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VCE Chemistry ¾
Electroplating [2.4]
Workbook

Outline:

Learning Objectives:

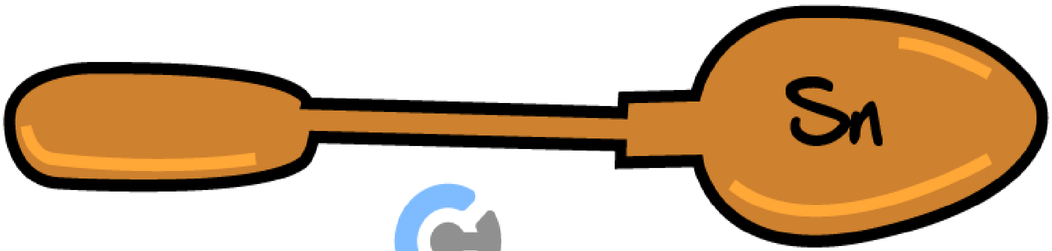
- ☐ **CH34 [2.4.1]** - Identify the Electroplating Setup (Location of the Object) & Find the Electroplating Reactions
- ☐ **CH34 [2.4.2]** - Find Next-Order Reactions During Electrolysis
- ☐ **CH34 [2.4.3]** - Apply Faraday's Laws to Electroplating Calculations

Section A: Electroplating

Sub-Section: Introduction to Electroplating

Context

- Electroplating is the process of coating a metal onto another conductive substance.



Electroplating	
'Electro- '	'Plating'
Using ____________________.	__________________ a metal onto another one.

Discussion: What are some uses of electroplating?

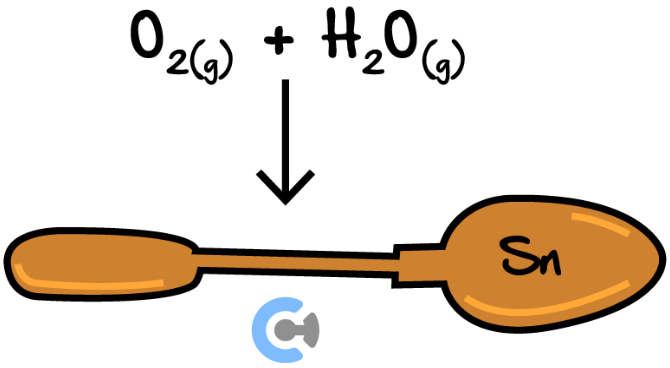
- Electroplating helps to prevent ___________________________________.
- Improve ______________________________________.
- Gain some ____________________ of the coated metal. For instance:
  - Copper metal (Cu) → ________________________.
  - Zinc metal (Zn) → ________________________.
  - Nickel metal (Ni) → ________________________.

NOTE: Electroplating is also _______________ and ___________ to set up!

How does electroplating a metal such as silver or onto a metal such as tin stop rusting?

<u>Exploration</u>: How does electroplating stop rusting?

- Substances that cause rust:
  - ________________________
  - ________________________
- Consider if a **tin (Sn) spoon** comes in contact with **water vapour** and **oxygen**.



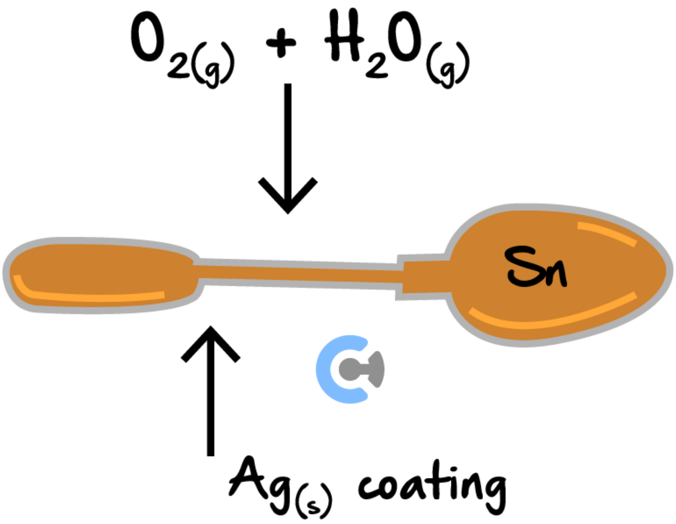
- Oxygen, water and tin on electrochemical series: *(Circle Below.)*

$\mathbf{O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)}$	$\mathbf{+1.23}$
$Br_2(l) + 2e^- \rightleftharpoons \mathbf{2Br^-(aq)}$	$+1.09$
$Ag^+(aq) + e^- \rightleftharpoons \mathbf{Ag(s)}$	$+0.80$
$Fe^{3+}(aq) + e^- \rightleftharpoons \mathbf{Fe^{2+}(aq)}$	$+0.77$
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons \mathbf{H_2O_2(aq)}$	$+0.68$
$I_2(s) + 2e^- \rightleftharpoons \mathbf{2I^-(aq)}$	$+0.54$
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons \mathbf{4OH^-(aq)}$	$+0.40$
$Cu^{2+}(aq) + 2e^- \rightleftharpoons \mathbf{Cu(s)}$	$+0.34$
$Sn^{4+}(aq) + 2e^- \rightleftharpoons \mathbf{Sn^{2+}(aq)}$	$+0.15$
$2H^+(aq) + 2e^+ \rightleftharpoons \mathbf{H_2(g)}$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons \mathbf{Pb(s)}$	-0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons \mathbf{Sn(s)}$	-0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons \mathbf{Ni(s)}$	-0.25

