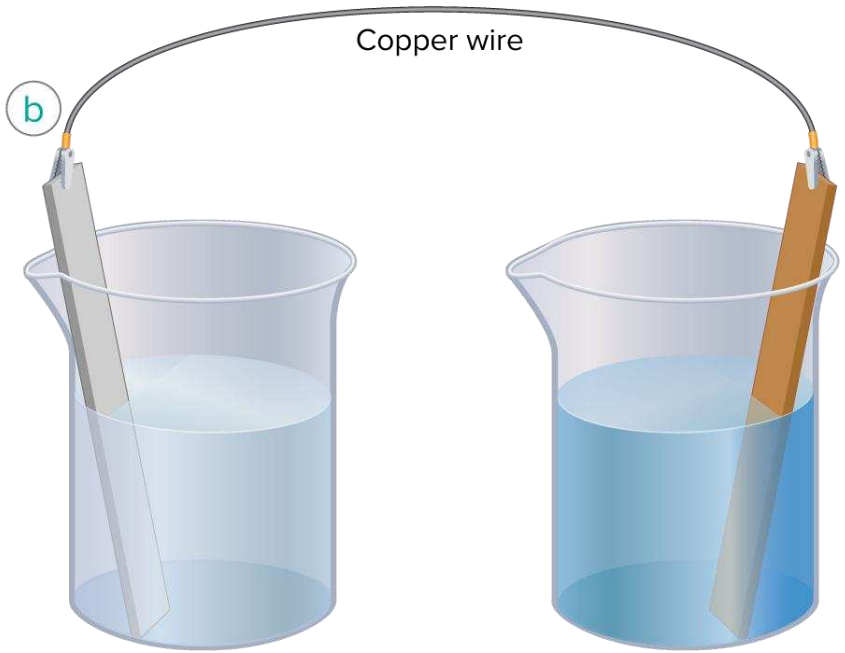
- A **voltaic cell** is a type of electrochemical cell that **converts chemical energy to electrical energy** by a **spontaneous** redox reaction.




A wire joining the zinc and copper strips provides a pathway for the flow of electrons, but the pathway is not complete. Electron transfer is still not possible. (**Ions build up on the electrodes.)**

How to fix this? Use a **Salt Bridge**

# Redox in Electrochemistry

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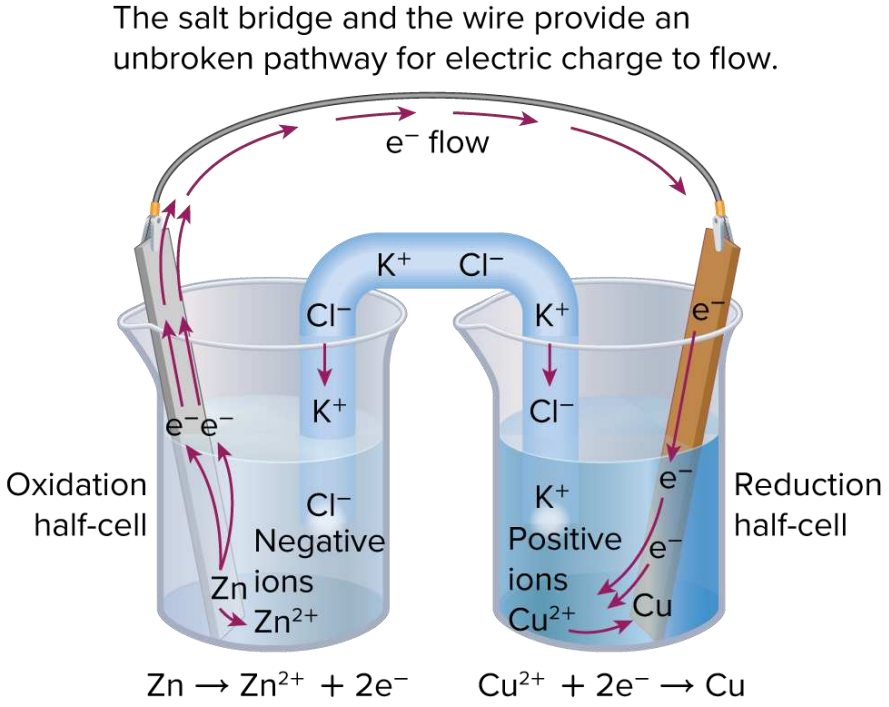
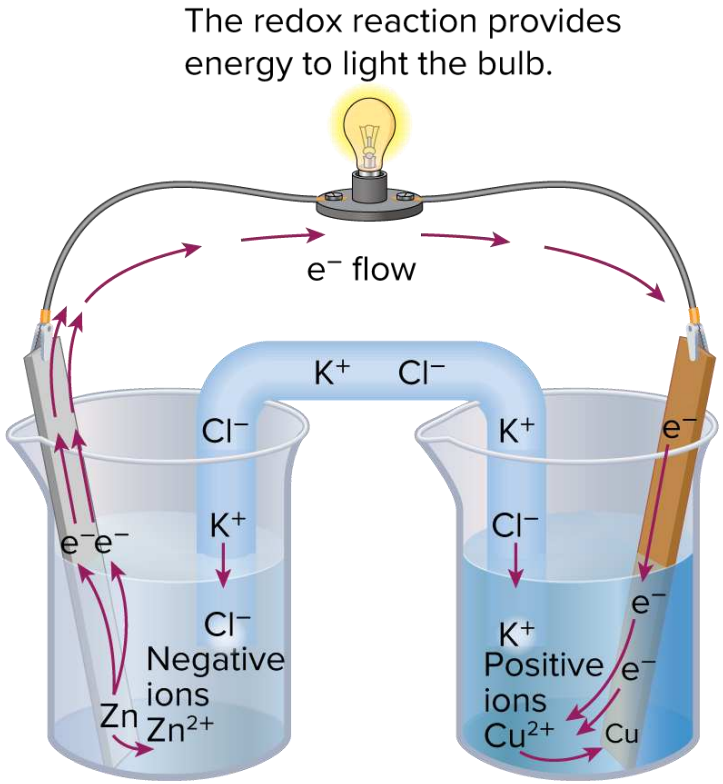
- A **voltaic cell** is a type of electrochemical cell that **converts chemical energy to electrical energy** by a **spontaneous** redox reaction.

- A **salt bridge** is a pathway to allow the passage of ions from one side to another, ***so that ions do not build up around the electrodes.***




The addition of a salt bridge completes the pathway. Negative ions move through the salt bridge to the zinc side. Positive ions move through the bridge to the copper side.

# Quiz

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1. Which is a pathway to allow the passage of ions from one side of a redox process to the other?

1. voltaic cell
2. half-cell
3. salt bridge **CORRECT**
4. standard hydrogen electrode

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# Quiz

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2. Which is an apparatus that uses a redox reaction to produce electrical energy or uses electrical energy to cause a chemical reaction?

A salt bridge

B half-cell

C standard hydrogen electrode

D electrochemical cell

# Quiz

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3. Which is a type of electrochemical cell that converts chemical energy to electrical energy by a spontaneous redox reaction?

A voltaic cell

C half-cell

B salt cell

D oxidation cell

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# Chemistry of Voltaic Cells: Main points

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- An electrochemical cell consists of two parts called **half-cells**, in which the separate oxidation and reduction reactions take place.
- The electrode where ***oxidation*** takes place is called the ***anode.***
- The ***cathode*** is the electrode where ***reduction*** occurs.

- **- Salt bridge:** It is a pathway to maintain solution neutrality by allowing the passage of ions from one side to another, where anions migrate towards anode and cations migrate towards cathode.

- It usually contains a conducting neutral soluble solution as KCl, NaCl or NaNO3
- *- The spontaneous reaction starts when the connecting metal wire and salt bridge are in place*

# Chemistry of Voltaic Cells: Main points

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- Electrons flow through the wire from the oxidation-half reaction (anode) to the reduction-half reaction (cathode) while positive and negative ions move through the salt bridge

*- Current flows from cathode to anode*

- The flow of electrons through the wire and the flow of ions through the salt bridge make up the electric current

- ***Electric potential energy*** is a measure of the amount of current that can be generated from a voltaic cell to do work.

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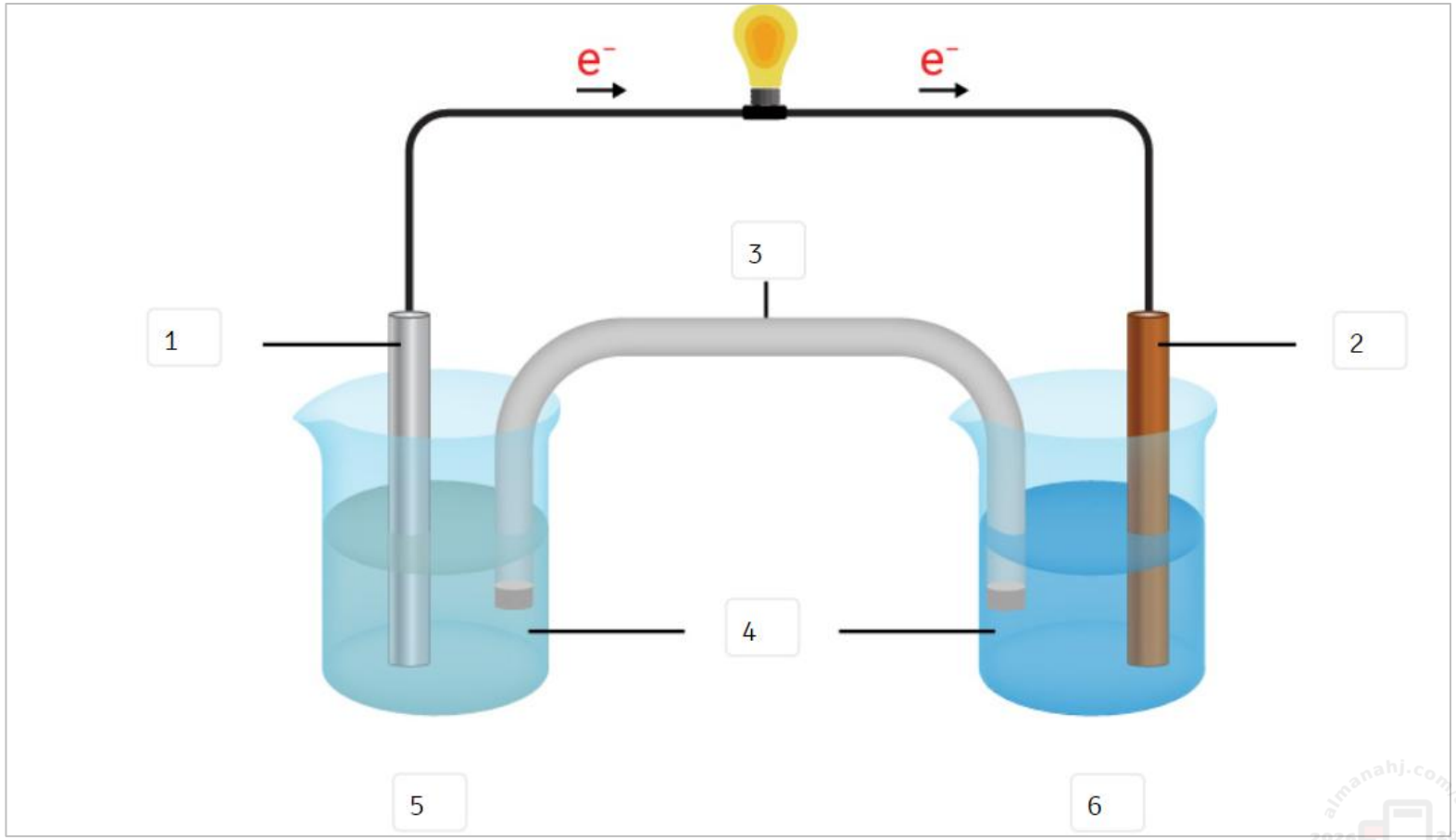
# Check Point (Paired Activity)

Assign the components to the numbers in the image.

