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Inspire Chemistry
Module 19

"Electrochemistry"

Redox Reactions

Lesson 1: "Voltaic Cells"

الخلايا الفولتية

Learning Outcomes:

- **Describe** a way to obtain electrical energy from a redox reaction.
- **Identify** the parts of a voltaic cell and explain how each part operates.
- **Calculate** cell potentials and determine the spontaneity of redox reactions.

Part 1

Learning objectives:

Define Electrochemistry.

Describe an electrochemical cell while specifying its types.

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

BIG IDEA: THINK-PAIR0SHARE

What is **"Electrochemistry"**?

How do **voltaic cells** harness **energy from chemical reactions**?

redox

New Vocabulary

salt bridge

electrochemical cell

voltaic cell

half-cell

anode

cathode

reduction potential

standard hydrogen electrode

Review Vocabulary

oxidation: the complete or partial loss of electrons from a reacting substance; increases an atom's or ion's oxidation number

reduction: the complete or partial gain of electrons by the atoms of a substance; decreases the atom's or ion's oxidation number

Big Idea

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

Red
redox Reaction
ox

What are the 2 half reactions?

Half-reactions Two half-reactions make up this redox process:

$Zn \rightarrow Zn^{2+} + 2e^-$ (oxidation half-reaction: electrons lost)

$Cu^{2+} + 2e^- \rightarrow Cu$ (reduction half-reaction: electrons gained)

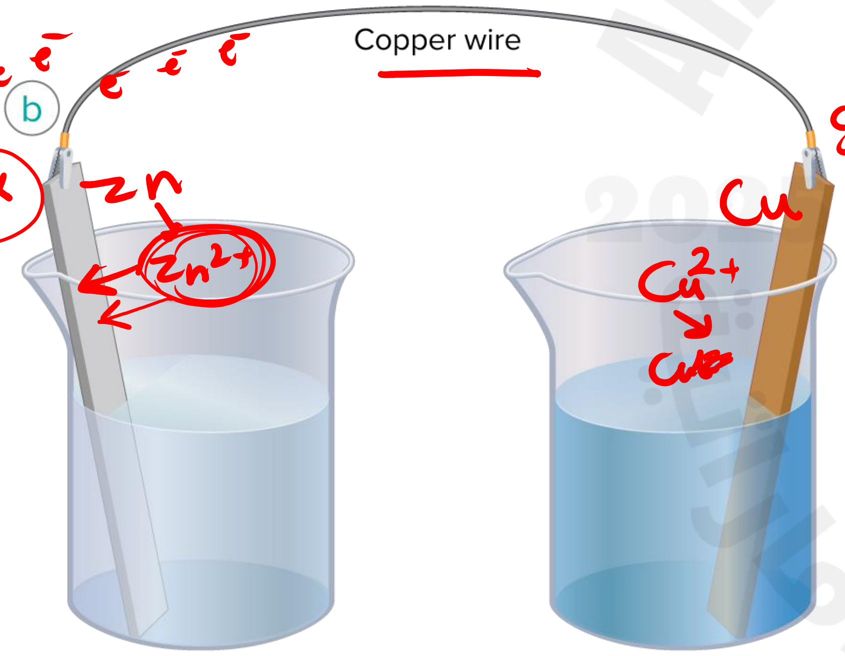
redox Reactions
Transfer of e^-

Can we use "Redox Reactions" to generate an **electrical current? How?**

Hint: Electricity is generated from the movement of electrons ________________

Redox in Electrochemistry

- **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.