- **Spontaneous Reaction:**

<u>Reduction</u>	<u>Oxidation</u>

- Consider a spoon electroplated with silver:



- Can the water and oxygen gas come in contact with the tin on the inside? [Yes] / [No]
- Can the water and oxygen gas react with the silver metal ($Ag(s)$) spontaneously? [Yes] / [No]

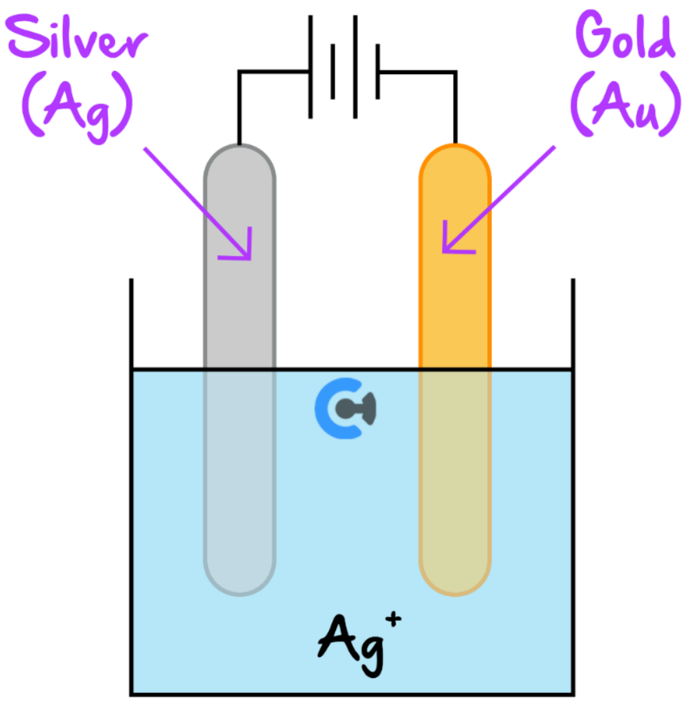
How do we actually set up electroplating?

Space for Personal Notes

Sub-Section: Electroplating Setup

Exploration: Electroplating Setup

- Electroplating is carried out in an **electrolytic cell** and thus, the **setup** is very **similar**.
- **Setup**: Two electrodes, a silver solid anode and a gold solid cathode with a silver ion $(Ag^+(aq))$ electrolyte.



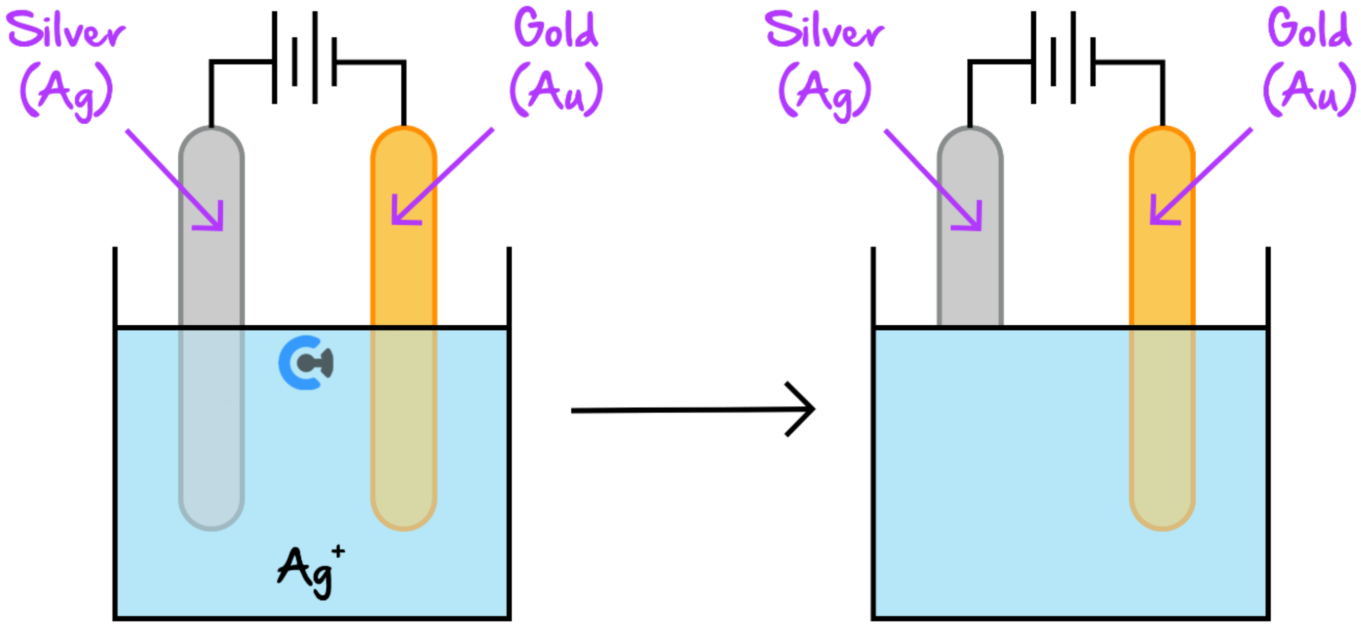
Reaction	Standard Electrode Potential ($E°$) in volts at 25°C
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	$+1.23$
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	$+1.09$
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	$+0.80$
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	-0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	-0.83
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	-1.18