*Cathode
*Anode
*Oxidation-half cell
*Reduction-half cell
*Salt bridge
*electrolyte



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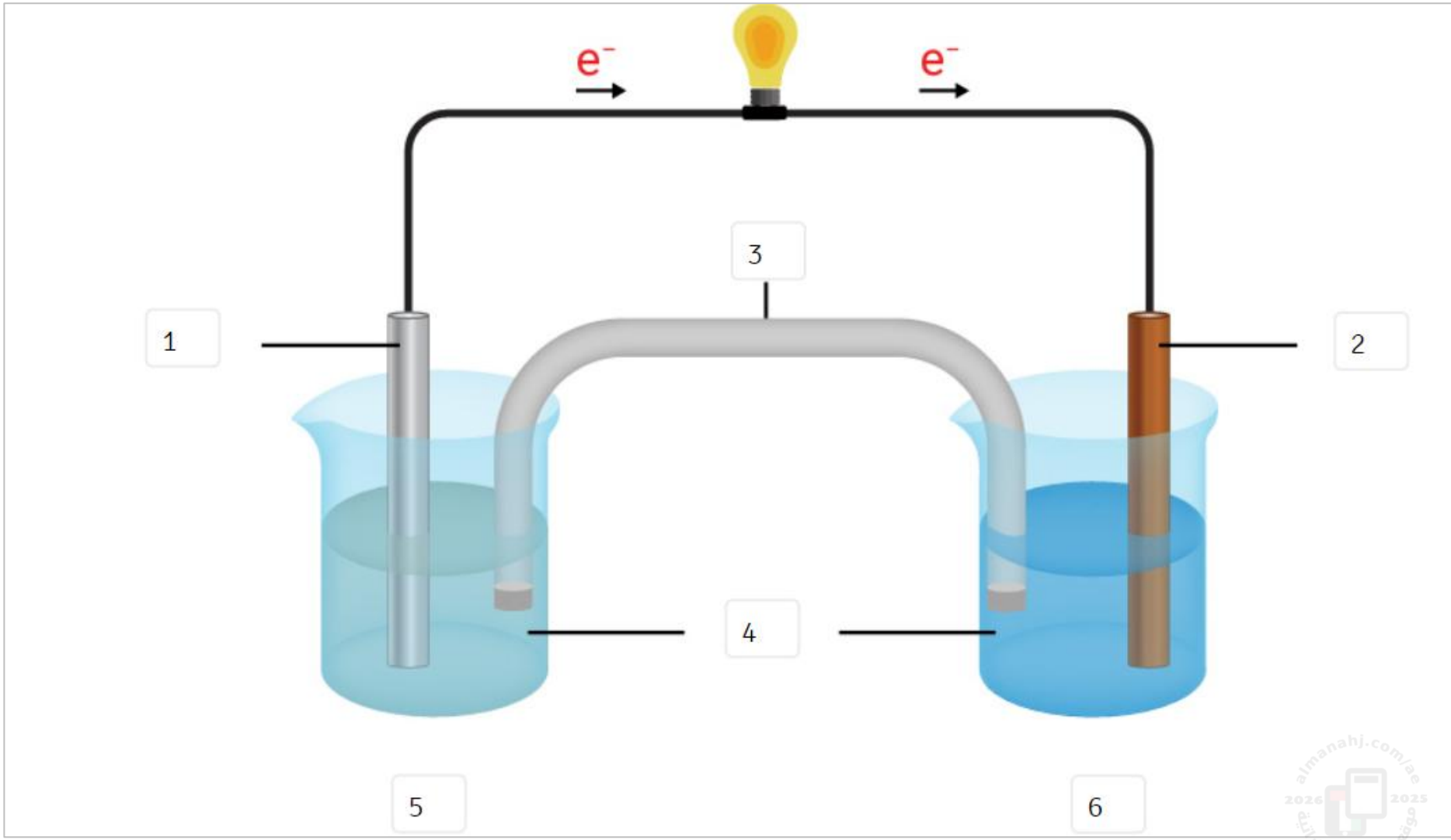
1. anode
2. cathode
3. salt bridge
4. electrolyte
5. oxidation half-cell
6. reduction half-cell



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| | | | |
|---|---|---|---|
| Q20 | Write the oxidation and reduction half-reactions occurring at cathode and anode for a voltaic cell | textbook+fiqure 6 | 27,31 |
| Q21 | Define the reduction potential and standard electrode potential (E°) | table 1+ textbook | 29,30 |
| Q22 | Describe standard hydrogen electrode (SHE), while identifying the importance of its E° value and writing the half-cell reactions of the two possible reactions that could occur at the hydrogen electrode | figures 5,6+ textbook | 29,31 |

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# What is "Standard Electrode Potential" $E^o$ ?

**Standard electrode potential($E^o$)** is the potential or voltage developed by a metal or other material immersed in an electrolyte solution relative to the **potential of the hydrogen electrode which is zero.**

Why is the standard potential of hydrogen ($H_2$) zero??

***Because it is used as the reference point of the scale.***

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# *Example from table 1 in the book (Standard Reduction potentials $E^o$)*

**Table 1 Standard Reduction Potentials**

| Half-Reaction | $E^o$ (V) | Half-Reaction | $E^o$ (V) |
|---|---|---|---|
| $Li^+ + e^- \rightleftarrows Li$ | −3.0401 | $Cu^{2+} + e^- \rightleftarrows Cu^+$ | +0.153 |
| $Ca^{2+} + 2e^- \rightleftarrows Ca$ | −2.868 | $Cu^{2+} + 2e^- \rightleftarrows Cu$ | +0.3419 |
| $Na^+ + e^- \rightleftarrows Na$ | −2.71 | $O_2 + 2H_2O + 4e^- \rightleftarrows 4OH^-$ | +0.401 |
| $Mg^{2+} + 2e^- \rightleftarrows Mg$ | −2.372 | $I_2 + 2e^- \rightleftarrows 2I^-$ | +0.5355 |
| $Be^{2+} + 2e^- \rightleftarrows Be$ | −1.847 | $Fe^{3+} + e^- \rightleftarrows Fe^{2+}$ | +0.771 |
| $Al^{3+} + 3e^- \rightleftarrows Al$ | −1.662 | $NO_3^- + 2H^+ + e^- \rightleftarrows NO_2 + H_2O$ | +0.775 |
| $Mn^{2+} + 2e^- \rightleftarrows Mn$ | −1.185 | $Hg_2^{2+} + 2e^- \rightleftarrows 2Hg$ | +0.7973 |
| $Cr^{2+} + 2e^- \rightleftarrows Cr$ | −0.913 | $Ag^+ + e^- \rightleftarrows Ag$ | +0.7996 |
| $2H_2O + 2e^- \rightleftarrows H_2 + 2OH^-$ | −0.8277 | $Hg^{2+} + 2e^- \rightleftarrows Hg$ | +0.851 |
| $Zn^{2+} + 2e^- \rightleftarrows Zn$ | −0.7618 | $2Hg^{2+} + 2e^- \rightleftarrows Hg_2^{2+}$ | +0.920 |

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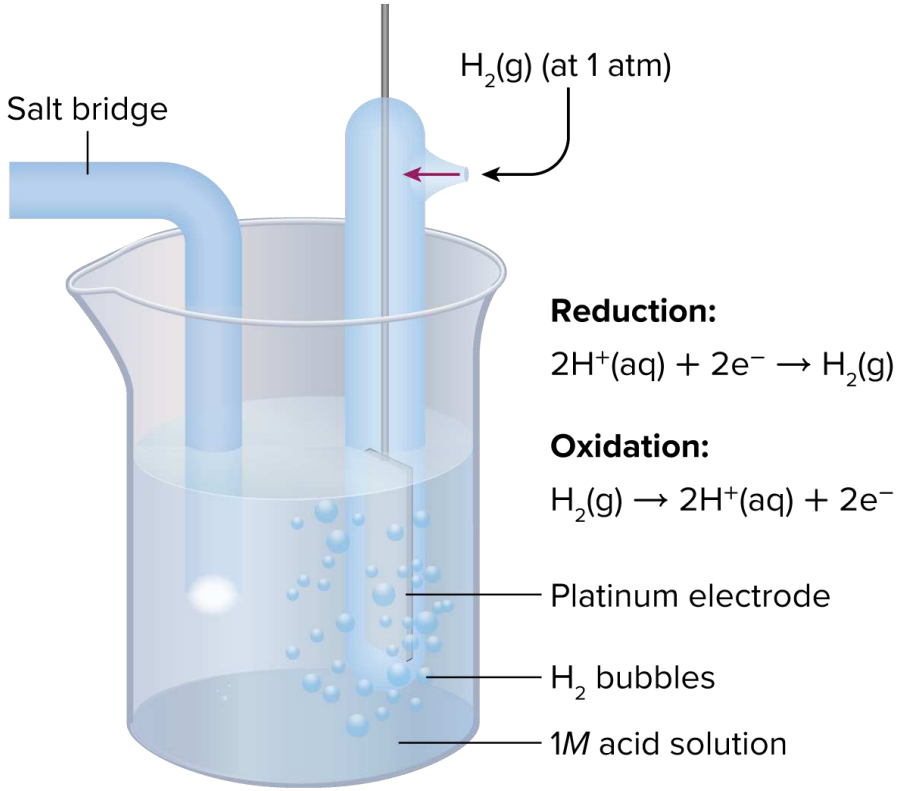


$$2H^+(aq) + 2e^- \underset{\text{Oxidation}}{\overset{\text{Reduction}}{\rightleftharpoons}} H_2(g) \qquad E^0 = 0.000\ \text{V}$$

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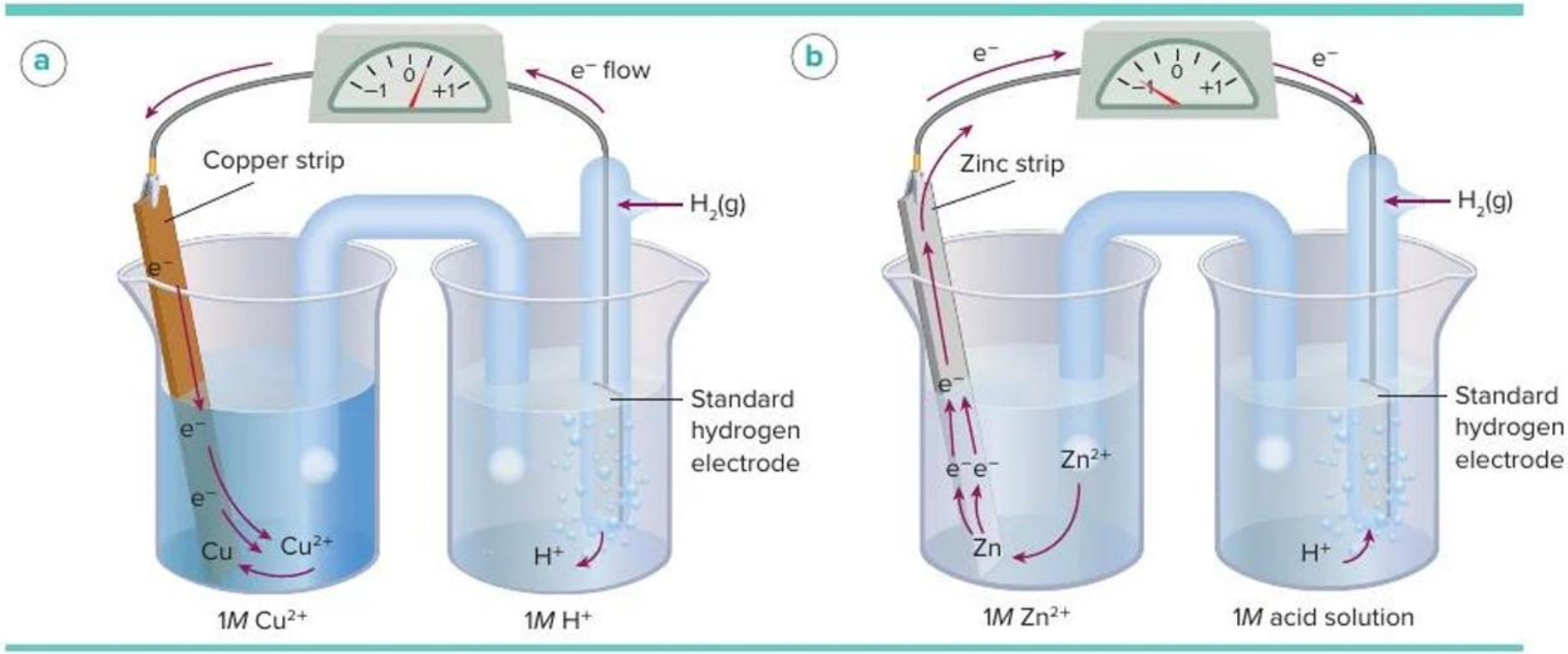


**Figure 6** **a.** When a $Cu^{2+}$ | Cu electrode is connected to the hydrogen electrode, electrons flow toward the copper strip and reduce $Cu^{2+}$ ions to Cu atoms. **b.** When a Zn | $Zn^{2+}$ electrode is connected to the hydrogen electrode, electrons flow away from the zinc strip and zinc atoms are oxidized to $Zn^{2+}$ ions.