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

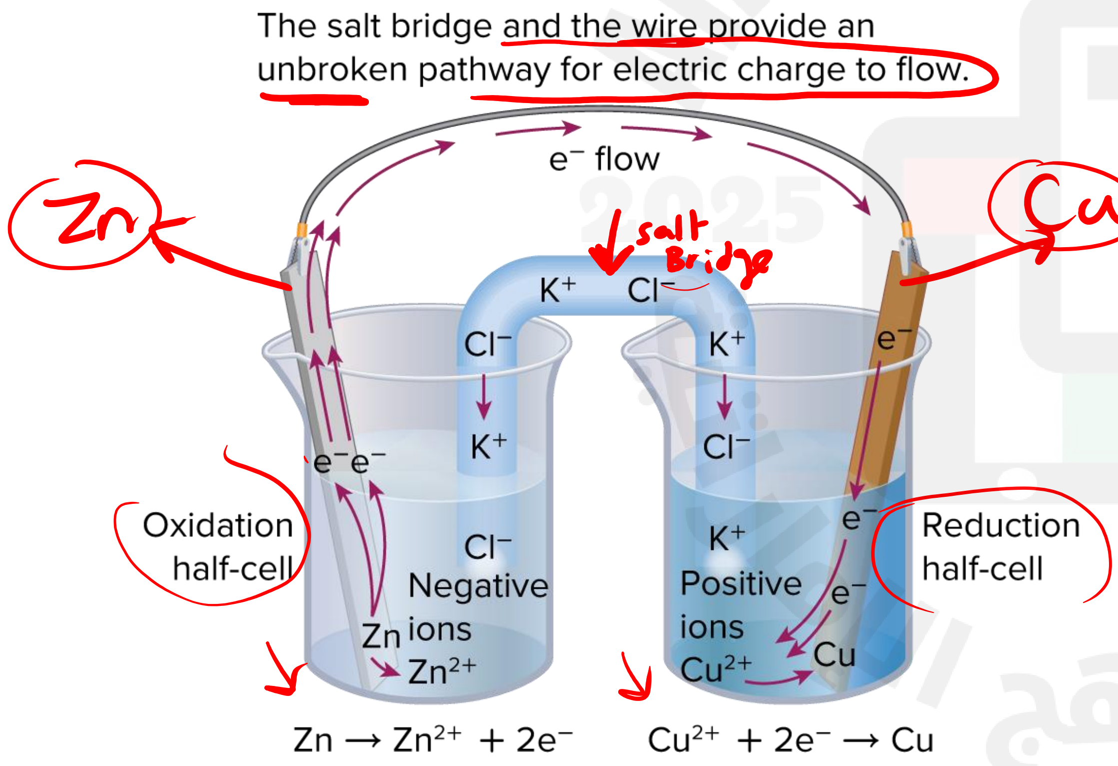
- An **electrochemical cell** is an apparatus that uses a redox reaction to produce electrical energy or uses electrical energy to cause a chemical reaction.

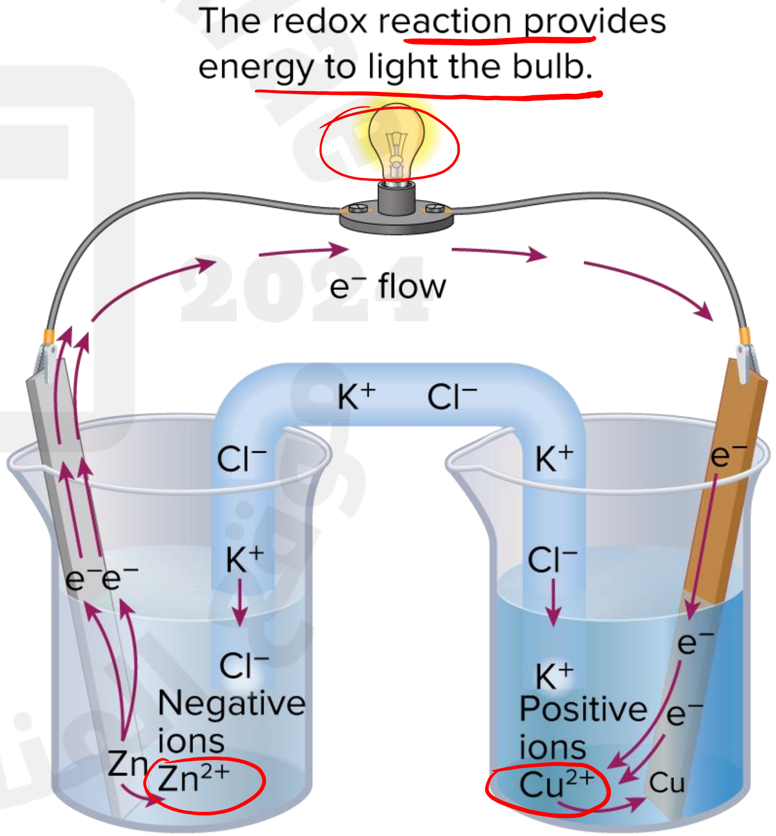
- Two types:
  - **1-Electrolytic and**
  - **2-Voltaic Cells.**