Cathode Reduction Half-Equation	Anode Oxidation Half-Equation

- **Overall Reaction:** ____.

Let's have a look at what actually goes on in the cell!

Anode Oxidation Half-Equation	Cathode Reduction Half-Equation
$Ag(s) \rightarrow Ag^{+}(aq) + e^{-}$	$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$



- **Observations:**
  - **Size/Mass** of the **anode** over time: [Increases] / [Decreases]
  - **Size/Mass** of the **cathode** over time: [Increases] / [Decreases]
- **Conclusion**: While there is no 'overall reaction' occurring, silver metal (Ag) is ______________________ from one electrode to the other.
- Silver ions (Ag^{+}): [Cation] / [Anion]
- Electrode silver ions (Ag^{+}) moving to: [Cathode] / [Anode]

Definition

<u>Electroplating Process</u>

- **Definition**: **Plating a metal** onto **another metal** via electrolysis by **transferring** the metal from one electrode to the other.
- **Process**: Anode metal oxidises into its ion form moves to the cathode and reacts.
- **Ion Flow:** Cations → Cathode, Anions → Anode.

Let's look at an example together!

<u>Exploration</u>: **Placing Objects during Electroplating**

- **Electroplating Objects:**
  - **Object** that is to be electroplated. (e.g., key, spoon.)
  - **Metal is used** to coat the other object.

For the <u>object</u> to be electroplated:

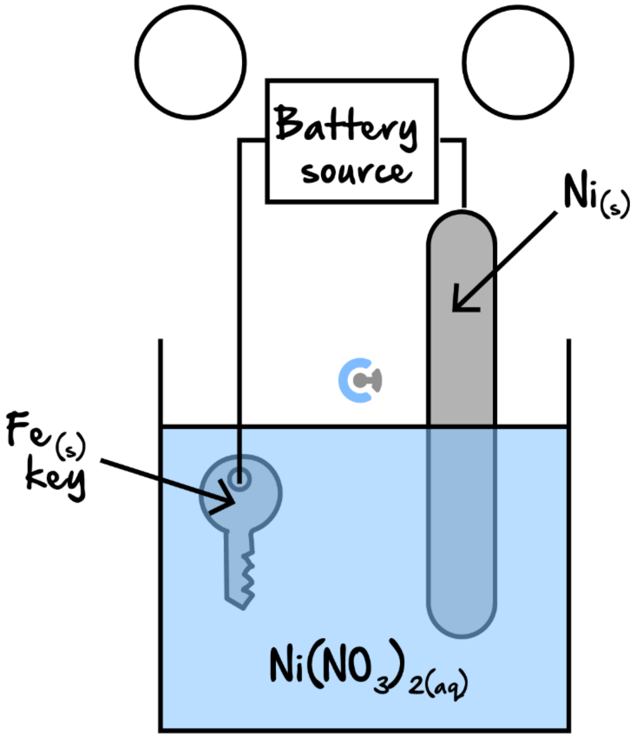
- Change in Mass Desired: [Increase] / [Decrease]
- Location Placed: [Cathode] / [Anode]
- Terminal Connected: [Positive] / [Negative]

For the <u>metal</u> used to electroplate:

- Change in Mass Desired: [Increase] / [Decrease]
- Location Placed: [Cathode] / [Anode]
- Terminal Connected: [Positive] / [Negative]

Example:

➤ **Scenario**: Electroplating an iron key with a nickel electrode and nickel nitrate solution $Ni(NO_3)_2(aq)$:



Reaction	Standard Electrode Potential ($E°$) in volts at 25℃
$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$	+1.77
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.25
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.44
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83