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# *Example from table 1 in the book (Standard Reduction potentials $E^o$)*

**Table 1 Standard Reduction Potentials**

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| $Na^+ + e^- \rightleftarrows Na$ | −2.71 | $O_2 + 2H_2O + 4e^- \rightleftarrows 4OH^-$ | +0.401 |
| $Mg^{2+} + 2e^- \rightleftarrows Mg$ | −2.372 | $I_2 + 2e^- \rightleftarrows 2I^-$ | +0.5355 |
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| $Mn^{2+} + 2e^- \rightleftarrows Mn$ | −1.185 | $Hg_2^{2+} + 2e^- \rightleftarrows 2Hg$ | +0.7973 |
| $Cr^{2+} + 2e^- \rightleftarrows Cr$ | −0.913 | $Ag^+ + e^- \rightleftarrows Ag$ | +0.7996 |
| $2H_2O + 2e^- \rightleftarrows H_2 + 2OH^-$ | −0.8277 | $Hg^{2+} + 2e^- \rightleftarrows Hg$ | +0.851 |
| $Zn^{2+} + 2e^- \rightleftarrows Zn$ | −0.7618 | $2Hg^{2+} + 2e^- \rightleftarrows Hg_2^{2+}$ | +0.920 |

# Calculating Electrochemical Cell Potentials

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- Formula for Cell Potential

$$E^0_{cell} = E^0_{reduction} - E^0_{oxidation}$$

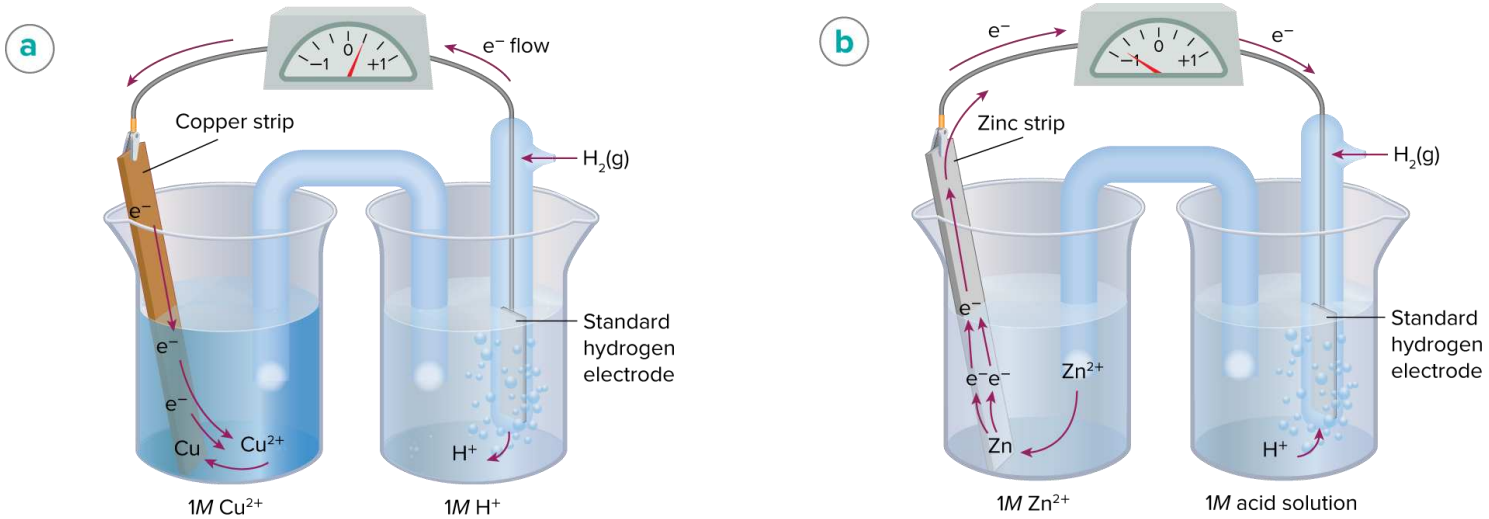
or

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$

- The standard potential of a cell is the **standard potential of the reduction half-cell** minus **the standard potential of the oxidation half-cell.**

If **E°cell is positive (>0),** the redox reaction is **spontaneous (تلقائي)**

If **E°cell is negative (< 0),** the redox reaction is **nonspontaneous (غير تلقائي)**



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## Table 1 Standard Reduction Potentials

| Half-Reaction | $E^0$ (V) | Half-Reaction | $E^0$ (V) |
|---|---|---|---|
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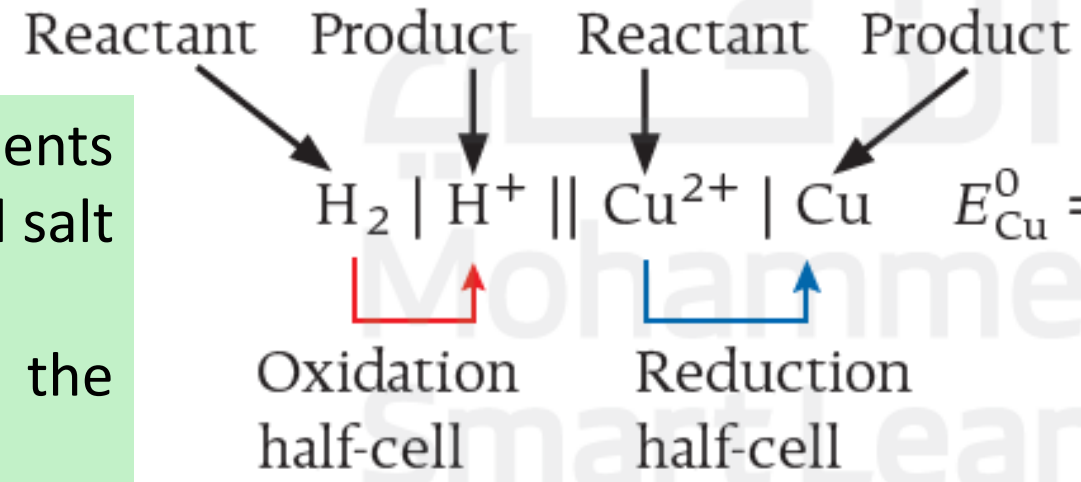
| $Na^+ + e^- \rightleftarrows Na$ | −2.71 | $O_2 + 2H_2O + 4e^- \rightleftarrows 4OH^-$ | +0.401 |
| $Mg^{2+} + 2e^- \rightleftarrows Mg$ | −2.372 | $I_2 + 2e^- \rightleftarrows 2I^-$ | +0.5355 |
| $Be^{2+} + 2e^- \rightleftarrows Be$ | −1.847 | $Fe^{3+} + e^- \rightleftarrows Fe^{2+}$ | +0.771 |
| $Al^{3+} + 3e^- \rightleftarrows Al$ | −1.662 | $NO_3^- + 2H^+ + e^- \rightleftarrows NO_2 + H_2O$ | +0.775 |
| $Mn^{2+} + 2e^- \rightleftarrows Mn$ | −1.185 | $Hg_2^{2+} + 2e^- \rightleftarrows 2Hg$ | +0.7973 |
| $Cr^{2+} + 2e^- \rightleftarrows Cr$ | −0.913 | $Ag^+ + e^- \rightleftarrows Ag$ | +0.7996 |
| $2H_2O + 2e^- \rightleftarrows H_2 + 2OH^-$ | −0.8277 | $Hg^{2+} + 2e^- \rightleftarrows Hg$ | +0.851 |
| $Zn^{2+} + 2e^- \rightleftarrows Zn$ | −0.7618 | $2Hg^{2+} + 2e^- \rightleftarrows Hg_2^{2+}$ | +0.920 |

**Write** the cell notation and the overall chemical equation for a redox reaction occurring in a voltaic cell. **Page 184**

$$H_2(g) \rightarrow 2H^+(aq) + 2e^- \text{ (oxidation half-cell reaction)}$$
$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \text{ (reduction half-cell reaction)}$$

$$H_2(g) + Cu^{2+}(aq) \rightarrow 2H^+(aq) + Cu(s) \text{ (overall redox reaction)}$$

This reaction can be written in a form called cell notation.




**||** represents the wire and salt bridge connecting the half-cells

***Cell Notation***

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| | | | |
|---|---|---|---|
| Q23 | Use the half-cell standard reduction potentials to calculate the electrochemical cell standard potential, while determining whether the redox reactions are spontaneous or non-spontaneous | textbook+ example problem1 + practice problems+table3 | 32,33,34 |
| Q24 | Write the cell notation and the overall chemical equation for a redox reaction occurring in a voltaic cell | textbook+ example problem1 + practice problems+table3 | 31,32,33, 34 |
| Q25 | Use the half-cell standard reduction potentials to calculate the electrochemical cell standard potential, while determining whether the redox reactions are spontaneous or non-spontaneous | textbook+ practice problems+check your progress | 34,35 |

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# Calculate a Cell Potential

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**IN-CLASS EXAMPLE**

*Use with Example Problem 1.*

**Problem**

The following reduction half-reactions represent the half-cells of a voltaic cell.

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$

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

Determine the **overall cell reaction** and the **standard cell potential**. **Describe the cell using cell notation.**

**KNOWN**

Standard reduction potentials for the half-cells

$E^0_{cell} = E^0_{reduction} - E^0_{oxidation}$

**UNKNOWN**

overall cell reaction = ?

$E^0_{cell} = ?$

cell notation = ?

*Given to you in the exam*

**The standard reduction potentials of each half-reaction in Table 1.**

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$ $\quad E^0 = +0.536$ V

$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ $\quad E^0 = -0.447$ V

**First, which one is oxidized, and which one is reduced?**

The half reaction with the **higher $E^0$** is **reduced**

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# Calculate a Cell Potential

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Easy Chemistry4all

**IN-CLASS EXAMPLE**

*Given to you in the exam*

**The standard reduction potentials of each half-reaction in Table 1.**

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$ $\quad E^0 = +0.536$ V

$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ $\quad E^0 = -0.447$ V

The half reaction with the **higher $E^0$** is **reduced**

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# Calculate a Cell Potential

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IN-CLASS EXAMPLE

*SOLVE FOR THE UNKNOWN (continued)*

- **Rewrite the iron half-reaction in the correct direction.**

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$ (reduction half-cell reaction) **because it has the higher $E^0$**

$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ (oxidation half-cell reaction)

- Add the two equations.

$$I_2(s) + Fe(s) \rightarrow Fe^{2+}(aq) + 2I^-(aq)$$ **Overall reaction**

Calculate the standard cell potential.

- State the formula for cell potential.

$$E^{\mathbf{0}}_{\mathbf{cell}} = E^{0}_{\text{reduction}} - E^{0}_{\text{oxidation}}$$

- Substitute $E^{0}_{I_2|I^-}$ and $E^{0}_{Fe^{2+}|Fe}$ in the generic equation.

$$E^{\mathbf{0}}_{\mathbf{cell}} = E^{0}_{I_2|I^-} - E^{0}_{Fe^{2+}|Fe}$$

# Calculate a Cell Potential

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IN-CLASS EXAMPLE

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$ $E^0 = +0.536\ V$ **(Reduction)**

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ $E^0 = -0.447\ V$ **(Oxidation)**

$$E^0_{cell} = +0.536\ V - (-0.447\ V)$$

$$E^0_{cell} = \mathbf{+0.983V}$$

Describe the cell using cell notation.

- First, write the oxidation half-reaction using cell notation: **reactant then product**.

$$Fe \mid Fe^{2+}$$

- Next, write the reduction half-reaction to the right. Separate the half-cells by a double vertical line.

$$Fe \mid Fe^{2+} \parallel I_2 \mid I^-$$

**Cell notation: $Fe \mid Fe^{2+} \parallel I_2 \mid I^-$**

# Calculate a Cell Potential

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IN-CLASS EXAMPLE

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$ $E^0 = +0.536\ V$ **(Reduction)**

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ $E^0 = -0.447\ V$ **(Oxidation)**

$$E^0_{cell} = +0.536\ V - (-0.447\ V)$$

$$E^0_{cell} = +0.983V$$

Is this reaction spontaneous or nonspontaneous?

**It is spontaneous because the cell potential $E^0$cell is positive**

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For each of these pairs of half-reactions, write the balanced equation for the overall cell reaction, and calculate the standard cell potential. Describe the reaction using cell notation. Refer to the module on redox reactions to review writing and balancing redox equations.

1. $Pt^{2+}(aq) + 2e^- \rightleftharpoons Pt(s)$ and $Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$
2. $Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$ and $Cr^{3+}(aq) + 3e^- \rightleftharpoons Cr(s)$
3. $Hg^{2+}(aq) + 2e^- \rightleftharpoons Hg(l)$ and $Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$
4. **CHALLENGE** Write the balanced equation for the cell reaction and calculate the standard cell potential for the reaction that occurs when these half-cells are connected. Describe the reaction using cell notation.