lesson

Redox in Electrochemistry

- 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.***

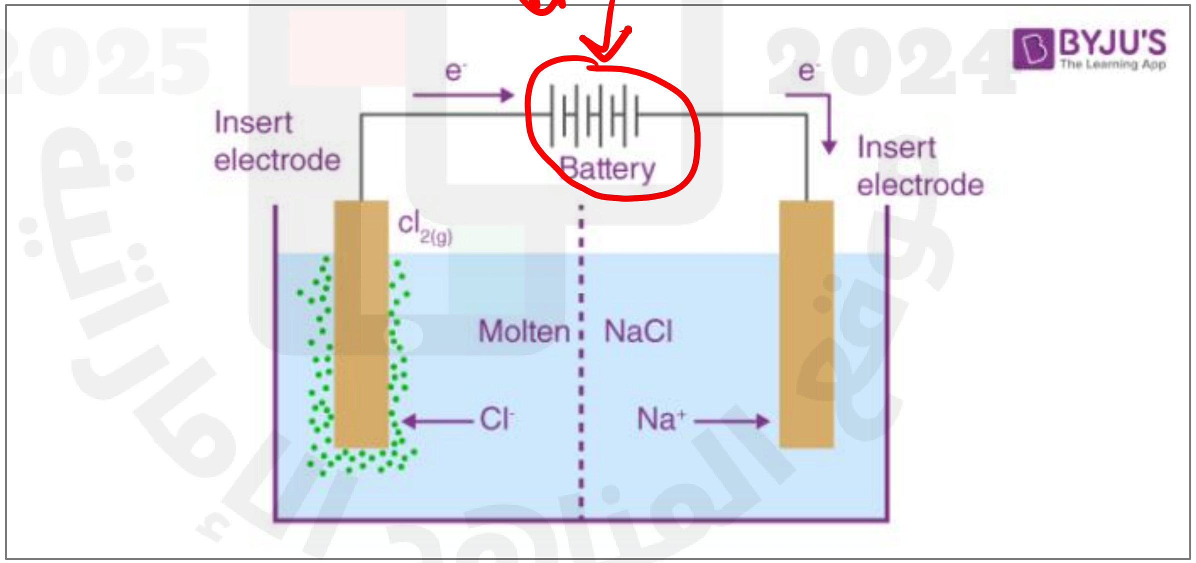




Redox in Electrochemistry

- **Electrolytic cell** is a type of electrochemical cell that converts electrical energy to chemical energy by a non-spontaneous redox reaction. It is an electrochemical cell where electrolysis occurs. Image below.
-

Electrolytic cells Needs an outer source of energy (Battery)



Quiz

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

Quiz

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

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

A voltaic cell

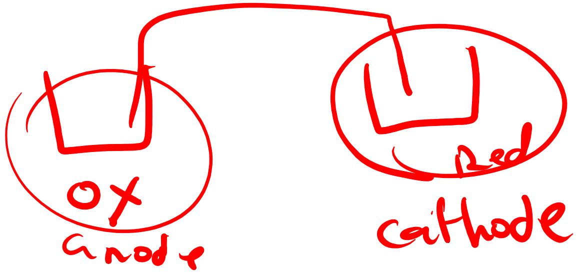
C half-cell

B salt cell

D oxidation cell

Chemistry of Voltaic Cells: Main points

- 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

cell

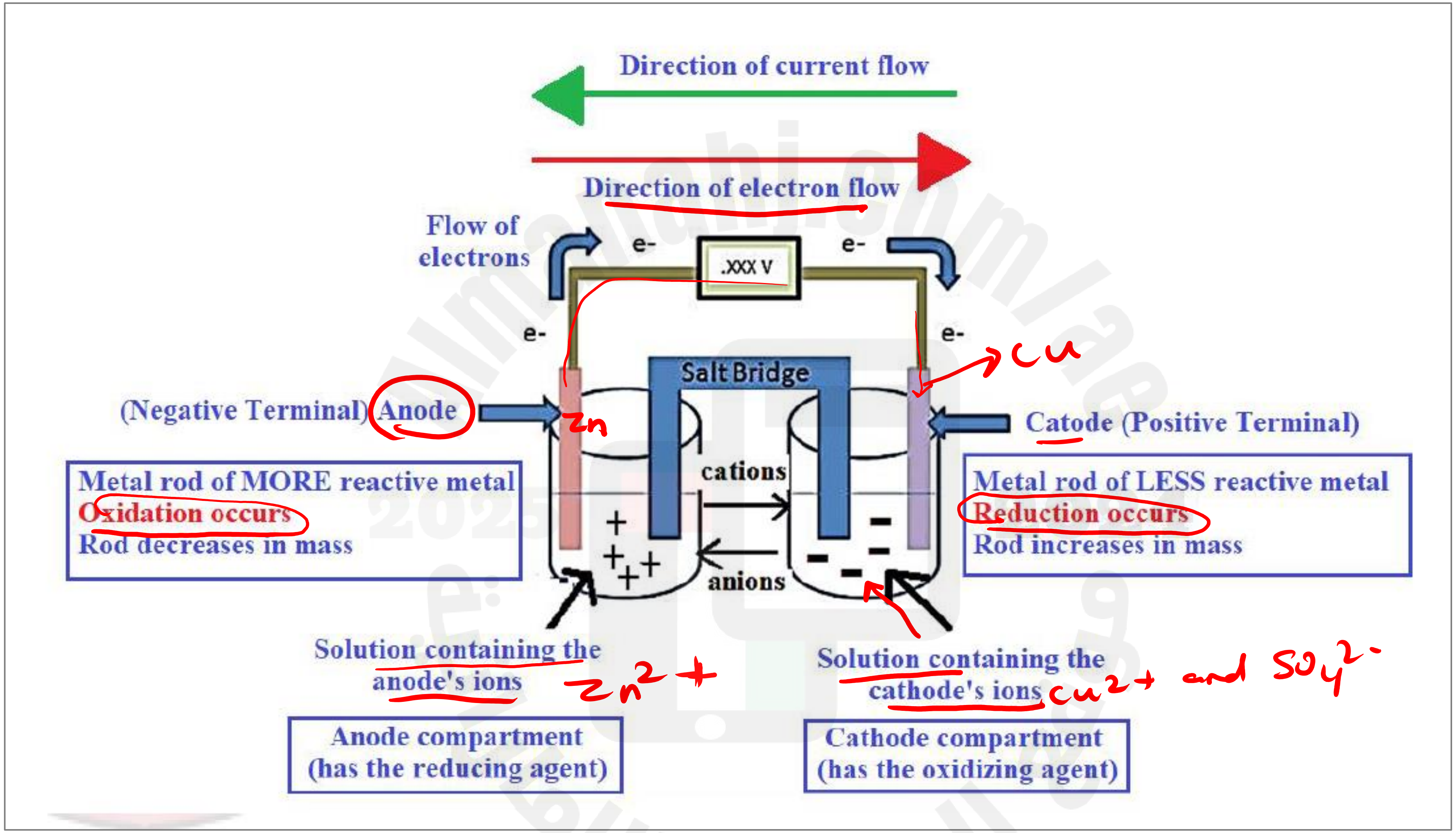
- Electrons flow through the wire from the oxidation-half reaction (anode) to the reduction-half reaction (cathode) while positive and negative ions move through the salt bridge