Cathode Reduction Half-Equation	Anode Oxidation Half-Equation

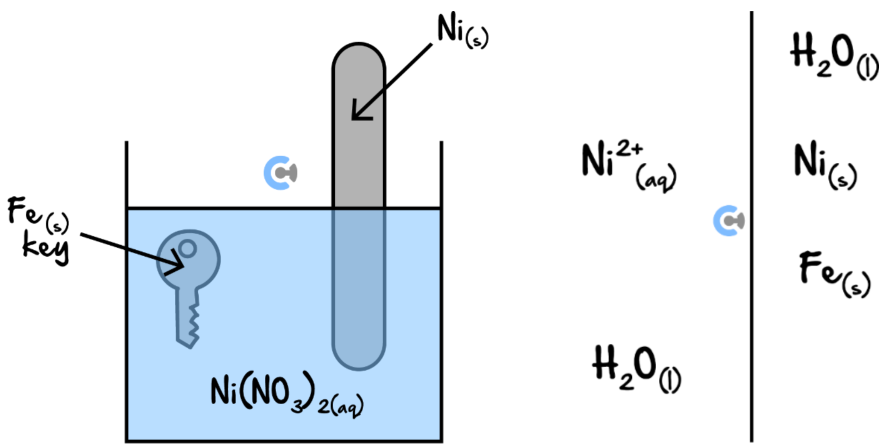
- Concentration of $Ni^{2+}(aq)$ overtime: ______________________________.

NOTE: While Ni^{2+} is being used at the cathode, the same amount of Ni^{2+} is being produced at the anode, resulting in **no overall change in the concentration** of nickel ions (Ni^{2+}).

What happens if no electrical energy is inputted into this cell?

Exploration: No Electrical Energy Inputted

- Visualisation without input of electrical energy:



- **Reaction:**

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

- **Spontaneity**: [Spontaneous] / [Non-spontaneous]
- **Process Name**: ____________________________ (which is not covered in VCAA).

<u>Extension:</u> **If galvanisation happens spontaneously, why don't we just allow galvanisation to take place in the previous scenario? Why bother with electroplating?**

- During galvanisation, the iron key oxidises:

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

- This causes the key to lose its shape and ridges, making the whole process not as smooth!

Electroplating Setup

Object to be Electroplated	Metal Used to Electroplate
[Cathode] / [Anode]	[Cathode] / [Anode]
Attached [positive] / [negative] terminal	Attached [positive] / [negative] terminal

- **Concentration**: No change.
- **Electrolyte Movement**: Cations → Cathode, Anions → Anode.

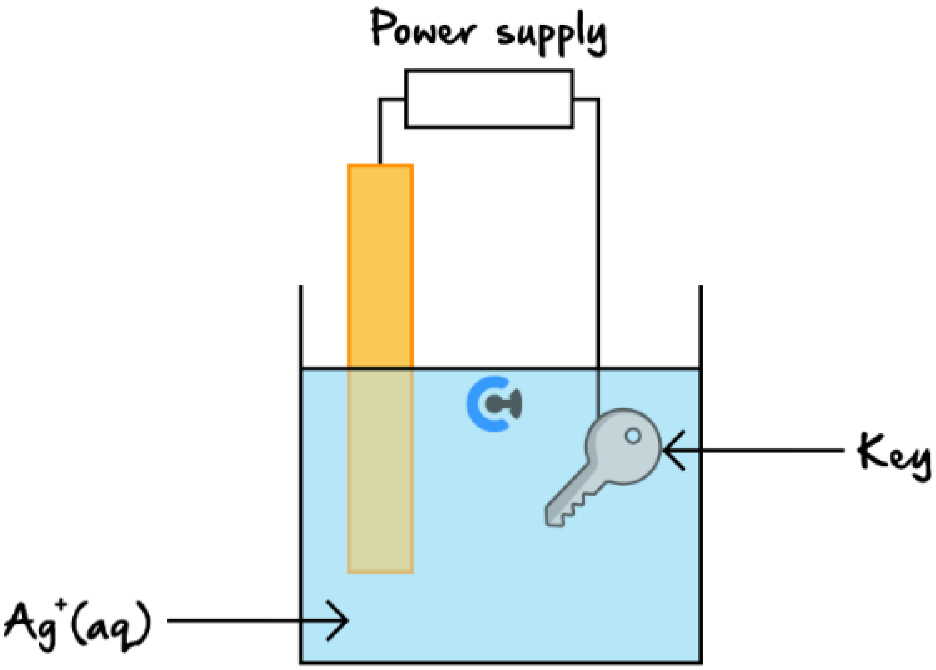
Space for Personal Notes

Your turn!

Question 1

Inspired from VCAA Chemistry Exam 2002
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/chem22002.pdf#page=4

A student decided to silver-plate a locker key using the apparatus shown:



In this cell, the key is the:

A. Anode, and is connected to the positive terminal of the power supply.

B. Anode, and is connected to the negative terminal of the power supply.

C. Cathode, and is connected to the positive terminal of the power supply.

D. Cathode, and is connected to the negative terminal of the power supply.

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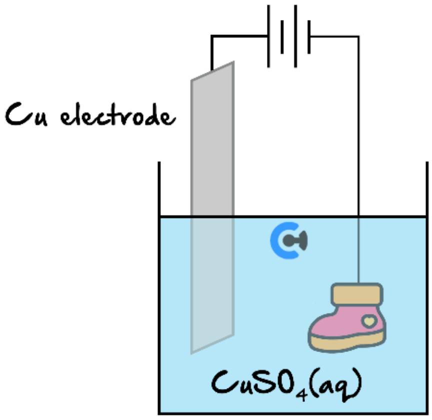
Question 2

Inspired from VCAA Chemistry Exam 2019
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2019/2019chem-w.pdf#page=7

There is a tradition of bronzing baby shoes. Before electroplating, the shoe is painted with a conductive material. The copper, Cu, electrode and copper sulphate, $CuSO_4$, solution cell used for electroplating a shoe is shown below.