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

   $O_2 + 2H_2O + 4e^- \rightleftharpoons 4OH^-$

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

$$E^\circ_{red} = +1.18\ V$$

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

$$E^\circ_{red} = -0.1375\ V$$

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

$$E^\circ_{red} = -0.28\ V$$

$$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$$

$$E^\circ_{red} = -0.744\ V$$

$$Hg^{2+}(aq) + 2e^- \rightarrow Hg(l)$$

$$E^\circ_{red} = +0.851\ V$$

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

$$E^\circ_{red} = -0.913\ V$$

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

$$E^\circ_{red} = +0.957\ V$$

$$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$$

$$E^\circ_{red} = +0.401\ V$$

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## PRACTICE Problems

ADDITIONAL PRACTICE

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For each of these pairs of half-reactions, write the balanced equation for the overall cell reaction, and calculate the standard cell potential. Describe the reaction using cell notation. Refer to the module on redox reactions to review writing and balancing redox equations.

1. $Pt^{2+}(aq) + 2e^- \rightleftharpoons Pt(s)$ and $Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$
2. $Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$ and $Cr^{3+}(aq) + 3e^- \rightleftharpoons Cr(s)$
3. $Hg^{2+}(aq) + 2e^- \rightleftharpoons Hg(l)$ and $Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$
4. **CHALLENGE** Write the balanced equation for the cell reaction and calculate the standard cell potential for the reaction that occurs when these half-cells are connected. Describe the reaction using cell notation.

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

   $O_2 + 2H_2O + 4e^- \rightleftharpoons 4OH^-$

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

$$E^\circ_{red} = +1.18\ V$$

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

$$E^\circ_{red} = -0.1375\ V$$

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## PRACTICE Problems

ADDITIONAL PRACTICE

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For each of these pairs of half-reactions, write the balanced equation for the overall cell reaction, and calculate the standard cell potential. Describe the reaction using cell notation. Refer to the module on redox reactions to review writing and balancing redox equations.

1. $Pt^{2+}(aq) + 2e^- \rightleftharpoons Pt(s)$ and $Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$
2. $Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$ and $Cr^{3+}(aq) + 3e^- \rightleftharpoons Cr(s)$
3. $Hg^{2+}(aq) + 2e^- \rightleftharpoons Hg(l)$ and $Cr^{2+}(aq) + 2e^- \rightleftharpoons Cr(s)$
4. **CHALLENGE** Write the balanced equation for the cell reaction and calculate the standard cell potential for the reaction that occurs when these half-cells are connected. Describe the reaction using cell notation.

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

   $O_2 + 2H_2O + 4e^- \rightleftharpoons 4OH^-$

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

$$E^\circ_{red} = +0.957\ V$$

$$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$$

$$E^\circ_{red} = +0.401\ V$$

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## Practice Problems

*page 631*

**For each of these pairs of half-reactions, write the balanced equation for the overall cell reaction and calculate the standard cell potential. Express the reaction using cell notation. Refer to the chapter on redox reactions to review writing and balancing redox equations.**

1. $Pt^{2+}(aq) + 2e^- \rightarrow Pt(s)$ and
   $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$

   **$Pt^{2+}(aq) + Sn(s) \rightarrow Pt(s) + Sn^{2+}(aq)$**

   **$E^0_{cell} = +1.18\ V - (-0.1375\ V)$**

   **$E^0_{cell} = +1.32\ V$**

   **$Sn|Sn^{2+}||Pt^{2+}|Pt$**

2. $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$

   $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$

   **$3Co^{2+}(aq) + 2Cr(s) \rightarrow 3Co(s) + 2Cr^{3+}(aq)$**

   **$E^0_{cell} = -0.28\ V - (-0.744\ V)$**

   **$E^0_{cell} = +0.46\ V$**

   **$Cr|Cr^{3+}||Co^{2+}|Co$**

3. $Hg^{2+}(aq) + 2e^- \rightarrow Hg(l)$

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

   **$Hg^{2+}(aq) + Cr(s) \rightarrow Hg(l) + Cr^{2+}(aq)$**

   **$E^0_{cell} = +0.851\ V - (-0.913\ V)$**

   **$E^0_{cell} = +1.764\ V$**

   **$Cr|Cr^{2+}||Hg^{2+}|Hg$**

4. **Challenge** Write the balanced equation for the cell reaction and calculate the standard cell potential for the reaction that occurs when these half-cells are connected. Describe the reaction using cell notation.

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

   $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

   **$4NO_3^-(aq) + 12OH^-(aq) + 16H^+(aq) \rightarrow 3O_2(g) + 4NO(g) + 14H_2O(l)$**

   **$E^0_{cell} = +0.957\ V - (+0.401\ V)$**

   **$E^0_{cell} = +0.556\ V$**

   **$OH^-|O_2||NO_3^-|NO$**

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## PRACTICE Problems

ADDITIONAL PRACTICE

Calculate the cell potential to determine if each of the following balanced redox reactions is spontaneous as written. Use **Table 1** to help you determine the correct half-reactions.

5. $Sn(s) + Cu^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cu(s)$
6. $Mg(s) + Pb^{2+}(aq) \rightarrow Pb(s) + Mg^{2+}(aq)$
7. $2Mn^{2+}(aq) + 8H_2O(l) + 10Hg^{2+}(aq) \rightarrow 2MnO_4^-(aq) + 16H^+(aq) + 5Hg_2^{2+}(aq)$
8. $2SO_4^{2-}(aq) + Co^{2+}(aq) \rightarrow Co(s) + S_2O_8^{2-}(aq)$
9. **CHALLENGE** Using **Table 1**, write the equation and determine the cell voltage ($E^0$) for the following cell. Is the reaction spontaneous?

$$Al \mid Al^{3+} \parallel Hg^{2+} \mid Hg_2^{2+}$$

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$
$$E^\circ_{red} = +0.3419\ V$$
$$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$$
$$E^\circ_{red} = -0.1375\ V$$

$$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$$
$$E^\circ_{red} = -0.1262\ V$$
$$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$$
$$E^\circ_{red} = -2.372\ V$$

$$Hg_2^{2+}(aq) + 2e^- \rightarrow 2Hg(l)$$
$$E^\circ_{red} = +0.920\ V$$
$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$
$$E^\circ_{red} = +1.507\ V$$

$$Co^{2+}(aq) + 2e^- \rightarrow Co(s)$$
$$E^\circ_{red} = -0.28\ V$$
$$S_2O_8^{2-}(aq) + 2e^- \rightarrow 2SO_4^{2-}(aq)$$
$$E^\circ_{red} = +2.010\ V$$

$$Hg_2^{2+}(aq) + 2e^- \rightarrow 2Hg(l)$$
$$E^\circ_{red} = +0.920\ V$$
$$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$$
$$E^\circ_{red} = -1.662\ V$$

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## PRACTICE Problems

ADDITIONAL PRACTICE

Calculate the cell potential to determine if each of the following balanced redox reactions is spontaneous as written. Use **Table 1** to help you determine the correct half-reactions.

5. $Sn(s) + Cu^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cu(s)$
6. $Mg(s) + Pb^{2+}(aq) \rightarrow Pb(s) + Mg^{2+}(aq)$
7. $2Mn^{2+}(aq) + 8H_2O(l) + 10Hg^{2+}(aq) \rightarrow 2MnO_4^-(aq) + 16H^+(aq) + 5Hg_2^{2+}(aq)$
8. $2SO_4^{2-}(aq) + Co^{2+}(aq) \rightarrow Co(s) + S_2O_8^{2-}(aq)$
9. **CHALLENGE** Using **Table 1**, write the equation and determine the cell voltage ($E^0$) for the following cell. Is the reaction spontaneous?

$$Al \mid Al^{3+} \parallel Hg^{2+} \mid Hg_2^{2+}$$

$$Hg_2^{2+}(aq) + 2e^- \rightarrow 2Hg(l)$$

$$E^\circ_{red} = +0.920\ V$$

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

$$E^\circ_{red} = +1.507\ V$$

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## PRACTICE Problems

ADDITIONAL PRACTICE

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9. **CHALLENGE** Using **Table 1**, write the equation and determine the cell voltage ($E^0$) for the following cell. Is the reaction spontaneous?

$$Al \mid Al^{3+} \parallel Hg^{2+} \mid Hg_2^{2+}$$

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

$$E^\circ_{red} = +0.3419\ V$$

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

$$E^\circ_{red} = -0.1375\ V$$

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## PRACTICE Problems

ADDITIONAL PRACTICE

Calculate the cell potential to determine if each of the following balanced redox reactions is spontaneous as written. Use **Table 1** to help you determine the correct half-reactions.

5. $Sn(s) + Cu^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cu(s)$
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8. $2SO_4^{2-}(aq) + Co^{2+}(aq) \rightarrow Co(s) + S_2O_8^{2-}(aq)$
9. **CHALLENGE** Using **Table 1**, write the equation and determine the cell voltage ($E^0$) for the following cell. Is the reaction spontaneous?

$$Al \mid Al^{3+} \parallel Hg^{2+} \mid Hg_2^{2+}$$

$$Hg_2^{2+}(aq) + 2e^- \rightarrow 2Hg(l)$$

$$E^\circ_{red} = +0.920\ V$$

$$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$$

$$E^\circ_{red} = -1.662\ V$$

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**Calculate the cell potential to determine if each of the following balanced redox reactions is spontaneous as written. Use Table 1 to help you determine the correct half-reactions.**

**5.** $Sn(s) + Cu^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cu(s)$

$E^0_{cell} = +0.3419\ V - (-0.1375\ V)$

$E^0_{cell} = +0.4794\ V$

$E^0_{cell} > 0$; **spontaneous**

**6.** $Mg(s) + Pb^{2+}(aq) \rightarrow Pb(s) + Mg^{2+}(aq)$

$E^0_{cell} = -0.1262\ V - (-2.372\ V)$

$E^0_{cell} = +2.246\ V$

$E^0_{cell} > 0$; **spontaneous**

**7.** $2Mn^{2+}(aq) + 8H_2O(l) + 10Hg^{2+}(aq) \rightarrow 2MnO_4^-(aq) + 16H^+(aq) + 5Hg_2^{2+}(aq)$

$E^0_{cell} = 0.920\ V - (+1.507\ V)$

$E^0_{cell} = -0.587\ V$

$E^0_{cell} < 0$; **not spontaneous**

**8.** $2SO_4^{2-}(aq) + Co^{2+}(aq) \rightarrow Co(s) + S_2O_8^{2-}(aq)$

$E^0_{cell} = -0.28\ V - (+2.010\ V)$

$E^0_{cell} = -2.29\ V$

$E^0_{cell} < 0$; **not spontaneous**

**9. Challenge** Using **Table 1**, write the equation and determine the cell voltage ($E^0$) for the following cell. Is the reaction spontaneous?

$Al|Al^{3+}||Hg^{2+}|Hg_2^{2+}$

$2Al(s) + 6Hg^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Hg_2^{2+}(aq)$

$E^0_{cell} = 0.920\ V - (-1.662\ V) = +2.582\ V$

**The reaction is spontaneous.**

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# Check Your Progress

## Summary

- In a voltaic cell, oxidation and reduction take place at electrodes separated from each other.
- The standard potential of a half-cell reaction is its voltage when paired with a standard hydrogen electrode under standard conditions.
- The reduction potential of a half-cell is negative if it undergoes oxidation when connected to a standard hydrogen electrode. The reduction potential of a half-cell is positive if it undergoes reduction when connected to a standard hydrogen electrode.
- The standard potential of a voltaic cell is the difference between the standard reduction potentials of the half-cell reactions.

## Demonstrate Understanding

10. **Describe** the conditions under which a redox reaction causes an electric current to flow through a wire.
11. **Identify** the components of a voltaic cell. Explain the role of each component in the operation of the cell.
12. **Write** the balanced equation for the spontaneous cell reaction that occurs in a cell with these reduction half-reactions.
    a. $Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ and $Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$
    b. $Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$ and $2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$
    c. $Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$ and $Fe^{3+}(aq) + 3e^- \rightleftharpoons Fe(s)$
    d. $PbI_2(s) + 2e^- \rightleftharpoons Pb(s) + 2I^-(aq)$ and $Pt^{2+}(aq) + 2e^- \rightleftharpoons Pt(s)$
13. **Determine** the standard potential for electrochemical cells in which each equation represents the overall cell reaction. Identify the reactions as spontaneous or nonspontaneous as written.
    a. $2Al^{3+}(aq) + 3Cu(s) \rightarrow 3Cu^{2+}(aq) + 2Al(s)$
    b. $Hg^{2+}(aq) + 2Cu^+(aq) \rightarrow 2Cu^{2+}(aq) + Hg(l)$
    c. $Cd(s) + 2NO_3^-(aq) + 4H^+(aq) \rightarrow Cd^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$
14. **Design** a concept map for Lesson 1, starting with the term *electrochemical cell*. Incorporate all the new vocabulary terms in your map.

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# Check Your Progress

## Summary

- In a voltaic cell, oxidation and reduction take place at electrodes separated from each other.

## Demonstrate Understanding

10. **Describe** the conditions under which a redox reaction causes an electric current to flow through a wire.
11. **Identify** the components of a voltaic cell. Explain the role of each component in the operation of the cell.