- *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.

Chemistry of Voltaic or Galvanic Cells: Summary



Part 2

Learning objectives:

Write the oxidation and reduction half-reactions occurring at cathode and anode for a voltaic cell.

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.

Define the reduction potential and **standard electrode potential (E°)**

Write the oxidation and reduction half-reactions occurring at cathode and anode for a voltaic cell ***(From The previous chapter)***

Previous video

Activity: The overall reaction occurs in a cell made up of an aluminium electrode and a tin electrode:

$$2Al(s) + 3Sn^{2+}(aq) \longrightarrow 2Al^{3+}(aq) + 3Sn(s)$$

Ox Red

Complete the reactions. Remember that the number of electrons released in the oxidation process must be the same as the number of electrons consumed in the reduction process. To obtain the overall cell reaction, we add the two half-reactions.

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Anode (ox)

$2Al \longrightarrow 2Al^{3+} + 6e^-$

0 → +6, −6 → 0

Cathode (Reduction)

$3Sn^{2+} + 6e^- \longrightarrow 3Sn$

+6, −6 → 0 → 0

Write the oxidation and reduction half-reactions occurring at cathode and anode for a voltaic cell ***(From The previous chapter)***

Activity: The overall reaction occurs in a cell made up of an aluminium electrode and a tin electrode:

$$2Al(s) + 3Sn^{2+}(aq) \longrightarrow 2Al^{3+}(aq) + 3Sn(s)$$

Complete the reactions. Remember that the number of electrons released in the oxidation process must be the same as the number of electrons consumed in the reduction process. To obtain the overall cell reaction, we add the two half-reactions.

Sn	$Sn^{2+}(aq)$	$Sn(s)$	Al	$Al(s)$	$Al^{3+}(aq)$	3	2	$6e^-$	$2e^-$

Red

Cathode: Sn element metal (No charge)

Reaction at the cathode:

$3\ Sn^{2+} + 6e^- \longrightarrow 3\ Sn$

ox

Anode: Al

Reaction at the anode:

$2\ Al \longrightarrow 2\ Al^{3+} + 6e^-$

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Page 3

Quiz

Which is the name of the two parts of an electrochemical cell, in which the separate oxidation and reduction reactions take place?

A voltaic cells

B half-cells

C anodes

D cathodes

Quiz

Which is the name of the two parts of an electrochemical cell, in which the separate oxidation and reduction reactions take place?

- A voltaic cells
- B half-cells
- C anodes
- D cathodes

Chemistry of Voltaic Cells

- Electric charge can flow between two points only **when a difference in electric potential (Voltage)** energy exists between the two points. (From your physics class ;)
- A volt is a unit used to measure cell potential—the force from the difference in electric potential energy between two electrodes.

HOW DO WE MEASURE THE VOLTAGE FROM A VOLATIC CELL?

In other words, how is the Cell Potential "E°" Calculated?

Let's learn some terminologies before we do that!