During the electroplating process:

A. The copper electrode is oxidised and its mass is unchanged.

B. The shoe is coated with copper metal at the cathode.

C. The copper electrode is the oxidising agent.

D. Oxygen is produced at the cathode.

Question 3

Tin metal is electroplated onto sheets of iron using a tin anode and a well-stirred solution containing Sn^{2+} ions. During this process:

A. The anode increases in mass.

B. The concentration of Sn^{2+} in the solution decreases.

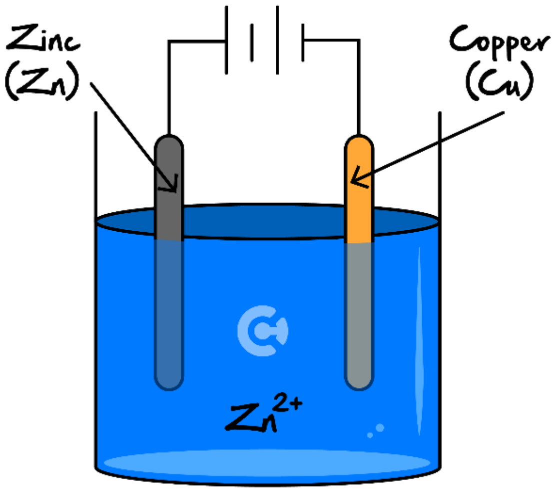
C. The concentration of Sn^{2+} in the solution increases.

D. Sn^{2+} ions move towards the cathode.

REMINDER: The concentration of the electrolyte does not change!

Sub-Section: Voltage Input

Exploration: Electroplating of Zinc Nitrate



Cathode: $Zn^{+}(aq) + e^{-} \rightarrow Zn(s)$

Anode: $Zn(s) \rightarrow Zn^{2+}(aq) + e^{-}$

EMF of the cell	Voltage to be inputted

- Usually, **very low voltage are inputted!**

Voltage Input During Electroplating

- If $EMF = 0\ V$, voltage inputted is very low.

Extension: What happens if a high voltage is inputted?

- **Reactivity around the electrodes at higher energies?** [Increases] / [Decreases]
- **Likelihood of Substances to React:** [More] / [Less] likely
- **If higher voltages (over $4.0\ V$) are used:** May cause other species other than the strongest oxidant and reductant **to react!**
- **Covered in-depth in:** Electroplating practical booklet

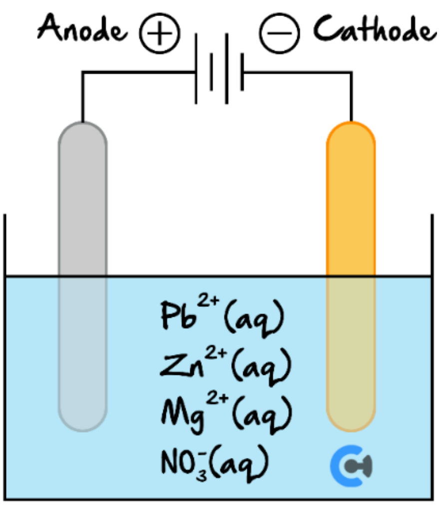
Section B: Next Order Reactions

Context

- Sometimes, as the cell operates, the reactants start to **run out**.
- **Result**: Reactions which occur at each electrode start to ____________.
- **Known as:** ______________________ reaction.

Exploration: Next Order Reaction

- Consider the following electrolytic cell:



- **Strongest Oxidant & Strongest Reductant:** *(Label Below.)*

First Order Reaction:

Reduction Cathode	Oxidation Anode

- **Scenario**: Low concentrations of the metal ions are provided.
- **Amount of $Pb^{2+}(aq)$ and $H_2O(l)$ over time:** ____________________
- **Conclusion**: Reaction at the cathode will **change** over time.

Second Order Reaction:

- After $Pb^{2+}(aq)$ **has run out completely**, new strongest oxidant: __________

Reduction Cathode	Oxidation Anode
	$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

Third Order Reaction:

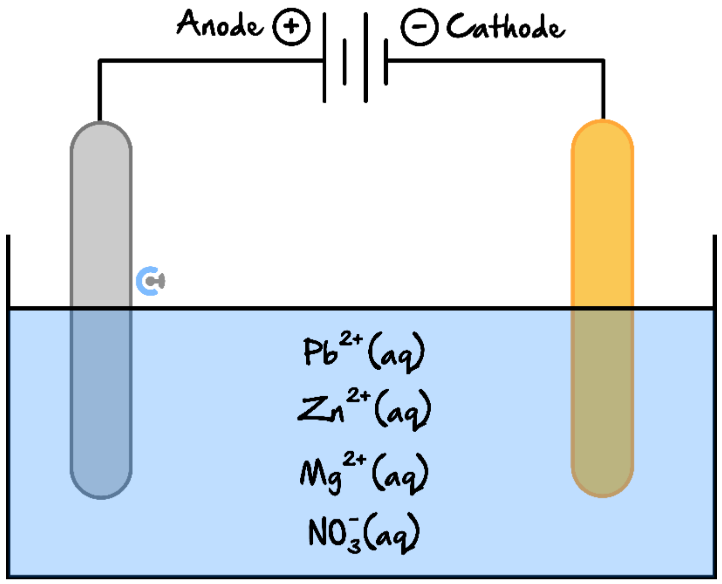
- **Scenario**: $Zn^{2+}(aq)$ also runs out after some time.
- **New Strongest Oxidant:** ____________