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12. **Write** the balanced equation for the spontaneous cell reaction that occurs in a cell with these reduction half-reactions.

a. $Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ and $Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$

b. $Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$ and $2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$

c. $Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$ and $Fe^{3+}(aq) + 3e^- \rightleftharpoons Fe(s)$

d. $PbI_2(s) + 2e^- \rightleftharpoons Pb(s) + 2I^-(aq)$ and $Pt^{2+}(aq) + 2e^- \rightleftharpoons Pt(s)$

$$Ag^+(aq) + e^- \rightarrow Ag(s)$$
$$E^\circ_{red} = +0.799\ \text{V}$$
$$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$$
$$E^\circ_{red} = -0.250\ \text{V}$$

$$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$$
$$E^\circ_{red} = -0.1375\ \text{V}$$
$$Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$$
$$E^\circ_{red} = -0.036\ \text{V}$$

$$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$$
$$E^\circ_{red} = -2.372\ \text{V}$$
$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$
$$E^\circ_{red} = 0.000\ \text{V}$$

$$PbI_2(s) + 2e^- \rightarrow Pb(s) + 2I^-(aq)$$
$$E^\circ_{red} \approx -0.365\ \text{V}$$
$$Pt^{2+}(aq) + 2e^- \rightarrow Pt(s)$$
$$E^\circ_{red} = +1.18\ \text{V}$$

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13. **Determine** the standard potential for electrochemical cells in which each equation represents the overall cell reaction. Identify the reactions as spontaneous or nonspontaneous as written.

a. $2Al^{3+}(aq) + 3Cu(s) \rightarrow 3Cu^{2+}(aq) + 2Al(s)$

b. $Hg^{2+}(aq) + 2Cu^{+}(aq) \rightarrow 2Cu^{2+}(aq) + Hg(l)$

c. $Cd(s) + 2NO_3^{-}(aq) + 4H^{+}(aq) \rightarrow Cd^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$

$$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$$

$$E^\circ_{\text{red}} = -1.662 \text{ V}$$

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

$$E^\circ_{\text{red}} = +0.3419 \text{ V}$$

$$Hg^{2+}(aq) + 2e^- \rightarrow Hg(l)$$

$$E^\circ_{\text{red}} = +0.851 \text{ V}$$

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

$$E^\circ_{\text{red}} = +0.153 \text{ V}$$

$$NO_3^-(aq) + 2H^+(aq) + e^- \rightarrow NO_2(g) + H_2O(l)$$

$$E^\circ_{\text{red}} = +0.775 \text{ V}$$

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

$$E^\circ_{\text{red}} = -0.4030 \text{ V}$$

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## Lesson 1 Check Your Progress

*page 632*

**10. Describe** the conditions under which a redox reaction causes an electric current to flow through a wire.

**An electrochemical cell in which an oxidation half-reaction and a reduction half-reaction are connected by a salt bridge and a conducting wire results in a flow of electrons (electric current) through the conducting wire.**

**11. Identify** the components of a voltaic cell. Explain the role of each component in the operation of the cell.

**A voltaic cell consists of an anode, a cathode, a salt bridge, and a connecting wire between the two electrodes. Oxidation takes place at the anode, reduction takes place at the cathode, the salt bridge allows movement of ions from one solution to the other, and the wire allows the passage of electrons from the anode to the cathode.**

**12. Write** the balanced equation for the spontaneous reaction that occurs in a cell with these reduction half-reactions.

**a.** $Ag^+(aq) + e^- \rightarrow Ag(s)$ and $Ni^{2+} + 2e^- \rightarrow Ni(s)$

**$2Ag^+ + Ni \rightarrow 2Ag + Ni^{2+}$**

**b.** $Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$ and $2H^+(aq) + 2e^- \rightarrow H_2(g)$

**$Mg + 2H^+ \rightarrow Mg^{2+} + H_2$**

**c.** $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$ and $Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$

**$2Fe^{3+}(aq) + 3Sn(s) \rightarrow 2Fe(s) + 3Sn^{2+}(aq)$**

**d.** $PbI_2(s) + 2e^- \rightarrow Pb(s) + 2I^-(aq)$ and $Pt^{2+}(aq) + 2e^- \rightarrow Pt(s)$

**$Pb(s) + 2I^-(aq) + Pt^{2+}(aq) \rightarrow PbI_2(s) + Pt(s)$**

**13. Determine** the standard potential for electrochemical cells in which each equation represents the overall cell reaction. Identify the reactions as spontaneous or nonspontaneous as written.

**a.** $2Al^{3+}(aq) + 3Cu(s) \rightarrow 3Cu^{2+}(aq) + 2Al(s)$

**$E^0_{cell} = -1.662\ V - (+0.3419\ V) = -2.004\ V$**

**nonspontaneous**

**b.** $Hg^{2+}(aq) + 2Cu^+(aq) \rightarrow 2Cu^{2+}(aq) + Hg(l)$

**$E^0_{cell} = +0.851\ V - (+0.153\ V) = +0.698\ V$**

**spontaneous**

**c.** $Cd(s) + 2NO_3^-(aq) + 4H^+(aq) \rightarrow Cd^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$

**$E^0_{cell} = +0.775\ V - (-0.4030\ V) = +1.178\ V$**

**spontaneous**

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| | | | |
|---|---|---|---|
| Q26 | Compare between electrolytic cell and voltaic cell in terms of identifying where will reduction and oxidation processes take place, anode, cathode, direction of electron flow and current flow and spontaneity of the reaction occurring | figure 18+ textbook | 47 |
| Q27 | Describe how a spontaneous redox reaction of an electrochemical cell can be reversed | figure 18+ textbook | 47,48 |

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# Electrolysis

- **Reversing Redox Reaction**

- The use of electrical energy to bring about a chemical reaction is called **electrolysis**.
- An electrochemical cell in which electrolysis occurs is called an **electrolytic cell**.
- For example, when a secondary battery is recharged, it is acting as an electrolytic cell.

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# Reversing Redox Reactions

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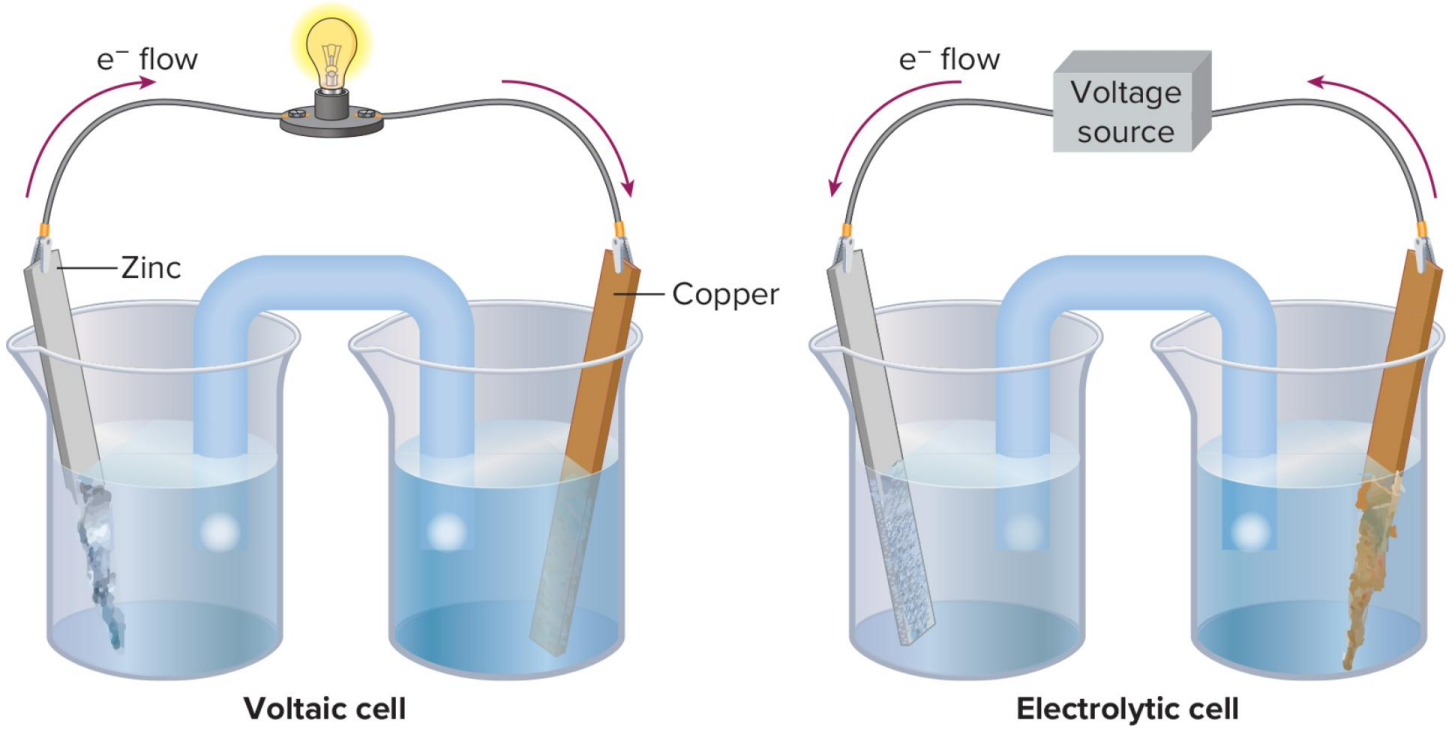
- The use of electrical energy to bring about a chemical reaction is called **electrolysis.**
- An electrochemical cell in which electrolysis occurs is called an **electrolytic cell**.



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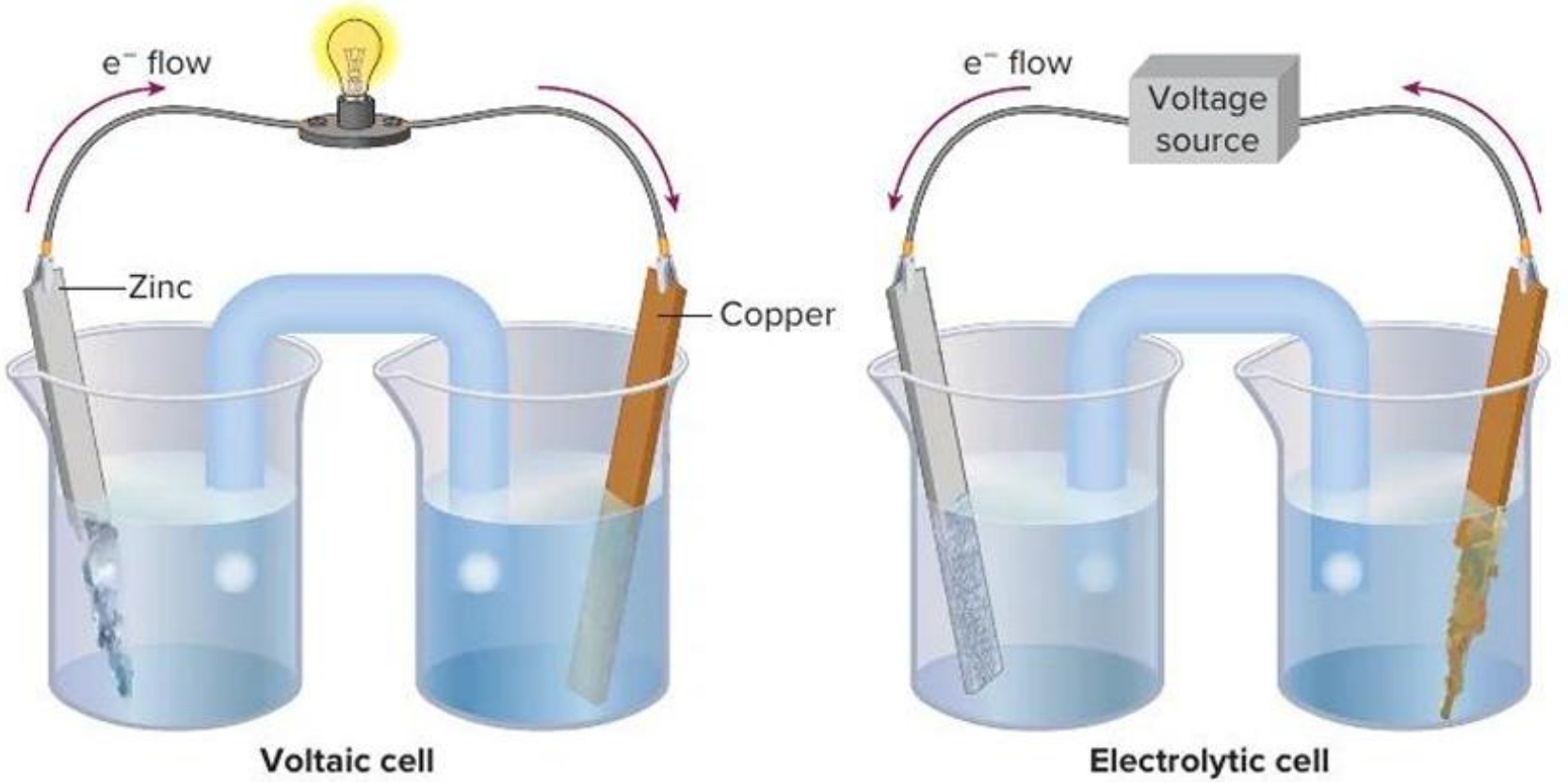


**Figure 18** The zinc-copper electrochemical cell can be a voltaic cell or an electrolytic cell.

**Infer** *In each electrochemical cell, which metal is oxidized? What is reduced?*

# Quiz

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1. Which is the use of electrical energy to bring about a chemical reaction?