Calculating Electrochemical Cell Potentials

- Formula for Cell Potential

voltage

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

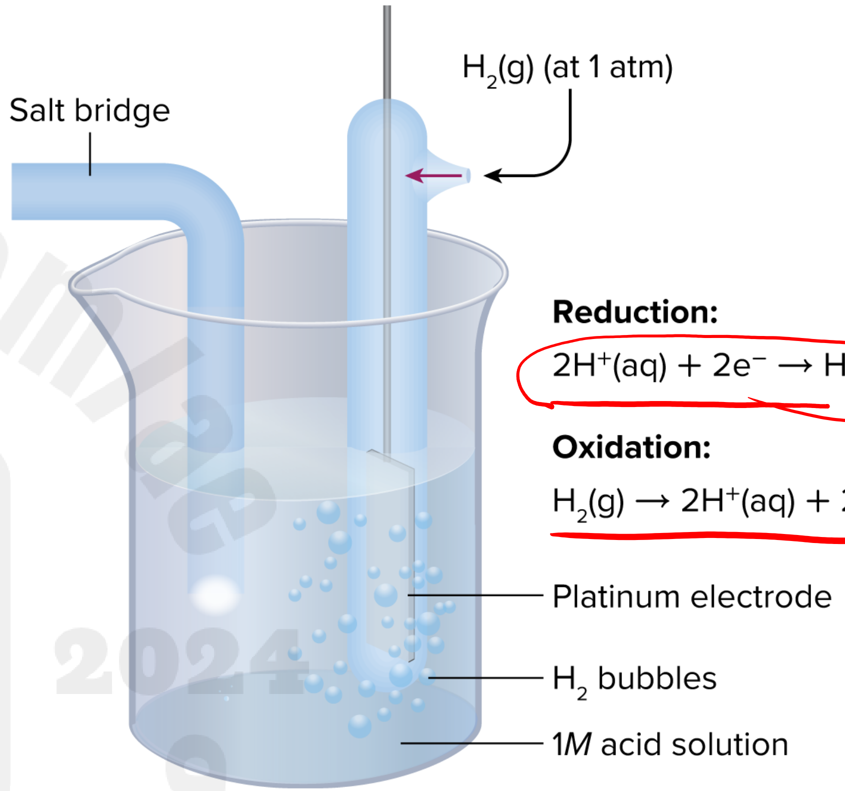
$$E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}}$$

- The standard potential of a cell (E^o) is the standard potential of the reduction half-cell minus the standard potential of the oxidation half-cell

Where do we find the standard potential of the cathode or the anode

Calculating Electrochemical Cell Potentials

- The tendency of a substance to gain electrons is its **reduction potential**.
- When two half-reactions are coupled, the **voltage generated corresponds to the difference in potential between the half-reactions. (ΔV or E° in chemistry)**
- **It is like the difference in voltage that you took in physics.**



Reference point

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

Quiz

Which is the tendency of a substance to **gain electrons**?

A reduction potential

C oxidation potential

B electrochemical potential

D reaction potential

Quiz

Which is the tendency of a substance to **gain electrons**?

A reduction potential

C oxidation potential

B electrochemical potential

D reaction potential

What is "Standard Electrode Potential" E^o ?

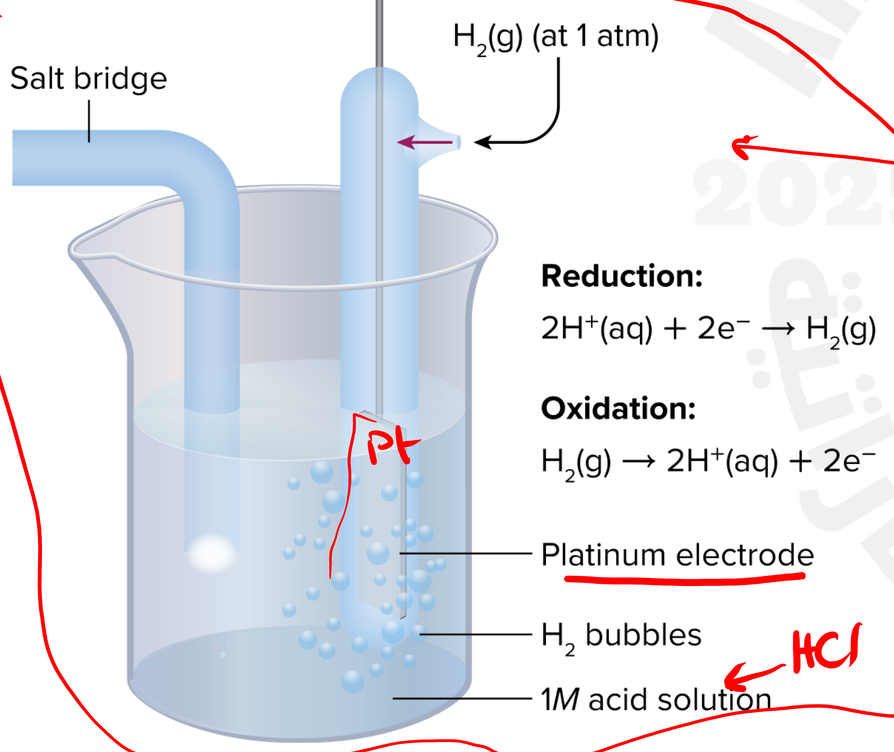
Standard electrode potential(E^o) is the potential or voltage developed by a metal or other material immersed in an electrolyte solution relative to the **potential of the hydrogen electrode which is zero.**