Reduction Cathode	Oxidation Anode
	$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

Overall Steps of Reactions:

- **Cathode Reactions:**
  - First order: $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$
  - Second order: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
  - Third order: $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

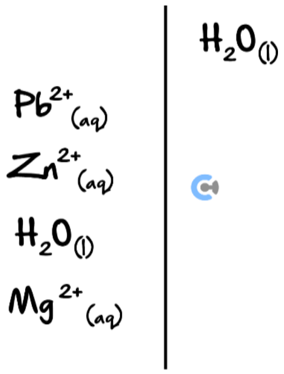
- Products Formed Over Time: *(Label Below.)*



- $Pb(s)$ reduced **first:** Forms __________ coating of electrode.
- $Zn(s)$ reduced **second**: Forms __________ coating of electrode.
- Water reduced third: ______________________________.

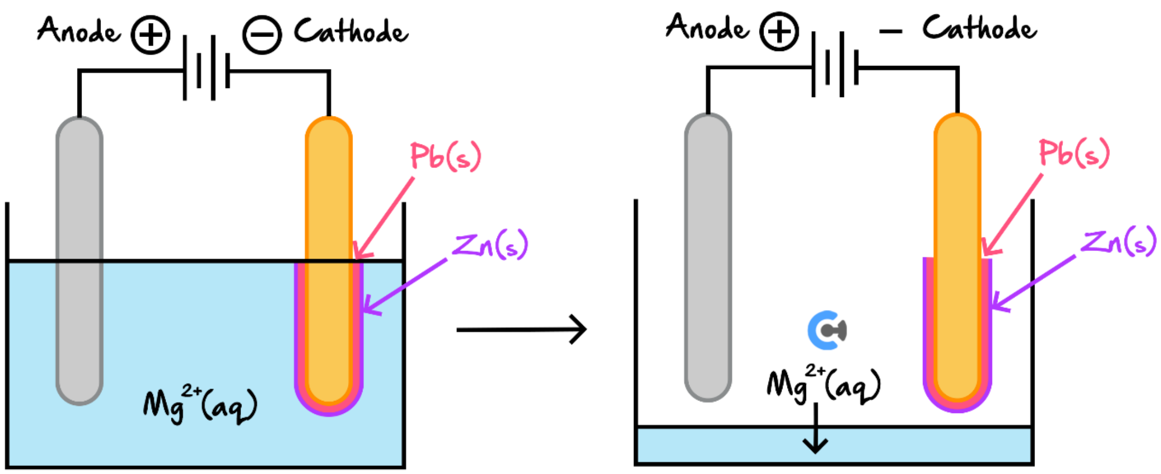
Next:

- **Scenario**: After third-order reaction, water reacts with water.
- What happens when we **run out of water**? Is there a fourth-order reaction with the Mg^{2+}?



[Yes, there is a fourth-order reaction] / [No there isn't]

- **Cell After Water Runs Out:**



- **Observation**: Water no longer _______________ the electrode.

- Can the magnesium cations (Mg^{2+}) react? [Yes] / [No]

Definition

<u>Next Order Reactions</u>

- **Predicting Next Order Reactions:** Assume current oxidant/reductant runs out, move to the next strongest oxidant/reductant.

- **Endgame Scenario**: Electrolysis of water, as after water is used up, electrodes no longer touches electrolyte and the electrolytic cell ____________________________ and _______________ longer operates.

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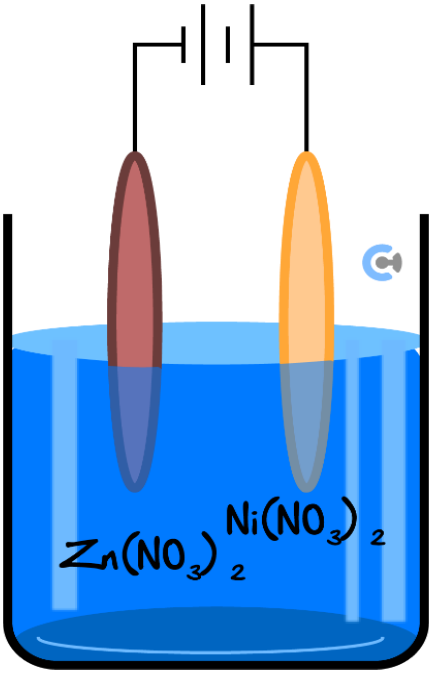
Let's have a look at a question together!