1. reduction
2. galvanization
3. electrolysis **CORRECT**
4. oxidation

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# Quiz

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2. Which is an electrochemical cell in which electrolysis occurs?

A a voltaic cell

C a half-cell

B a fuel cell

D an electrolytic cell

CORRECT

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# Differences between voltaic cells and electrolytic cells

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| Voltaic Cell | Electrolytic Cell |
|---|---|
| Chemical energy → Electrical energy | Electrical energy → Chemical energy |
| Voltaic cells convert chemical energy to electrical energy as a result of a spontaneous redox reaction. | Electrolytic cells use electrical energy to drive a nonspontaneous reaction. |
| In these electrochemical cells, the anode is negatively charged, and the cathode is positively charged. | These cells feature a positively charged anode and a negatively charged cathode. |
| The electrons originate from the species that undergoes oxidation. | Electrons originate from an external source (such as a battery). |

***NOTE:*** ***Cathode is ALWAYS where Reduction*** ***happens regardless if the cell is voltaic or Electrolytic.***

***Anode is Always where oxidation*** ***happens regardless if the cell is Voltaic or Electrolytic.***

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# Activity: concept check

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- What type of electrochemical cell is this?

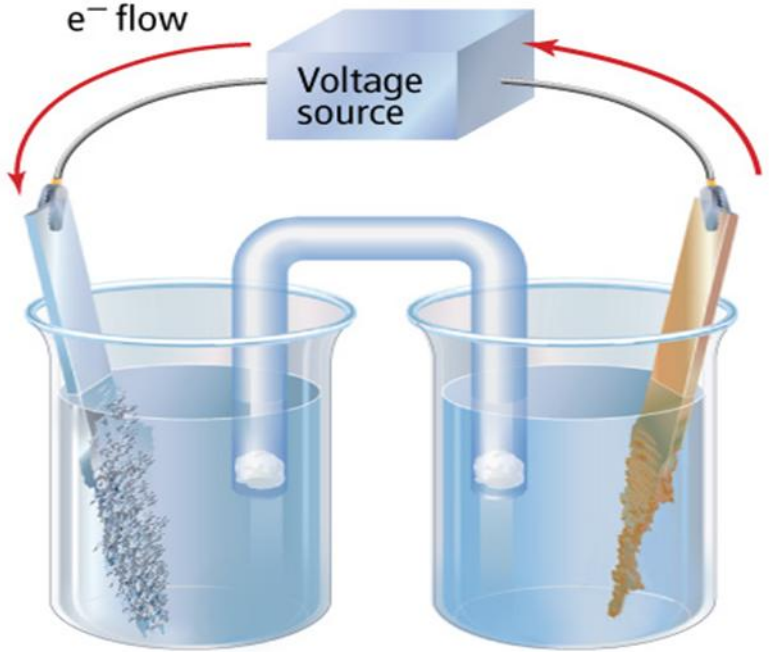
  Electrolytic cell

- What is the charge on anode and cathode?

  Anode is positive and cathode is negative

- Is the cell reaction spontaneous or nonspontaneous?

  Nonspontaneous




Electrolytic cell

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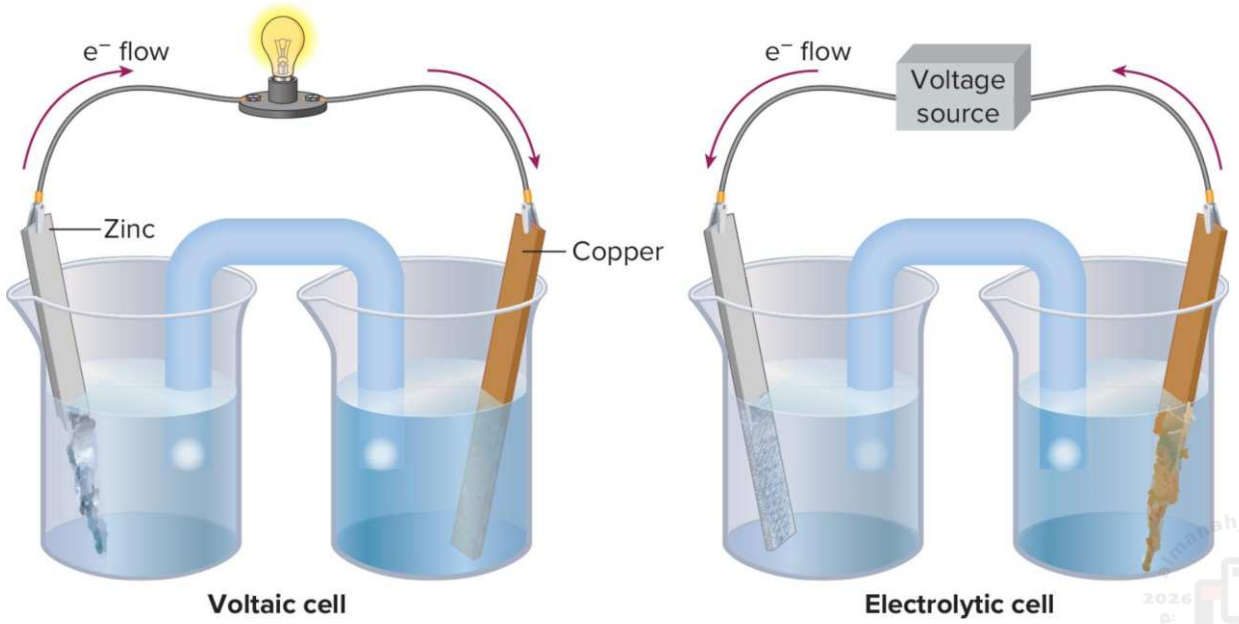
The reaction continues until the zinc strip is used up, and then the reaction stops. However, the cell can be regenerated if current is applied in the reverse direction using an external voltage source. The voltage source is required because the reverse reaction is nonspontaneous. If the voltage source is applied long enough, the cell will return to nearly its original strength.



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| | | | |
|---|---|---|---|
| Q28 | Write half-cell reactions while identifying the products of electrolysis of molten ionic compounds | figure 19+ textbook | 48 |
| Q29 | Write half-cell reactions while identifying the products of electrolysis of aqueous ionic compounds | figure 20+ textbook | 49 |

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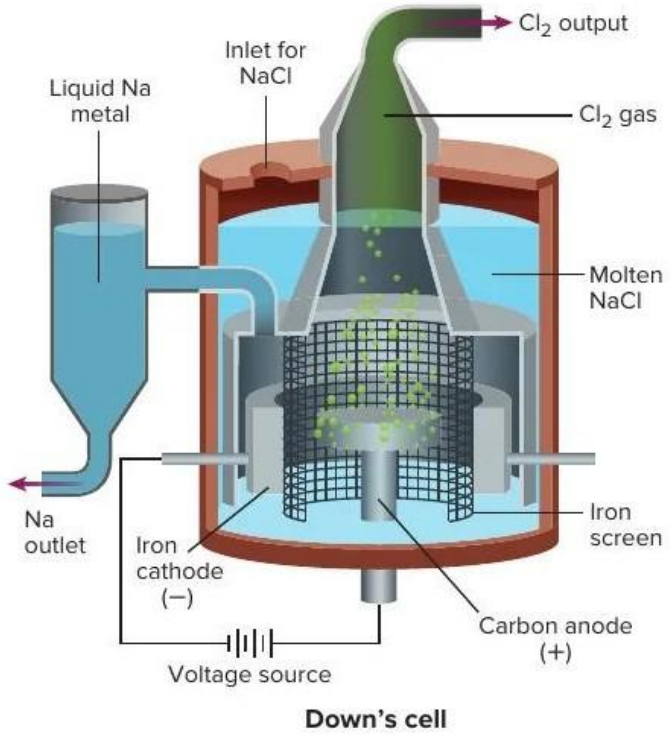
# Electrolysis of molten NaCl (Down's cell)




**Down's cell**

**Figure 19** In a Down's cell, electrons supplied by a generator are used to reduce sodium ions at the cathode. As electrons are removed from the anode, chloride ions are oxidized at the anode to chlorine gas.

- The electrolyte is the **molten NaCl** itself: its ions are free to move (solid NaCl does not conduct)

**Anode (+, carbon): oxidation**

$2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

**Cathode (−, iron): reduction**

$Na^+(l) + e^- \rightarrow Na(l)$

**Overall: $2Na^+(l) + 2Cl^-(l) \rightarrow 2Na(l) + Cl_2(g)$**

Products: **sodium metal** at the cathode and **chlorine gas** at the anode.

# Electrolysis of brine (aqueous NaCl)

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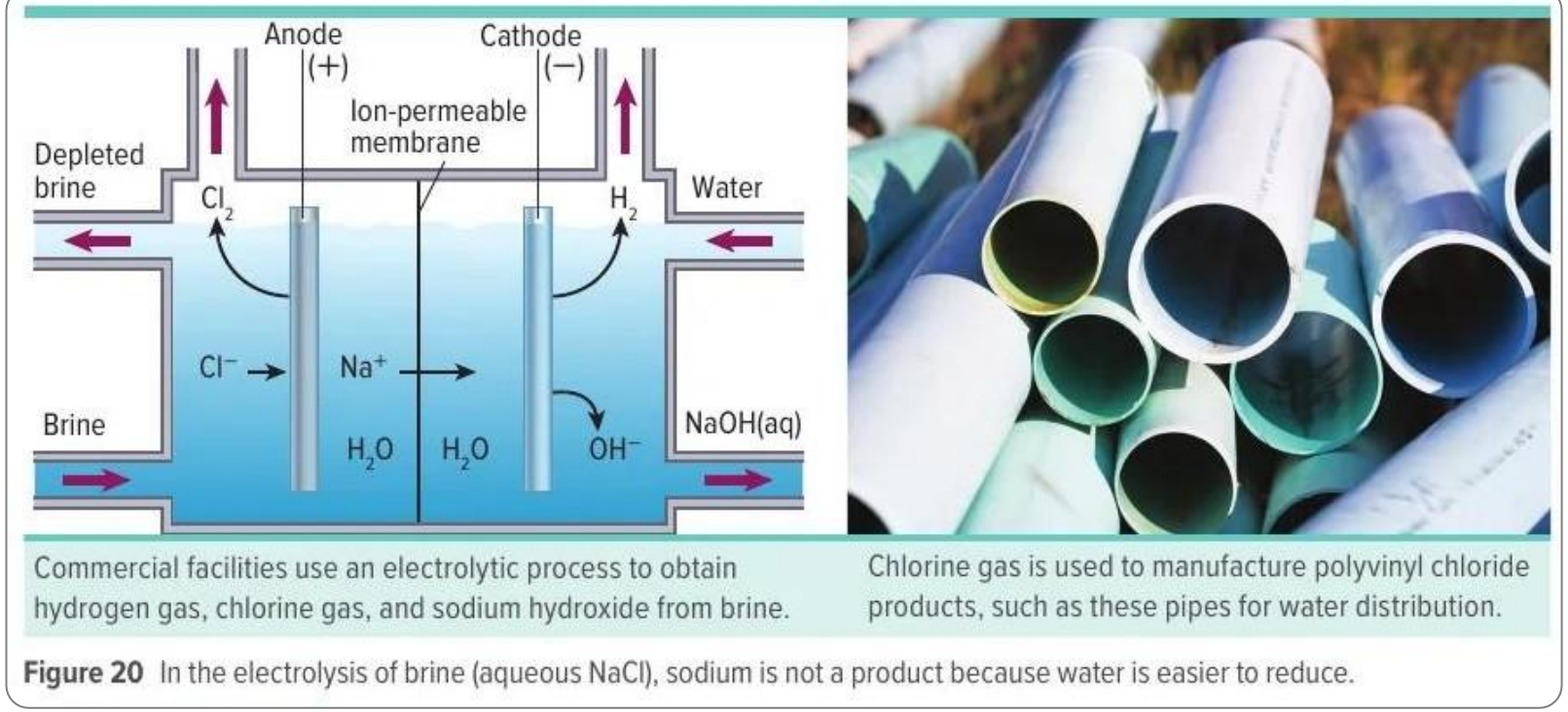
**Figure 20:** sodium is not a product because water is easier to reduce.