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

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

Calculating Electrochemical Cell Potentials

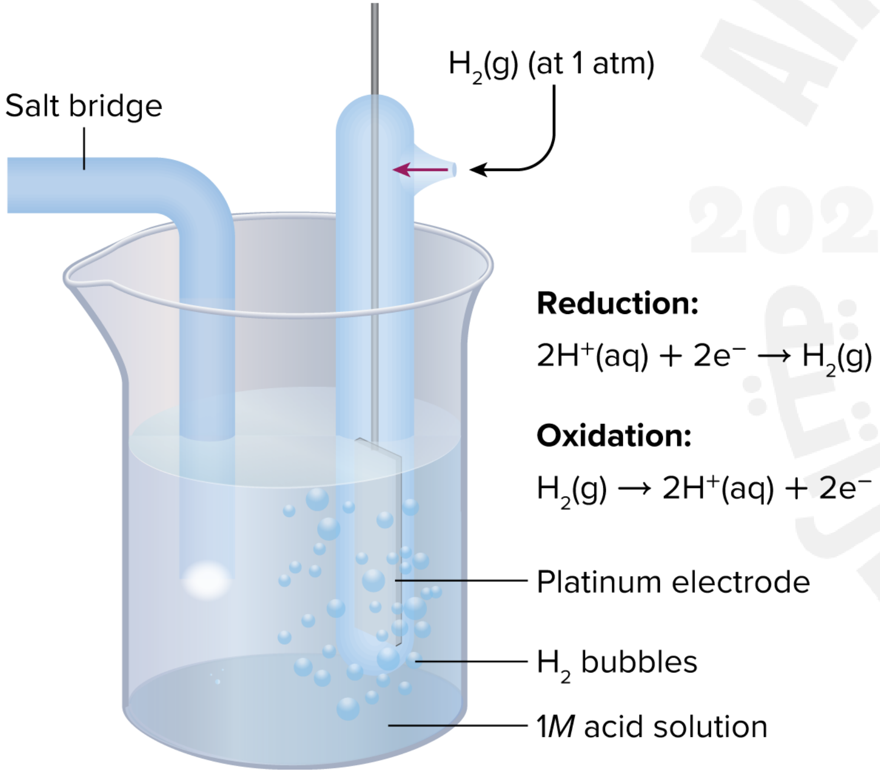
- The **standard hydrogen electrode (SHE)** consists of a small sheet of platinum immersed in a hydrochloric acid solution (HCl) that has a hydrogen ion concentration of 1 *M*. Hydrogen gas (H_2), at a pressure of 1 atm, is bubbled in and the temperature is maintained at 25°C.
- The standard hydrogen electrode is the standard against which all other reduction potentials are measured. **It is like point zero on the scale.**



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

Calculating Electrochemical Cell Potentials

- The standard reduction potential (potential of hydrogen electrode) E_o is **0.000 V**
- SHE can act as an oxidation-half reaction or a reduction-half reaction, depending on the half-cell to which it is connected