Question 4 (4 marks) **Walkthrough.**

Consider an aqueous solution that contains both zinc (II) nitrate and nickel (II) nitrate. The solution is electrolysed using inert electrodes, as shown below.



a. Write the balanced half-equations for the reaction which takes place at the:

i. Cathode. (1 mark)

____

ii. Anode. (1 mark)

____

b. After some time has elapsed, a new half-equation occurs at one of the electrodes. Write the half-equation for the new reaction which takes place. (1 mark)

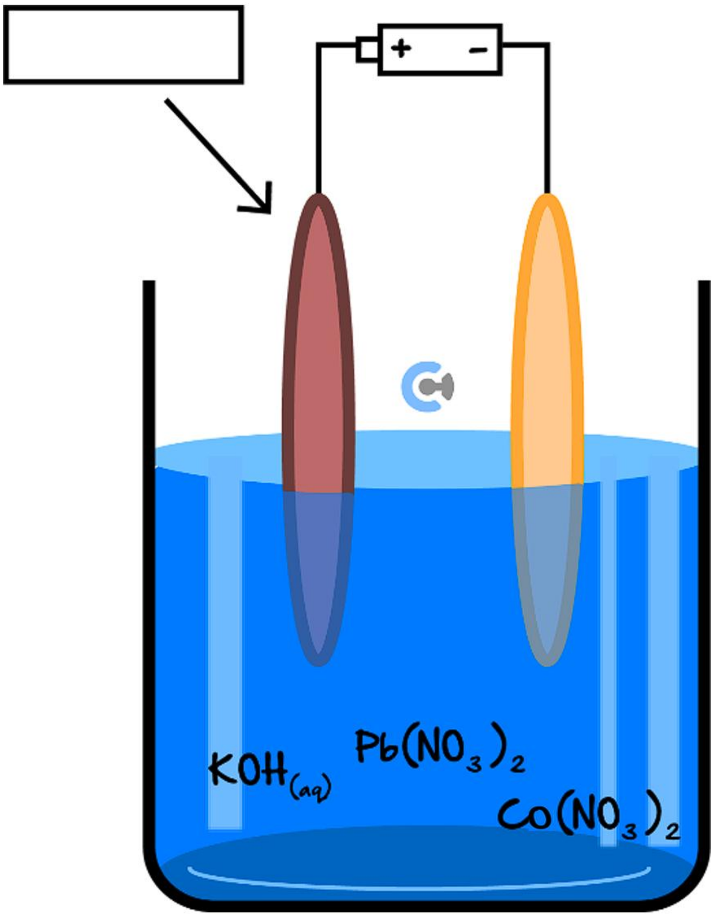
____

c. As the cell operates, a coating is seen to form over one of the electrodes. Draw the coating(s) that form at the electrode on the diagram above. (1 mark)

Your turn!

Question 5 (7 marks)

A solution of potassium hydroxide, cobalt (II) nitrate and lead (II) nitrate is electrolysed using inert electrodes. A setup of the cell is shown below.



a. Label the left electrode as the cathode or anode in the box provided above. (1 mark)

b. Initially, a set of reactions takes place at each electrode. Write the balanced half-equation which occurs at the:

i. Cathode. (1 mark)

____

ii. Anode. (1 mark)

____

c. After some time has elapsed, the reactions which take place at the negative electrode appear to **change**.

i. Write the balanced half-equation for the next reaction which takes place at the negative electrode. (1 mark)

____

ii. Write the balanced half-equation for the next reaction after the reaction indicated in **part c. i.** which takes place at the negative electrode. (1 mark)

____

iii. On the diagram above, draw the products which form at the cathode. (1 mark)

d. After a long time, the cell stops functioning. State the cations that will still be present in the electrolyte after this happens. (1 mark)

____

Question 6 Additional Question.

An aqueous solution was prepared by dissolving $0.100\ mol$ of each of $Co(NO_3)_2$, $Cu(NO_3)_2$, $Zn(NO_3)_2$, and $Ni(NO_3)_2$ in $1.00\ L$ of deionised water. The electrolysis of this aqueous solution was then carried out using platinum electrodes. At the end of the electrolysis, all the metal cations had been deposited onto the cathode in four successive coatings.

The order of the coatings on the cathode from **outside** to **inside** is:

A. Zn, Cu, Co, Ni

B. Zn, Co, Ni, Cu

C. Cu, Zn, Co, Ni

D. Cu, Ni, Co, Zn

Space for Personal Notes

Section C: Faraday's Laws of Electrolysis

While we've covered Faraday's Laws already, it is also present in AOS 2, and it is slightly different as well! Let's recap!

Electric Charge

- **Definition:** Electric charge is a physical property of matter that causes matter to experience a force when placed in an electromagnetic field.
- **Simplified:**

"How many positive or negative substances are there?"

Denoted By	SI Units

REMINDERS: 'q' - energy, 'Q' - electric charge.

Electric Current

- **Definition:** An electric current is a **flow** of charged particles, such as electrons or ions, moving through an electrical conductor or space.