Commercial facilities use an electrolytic process to obtain hydrogen gas, chlorine gas, and sodium hydroxide from brine.

Chlorine gas is used to manufacture polyvinyl chloride products, such as these pipes for water distribution.

**Figure 20** In the electrolysis of brine (aqueous NaCl), sodium is not a product because water is easier to reduce.

**Anode: oxidation**

Occurs: $2Cl^- \rightarrow Cl_2 + 2e^-$ ($Cl^-$ concentration kept high)

Possible: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

**Cathode: reduction**

Possible: $Na^+ + e^- \rightarrow Na$ (does **not** occur)

Occurs: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

**$2H_2O + 2NaCl \rightarrow H_2 + Cl_2 + 2NaOH$** (overall)

Products: **$H_2$**, **$Cl_2$** and **NaOH(aq)**.

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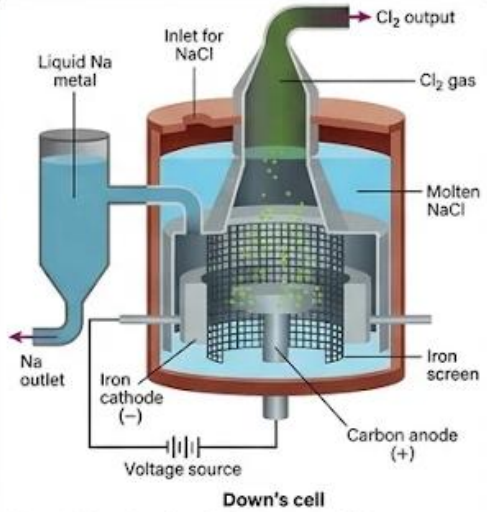
**Q1** What is produced at the cathode during the electrolysis of molten NaCl in a Down's cell?

ما الناتج المتكون عند المهبط (الكاثود) أثناء التحليل الكهربائي لمصهور NaCl في خلية داون؟




**Down's cell**

**Figure 19** In a Down's cell, electrons supplied by a generator are used to reduce sodium ions at the cathode. As electrons are removed from the anode, chloride ions are oxidized at the anode to chlorine gas.

| | |
|---|---|
| Chlorine gas | غاز الكلور |
| Sodium metal | فلز الصوديوم |
| Hydrogen gas | غاز الهيدروجين |
| Sodium hydroxide | هيدروكسيد الصوديوم |

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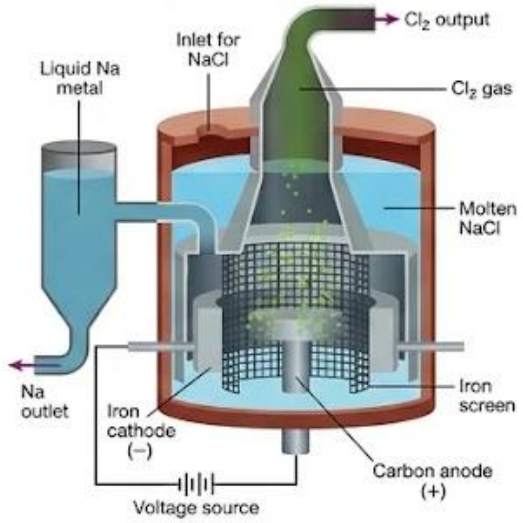
**Q5 Why must the NaCl electrolyte be molten in the Down's cell rather than solid?**

**Q5 لماذا يجب أن يكون إلكتروليت NaCl مصهورًا في خلية خلية داون بدلاً من أن يكون صلبًا؟**




**Down's cell**

**Figure 19** In a Down's cell, electrons supplied by a generator are used to reduce sodium ions at the cathode. As electrons are removed from the anode, chloride ions are oxidized at the anode to chlorine gas.

| | |
|---|---|
| Solid NaCl does not conduct electricity as its ions are not free to move | الصلب لا يوصل الكهرباء لأن أيوناته ليست حرة الحركة |
| Solid NaCl has a higher concentration of sodium ions | الصلب يحتوي على تركيز أعلى من أيونات الصوديوم |
| Molten NaCl has no chloride ions | المصهور لا يحتوي على أيونات الكلوريد |
| Solid NaCl prevents the formation of chlorine gas | الصلب يمنع تكون غاز الكلور |

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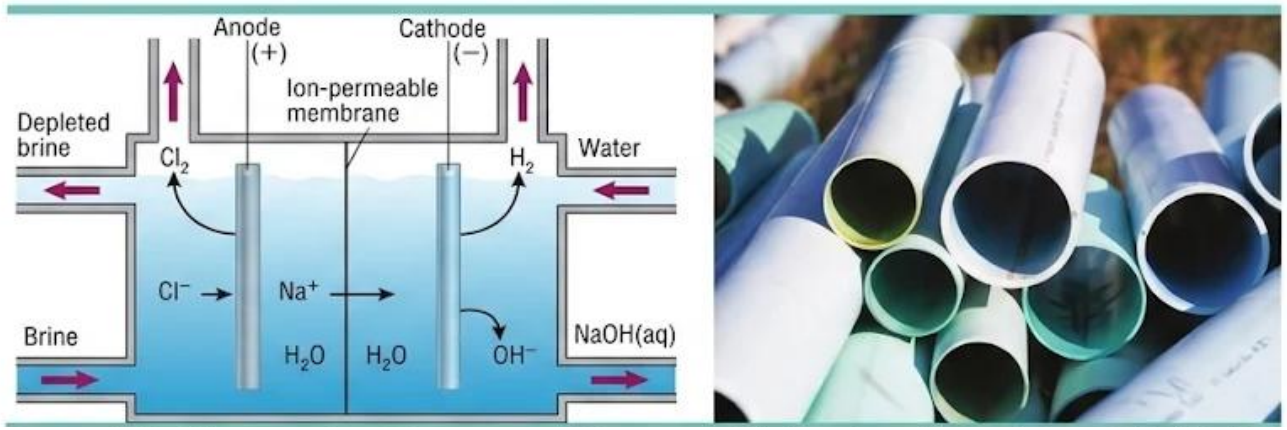
Q6 Which half-reaction occurs at the cathode during the electrolysis of brine?

Q6 ما التفاعل النصفي الذي يحدث عند المهبط (الكاثود) أثناء التحليل الكهربائي للماء المالح؟






| | |
|---|---|
| $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ | $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ ○ |
| $Na^+ + e^- \rightarrow Na$ | $Na^+ + e^- \rightarrow Na$ ○ |
| $2Cl^- \rightarrow Cl_2 + 2e^-$ | $2Cl^- \rightarrow Cl_2 + 2e^-$ ○ |
| $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ | $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ ○ |

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| | | | |
|---|---|---|---|
| Q30 | Define electroplating while describing how it works, identifying anode, cathode and electrolyte needed for an electrolytic cell in which a selected metal is to be plated on an object (car or spoon,... etc) | figure 22+ textbook | 51,52 |

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## Electroplating

Objects are electroplated when a uniform coating is deposited usually as a protective or decorative layer. The substance used is typically a metal. Electroplating with a metal such as silver is accomplished with a method similar to that used to refine copper. The object to be silver-plated is the cathode of an electrolytic cell. The anode is a silver bar or sheet, as shown in **Figure 22**.

At the anode, silver is oxidized to silver ions as electrons are removed by the power source. At the cathode, the silver ions are reduced to silver metal by electrons from the external power source. The silver forms a thin coating over the object being plated. Current passing through the cell must be carefully controlled in order to get a smooth, even metal coating.

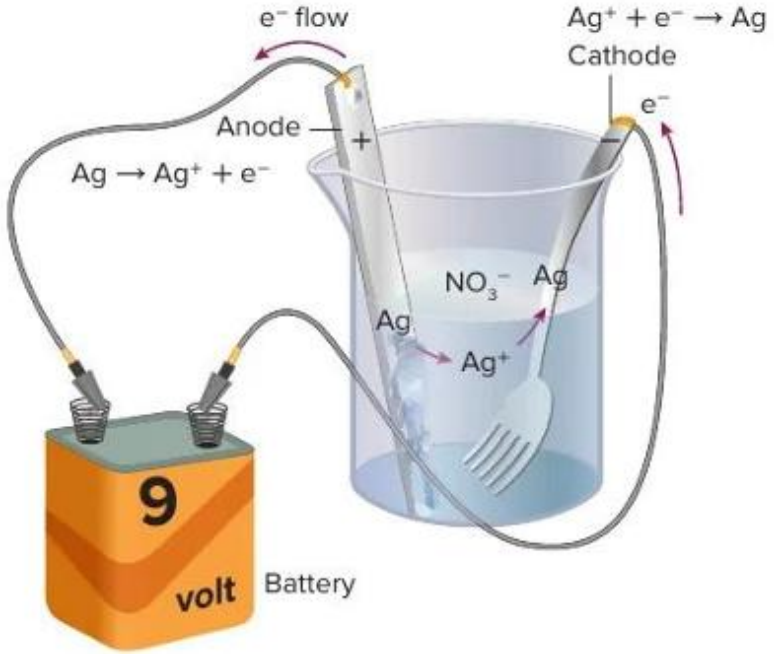


**Figure 22** Power is needed to oxidize silver at the anode and reduce silver at the cathode. In an electrolytic cell used for silver plating, the object to be plated is the cathode where silver ions in the electrolyte solution are reduced to silver metal and deposited on the object.

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# Extra Exam Practice on Q18 to 30

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In the voltaic cell shown below, where are the ions move from one side to another?

في الخلية الفولتية الموضحة ادناه، اين تتحرك الايونات من جهة إلى أخرى؟




| | |
|---|---|
| Through the silver strip | خلال قطب الفضة |
| Through the salt bridge | خلال القنطرة الملحية |
| Through the copper wire | خلال سلك النحاس |
| Through the zinc strip | خلال قطب الخارصين |

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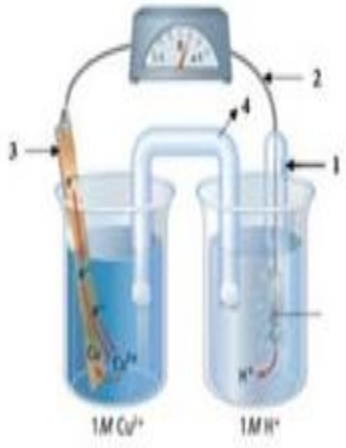
What number represents the part that prevents ions build up around the electrodes in the voltaic cell shown below?

ما الرقم الذي يُمثل الجزء الذي يمنع تراكم الأيونات حول القطبين في الخلية الفولتية الموضحة أدناه؟




| | |
|---|---|
| 1 | ○ |
| 2 | ○ |
| 3 | ○ |
| 4 | ○ |

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The following reaction takes place in a voltaic cell: يحدث التفاعل التالي في خلية فولتية:

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$

Which of the following takes place at the anode? أي مما يلي يحدث عند الأنود؟

- Reduction of $Cu^{2+}$ — اختزال $Cu^{2+}$
- Reduction of $Zn^{2+}$ — اختزال $Zn^{2+}$
- Oxidation of Cu — أكسدة Cu
- Oxidation of Zn — أكسدة Zn

In the voltaic cell shown below, where are electrons loosed? في الخلية الفولتية الموضحة أدناه، أين تفقد الإلكترونات؟




- At zinc strip — عند قطب الخارصين
- At salt bridge — عند القنطرة الملحية
- At both strips — عند كلا القطبين
- At copper strip — عند قطب النحاس

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When a solid copper wire is put into a solution of silver nitrate, a substance precipitates on the copper wire and blue solution forms. Which of the following is correct?

عند وضع سلك من النحاس الصلب في محلول نيترات الفضة، تترسب مادة على سلك النحاس ويتكون محلول أزرق اللون. أي مما يأتي صحيح؟

| المحلول الأزرق<br>The blue solution | نصف تفاعل الاختزال<br>Reduction - half-reaction | نصف تفاعل الأكسدة<br>Oxidation - half-reaction | الفلز الذي ترسبت على السلك<br>Metal deposited on the wire | |
|---|---|---|---|---|
| نيترات النحاس (II)<br>Copper(II) nitrate | $Ag^+ + e^- \longrightarrow Ag$ | $Cu \longrightarrow Cu^{2+} + 2e^-$ | Silver الفضة | 1 |
| نيترات الفضة<br>Silver nitrate | $Cu \longrightarrow Cu^{2+} + 2e^-$ | $Ag^+ + e^- \longrightarrow Ag$ | Copper النحاس | 2 |
| نيترات النحاس (II)<br>Copper(II) nitrate | $Cu^{2+} + 2e^- \longrightarrow Cu$ | $Ag + e^- \longrightarrow Ag^+$ | Silver الفضة | 3 |
| نيترات الفضة<br>Silver nitrate | $Ag + e^- \longrightarrow Ag^+$ | $Cu^{2+} + 2e^- \longrightarrow Cu$ | Copper النحاس | 4 |

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$Cu^{2+} + 2e^- \rightleftarrows Cu \qquad E^0(V) = +0.3419$

$Zn^{2+} + 2e^- \rightleftarrows Zn \qquad E^0(V) = -0.7618$




- The standard reduction potential ($E^0$) of the standard hydrogen electrode is defined as 1.104 V — جهد الاختزال القياسي ($E^0$) لقطب الهيدروجين يساوي 1.104 V

- (●) The standard reduction potential ($E^0$) of the standard hydrogen electrode is defined as 0.000 V — جهد الاختزال القياسي ($E^0$) لقطب الهيدروجين يساوي 0.000 V

- It will be a cathode when connected to $Cu|Cu^{2+}$ electrode — يكون كاثود عند توصيله مع قطب $Cu|Cu^{2+}$

- It will be an anode when connected to $Zn \mid Zn^{2+}$ electrode — يكون أنودًا عند توصيله مع قطب $Zn \mid Zn^{2+}$

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A voltaic cell consists of the following half-reduction reactions, which of the following is correct?