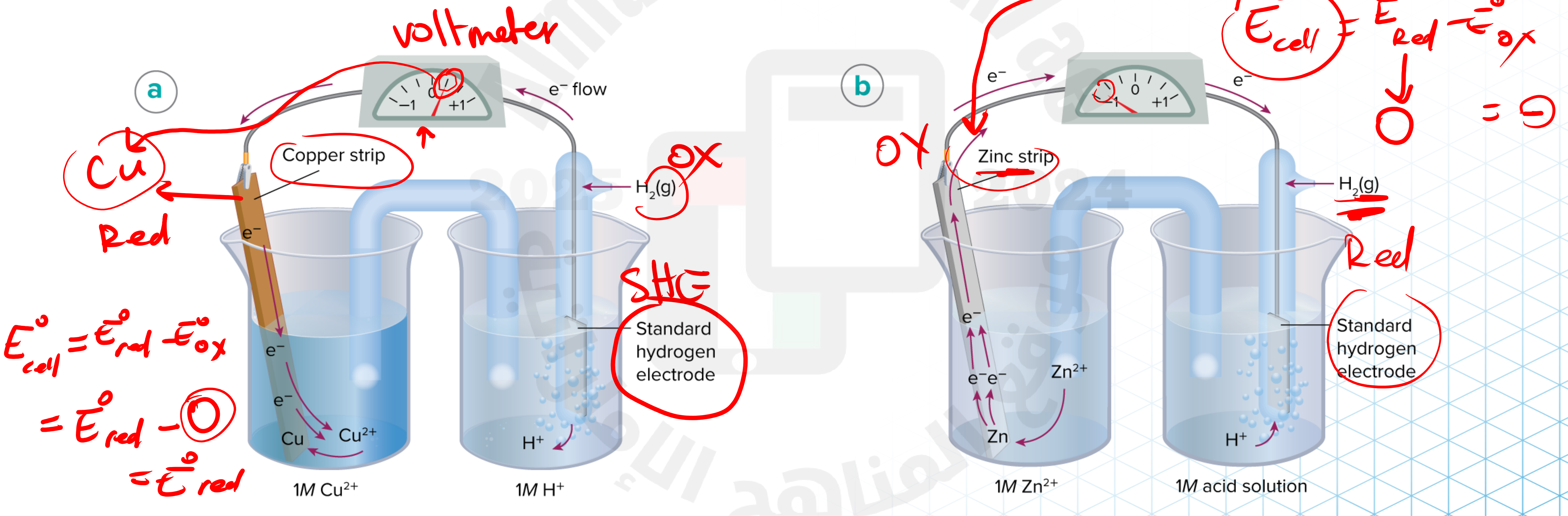


$$2H^+(aq) + 2e^- \underset{\text{Oxidation}}{\overset{\text{Reduction}}{\rightleftharpoons}} H_2(g) \qquad E^0 = 0.000\ V$$

$E^\circ_{cell} = E^\circ_{red} - E^\circ_{ox}$

What about the standard potential (E°) of the other elements? How can we calculate it?

By comparing it to “SHE”



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

$E^o_{cell} = E^o_{red} - E^o_{ox}$

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

Part 3

Learning objectives:

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

Use the half-cell standard reduction potentials to calculate the electrochemical cell standard potential **(E^0_{cell}),** while determining whether the redox reactions are spontaneous or nonspontaneous.

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.

Reactant Product Reactant Product

$$H_2 \mid H^+ \parallel Cu^{2+} \mid Cu \qquad E^0_{Cu} = +0.342\ V$$

→ *Cell Notation*

Oxidation half-cell — Reduction half-cell

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

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

Activity: In a cell consisting of a cadmium electrode and a gold electrode, the half-reactions take place:

$$Cd(s) \longrightarrow Cd^{2+}(aq) + 2e^-$$
$$Au^{3+}(aq) + 3e^- \longrightarrow Au(s)$$

Complete the equation for the overall reaction occurring in the cell.

$Cd(s)$ | $Cd^{2+}(aq)$ | $Au(s)$ | $Au^{3+}(aq)$ | 1 | 2 | 3 | 4

3 ____ + ____ ____ $\longrightarrow$ ____ ____ + 2 ____

Complete the notation for the cell.