تتكون خلية فولتية من تفاعلات الاختزال النصفية التالية، أي مما يأتي صحيح؟

$Ag^+ + e^- \rightleftharpoons Ag$ , $E^0$ (V) = +0.7996

$Cr^{3+} + 3e^- \rightleftharpoons Cr$ , $E^0$ (V) = −0.744

| | الأنود<br>Anode | الكاثود<br>Cathode | جهد الخلية ( $E^0$ )<br>Electrochemical Cell Potential ( $E^0$ ) | نوع التفاعل<br>Reaction Type |
|---|---|---|---|---|
| A | Cr | Ag | + 1.54 V | تلقائي<br>Spontaneous |
| B | Ag | Cr | + 0.0556 V | تلقائي<br>Spontaneous |
| C | Cr | Ag | − 1.54 V | غير تلقائي<br>Nonspontaneous |
| D | Ag | Cr | − 0.0556 V | غير تلقائي<br>Nonspontaneous |

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Easy Chemist

Which of the following represents the cell notation for the voltaic cell consisting of nickel and calcium electrodes?

أي مما يلي يمثل ترميز الخلية الفولتية التي تتكون من أقطاب النيكل والكالسيوم؟

| التفاعل النصفي Half-Reaction | $E^{\circ}(V)$ |
|---|---|
| $Ni^{2+} + 2e^{-} \rightarrow Ni$ | - 0.257 |
| $Ca^{2+} + 2e^{-} \rightarrow Ca$ | -2.868 |

- $Ca^{2+}|Ca\|Ni|Ni^{2+}$ ○
- $Ni^{2+}|Ni\|Ca^{2+}|Ca$ ○
- $Ni|Ni^{2+}\|Ca^{2+}|Ca$ ○
- $Ca|Ca^{2+}\|Ni^{2+}|Ni$ ○

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In the redox reaction below. **What represents the oxidation half-reaction and the reduction half-reaction?**

في تفاعل الأكسدة والاختزال أدناه. **ما الذي يُمثل نصفي تفاعل الأكسدة والاختزال؟**

$$2Br^- + Cl_2 \rightarrow Br_2 + 2Cl^-$$

| نصف تفاعل الاختزال<br>Reduction half-reaction | نصف تفاعل الأكسدة<br>Oxidation Half-Reaction | |
|---|---|---|
| $2Br^- + 2e^- \longrightarrow Br_2$ | $Cl_2 \longrightarrow 2Cl^- + 2e^-$ | 1 |
| $Cl_2 \longrightarrow 2Cl^- + 2e^-$ | $2Br^- + 2e^- \longrightarrow Br_2$ | 2 |
| $2Br^- \longrightarrow Br_2 + 2e^-$ | $Cl_2 + 2e^- \longrightarrow 2Cl^-$ | 3 |
| $Cl_2 + 2e^- \longrightarrow 2Cl^-$ | $2Br^- \longrightarrow Br_2 + 2e^-$ | 4 |

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| Which of the following is true for the two cells below? | أي مما يلي صحيح فيما يتعلق بالخليتين أدناه؟ |
|---|---|
|   | |
| A- The Cell number (1) converts the electrical energy into chemical energy<br>B- A spontaneous redox reaction occurs in the cell number (1)<br>C- The cell number (2) converts chemical energy into electrical energy<br>D- A spontaneous redox reaction occurs in the cell number (2) | A- تحوّل الخلية رقم (1) الطاقة الكهربائية إلى طاقة كيميائية<br>B- يحدث في الخلية رقم (1) تفاعل أكسدة واختزال تلقائي<br>C- تحوّل الخلية رقم (2) الطاقة الكيميائية إلى طاقة كهربائية<br>D- يحدث في الخلية رقم (2) تفاعل أكسدة واختزال تلقائي |

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1



2

- ○ In cell 1 the reaction is spontaneous
  في الخلية 1 يكون التفاعل تلقائيا
- ◉ In cell 2 the reaction continues until the zinc strip is used up and then the reaction stops
  في الخلية 2 يستمر التفاعل حتى يستهلك قطب الخارصين ثم يتوقف التفاعل
- ○ 2 is an electrolytic cell and coverts chemical energy to electrical energy
  الخلية 2 إلكتروليتية وتحول الطاقة الكيميائية إلى طاقة كهربائية

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What is the similarity between the two cells below?

ما وجه الشبه بين الخليتين أدناه؟




- A nonspontaneous reaction occurs — حدوث تفاعل غير تلقائي
- An oxidation-reduction reaction occurs — حدوث تفاعل أكسدة واختزال
- A neutralization reaction occurs — حدوث تفاعل تعادل
- A spontaneous reaction occurs — حدوث تفاعل تلقائي

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| Which of the following is true for the two cells below? | أي مما يلي صحيح فيما يتعلق بالخليتين أدناه؟ |
|---|---|
|   | |

| | |
|---|---|
| In cell number (1), the electrons flow from Cu to Zn | في الخلية رقم (1) تتدفق الإلكترونات من Cu إلى Zn |
| In cell number (1), the anode produces electrons from Zn | في الخلية رقم (1) يُنتج الأنود الإلكترونات من Zn |
| In cell number (2), the anode produces electrons from Zn | في الخلية رقم (2) يُنتج الأنود الإلكترونات من Zn |
| In cell number (2), the electrons flow from Zn to Cu | في الخلية رقم (2) تتدفق الإلكترونات من Zn إلى Cu |

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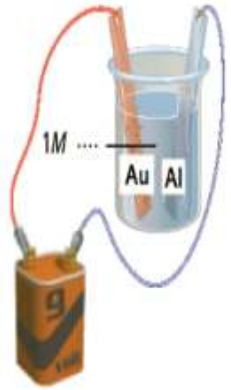
a layer of gold, as in the figure below. Which of the following is **correct**?

ادناه. أي مما يأتي **صحيح**؟




| | التغير في كتلة صفيحة الألمنيوم Change in mass of aluminum sheet | تفاعل الأنود Anode reaction | تفاعل الكاثود Cathode reaction | الإلكتروليت المستخدم Used electrolyte |
|---|---|---|---|---|
| 1 | تزداد Increases | $Au \longrightarrow Au^{3+} + 3e^-$ | $Au^{3+} + 3e^- \longrightarrow Au$ | يحتوي على $Au^{3+}$ Contains $Au^{3+}$ |
| 2 | تقل Decreases | $Al \longrightarrow Al^{3+} + 3e^-$ | $Al^{3+} + 3e^- \longrightarrow Al$ | يحتوي على $Al^{3+}$ Contains $Al^{3+}$ |
| 3 | تزداد Increases | $Au \longrightarrow Au^{3+} + 3e^-$ | $Al^{3+} + 3e^- \longrightarrow Al$ | يحتوي على $Au^{3+}$ Contains $Au^{3+}$ |
| 4 | تقل Decreases | $Au^{3+} + 3e^- \longrightarrow Au$ | $Al \longrightarrow Al^{3+} + 3e^-$ | يحتوي على $Al^{3+}$ Contains $Al^{3+}$ |

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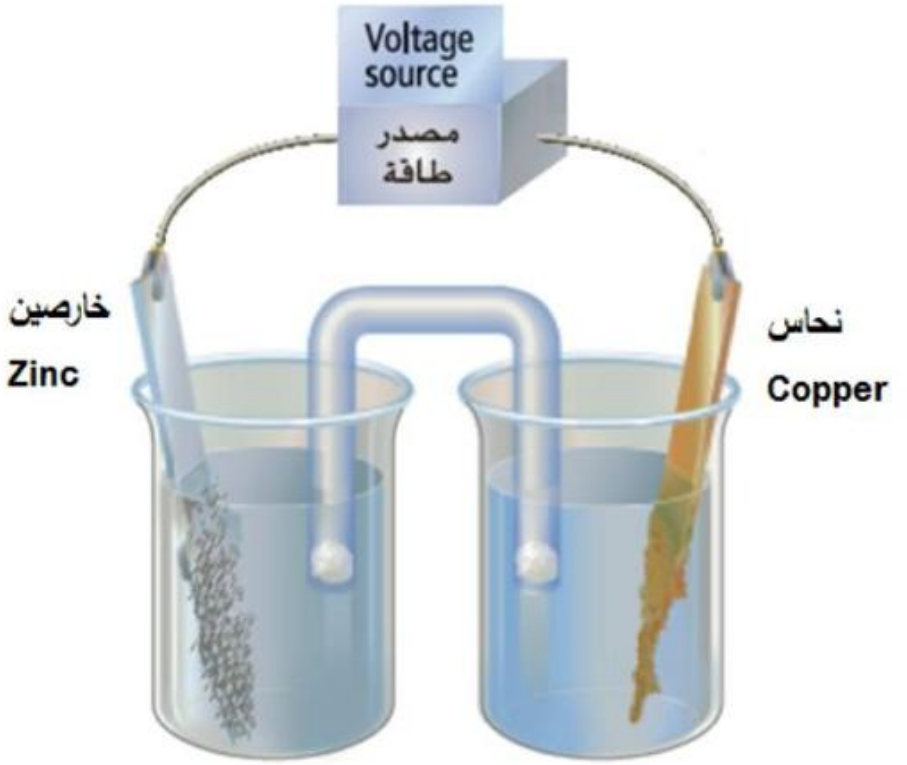
Regarding the cell in the figure below. Which of the following is **correct**?

فيما يتعلق بالخلية الموضحة في الشكل أدناه.
اي مما يأتي **صحيح**؟




| | |
|---|---|
| The reaction continues until the zinc strip is used up, and then the reaction stops | يستمر التفاعل حتى يستهلك قطب الخارصين ثم يتوقف التفاعل |
| Oxidation takes place at zinc electrode | تحدث عملية أكسدة عند قطب الخارصين |
| The external power supply (voltage) works on cell regeneration and the cell returns to almost its original strength | يعمل مصدر الطاقة (الجهد الكهربائي) الخارجي على تجديد الخلية وتعود الخلية إلى قوتها الأصلية تقريبًا |

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a. The object to be plated is the anode — يكون الجسم المراد طلاؤه هو الأنود

b. The metal used for coating is the cathode — يكون الفلز المراد الطلاء به هو الكاثود

c. Silver is oxidized to silver ions at the cathode — تتأكسد الفضة إلى أيونات الفضة عند الكاثود

d. Silver is oxidized to silver ions at the anode — تتأكسد الفضة إلى أيونات الفضة عند الأنود

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) The figure below shows a picture of a key being electroplated with copper in an electrolytic cell. Which of the following statements is true?

a. copper is oxidized to $Cu^{2+}$ ions at the cathode

b. reduction of copper ions occurs on the anode electrode

c. the mass of the copper electrode increases

d. electrical energy is used to cause a redox reaction



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**Why does the cathode produce hydrogen during the electrolysis of brine but sodium during the electrolysis of molten NaCl?**

**لماذا ينتج المهبط الهيدروجين عند التحليل الكهربائي للمحلول الملحي، بينما ينتج الصوديوم عند التحليل الكهربائي لكلوريد الصوديوم المنصهر؟**

Choose one correct answer.

اختر إجابة واحدة صحيحة.

**A** Sodium ions are absent from brine.

لا توجد أيونات الصوديوم في المحلول الملحي.

**B** Chloride ions are reduced more easily than sodium ions.

تُختزل أيونات الكلوريد بسهولة أكبر من أيونات الصوديوم.

**C** Water is oxidized at the cathode in brine.

يتأكسد الماء عند المهبط في المحلول الملحي.

**D** Water is easier to reduce than sodium ions and is present in brine.

الماء أسهل اختزالًا من أيونات الصوديوم، وهو موجود في المحلول الملحي.

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# END OF EOT1 Chemistry G12 Advanced

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