Cd | Au | Au^{3+} | Cd^{2+} | $|$ | $\|$

____ ____ ____ ____ ____ ____ ____

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Page 5

Calculating Electrochemical Cell Potentials

- Formula for Cell Potential

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

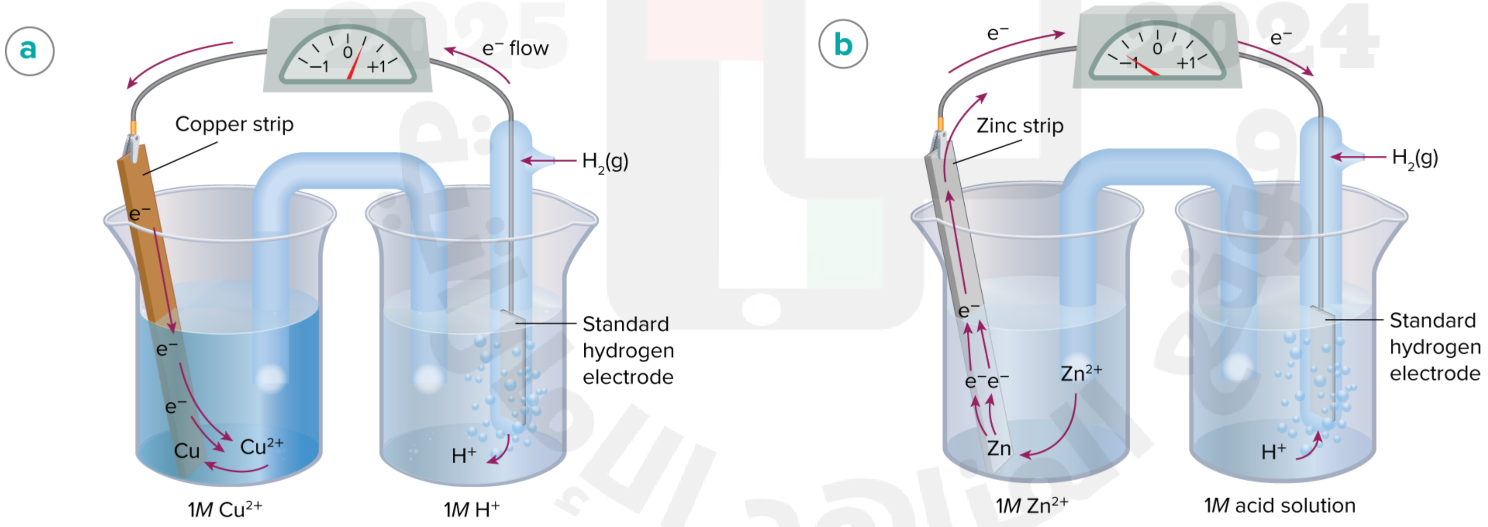
or

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

- 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)**



Rules to calculate E^0cell

- All half-reactions are given as **reduction half reactions** in a standard table (Table 1).

Table 1 Standard Reduction Potentials

Half-Reaction	E^0 (V)
$Li^+ + e^- \rightleftarrows Li$	−3.0401
$Ca^{2+} + 2e^- \rightleftarrows Ca$	−2.868

- When a half reaction is multiplied by an integer, *E*° **remains the same**.

$Li^+ + e^- \rightarrow$ **Li**	**E° = -3.0401 V**
2Li^+ + **2**e^- $\rightarrow$ **2Li**	**E° = -3.0401 V**
4Li^+ + **4**e^- $\rightarrow$ **4Li**	**E° = -3.0401 V**

- **Redox reactions are spontaneous if E°cell has a positive value** and **nonspontaneous if E° cell has a negative value**

Calculate a Cell Potential

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.

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

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

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

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

Calculate a Cell Potential

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.

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

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

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

Calculate a Cell Potential

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)**

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

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

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.

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

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

Calculate a Cell Potential

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)**

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

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

Is this reaction spontaneous or nonspontaneous?

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

APPLICATIONS

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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 chapter on redox reactions to review writing and balancing redox equations.

1. $Pt^{2+}(aq) + 2e^- \rightarrow Pt(s)$ and $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$
2. $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$ and $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$

Use Table 1 in page 183 to find the E^0 for the half cells

① $Pt^{2+} + 2e^- \rightarrow Pt$ $E^\circ = +1.18\ V$

$Sn^{2+} + 2e^- \rightarrow Sn$ $E^\circ = -0.1375\ V$

② $Co^{2+} + 2e^- \rightarrow Co$ $E^\circ = -0.28V$

$Cr^{3+} + 3e^- \rightarrow Cr$ $E^\circ = -0.744\ V$

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

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

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

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

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

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

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

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

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

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

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

More Practice

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Page 4

Use Standard Reduction Potentials

- Cell potentials can be used to determine if a proposed reaction under standard conditions will be spontaneous.
- If the **calculated potential is positive**, the reaction is **spontaneous**.
- If the calculated **potential is negative**, the reaction is not spontaneous.

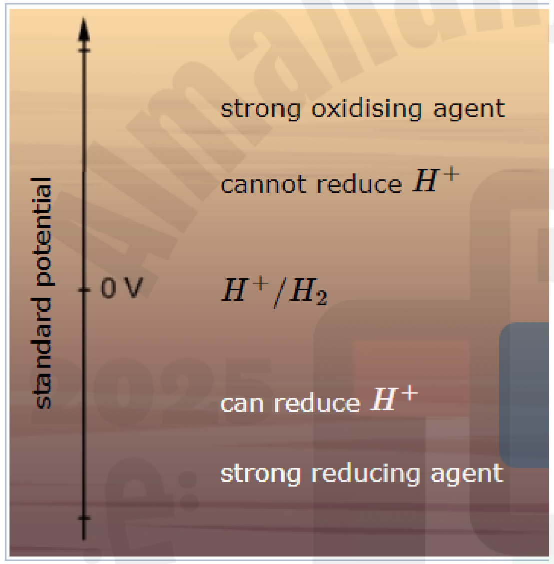
Use the standard reduction potentials in identifying the strongest reducing or oxidizing agent and the substance that is easily oxidized or reduced

Standard Electrochemical Potential 5/8

Applications of Standard Reduction Potentials

Reducing/Oxidising Ability of a Half-cell

The sign and magnitude of its standard potential define the ability of a species to lose or gain electrons.



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Page 5

Main Points:

- The **higher the standard potential (E^0) value**, the **stronger the oxidizing agent** (more easily reduced). Or the weaker the reducing agent

- The **lower the standard potential (E^0) value**, the **stronger the reducing agent** (more easily oxidized). Or the weaker the oxidizing agent)

APPLICATIONS

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?

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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

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