Denoted By	SI Units

Active Recall: What is the formula that links electric charge and electric current together?

____

Question 7

a. The current passing through a cell is at 2.50 A and the cell runs for 15 minutes. Calculate the amount of electrical charge that has passed through the cell.

---

b. The cell current is changed, whereby 20 C of charge is known to have passed through over a time period of 20 minutes. Find the current flowing through the cell in mA.

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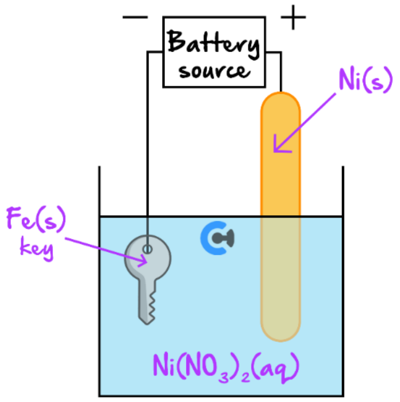
Sub-Section: Faraday's First Law of Electrochemistry

Exploration: Linking Electric Charge to Reactants/Products

- Consider the following electrolytic cell:



Cathode: $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$

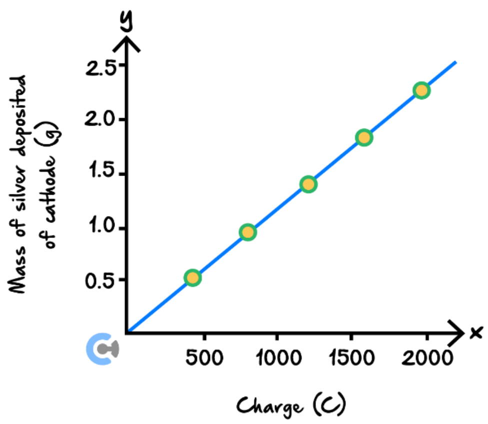
Anode: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$

- If **more electric charge** passes through, the **amount of electrons** arriving: [Increases] / [Decreases]
- Amount of nickel metal produced at the cathode: [Increases] / [Decreases]

Faraday's First Law of Electrochemistry

"The amount of chemical change being produced by a current at an electrode-electrolyte interface is proportional to the quantity of electricity used."

$$Q \propto m$$



- The **more electrical charge** passes through, the **amount of substances** reacting/produced **increases**.

Sub-Section: Faraday's Second Law of Electrochemistry

What does electric charge itself mean?

Exploration: Charge on a Mole of Electrons

- What is the charge of a single electron (in C)?

$$e^-$$

  - Electric Charge: ____________________

- Number electrons in one mole of electrons: ____________________

- Charge of **one mole** of electrons:

Faraday's Constant

- **Definition:** The quantity of **charge** carried by **one mole** of **electrons**.

- **Data Book:** Page 5.

Denoted By	Value	Formulae

Space for Personal Notes

Let's go through a question together!

Question 8 Walkthrough.

Given that 2.70 mol of electrons are passed through a cell, calculate the electrical charge passing through.

Your turn!

Question 9

Given that 0.150 mol of electrons pass through a cell, calculate the amount of electric charge which passes through.

Question 10

It is known that 1250 C of electric charge is produced in a galvanic cell. Find the amount of electrons which must have passed through.

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Definition

Faraday's Second Law of Electrochemistry

- **Law**: *"If the same amount of electricity is passed through different electrolytes, the masses of ions deposited at the electrodes are directly proportional to their chemical equivalents."*
- **Law Simplified**: Consider the **stoichiometric ratio** between the **electrons** and **metal**.
- To make $\mathbf{1}\ \boldsymbol{mol}$ of $\mathbf{M}$ from $\mathbf{M}^{\mathbf{z}^{+}}$, you need $\mathbf{z}\ \boldsymbol{mol}$ of electrons.

Misconception

"In the following equation, 1 mole of electrons produces 2 mol of zinc."

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

TRUTH:

- $1\ mol$ of electrons produces _______________ of zinc instead!

Let's look at a question together!

Question 11 Walkthrough.

Given that $6200\ C$ of charge was passed through a solution of zinc (II) chloride to cause it to electrolyse, find the change in mass of the negative electrode.

____

____

____

____

____

____

____

TIPS: To find the molar masses quickly, use **page 14 of the data book** which lists the elements in alphabetical order!

Your turn!

Question 12

7250 C of electric charge was passed through an electrolytic cell. Here is the reaction which occurs at the cathode.

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

a. Find the amount (in mol) of electrons which passed through.

---

b. Find the mass of iron which is deposited at the cathode.

Space for Personal Notes

Question 13

A solution of nickel (II) nitrate is electrolysed using inert electrodes.

a. Write the half-equation which occurs at the negative electrode.

____

b. The negative electrode is seen to increase in mass by 820 mg.

i. Find the amount of electrons which must have passed through the cell.

____

____

____

____

ii. Find the electric charge supplied by the power source.

____

____

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Sub-Section: Combining Equations Together

Active Recall: What are the two formulas to find electric charge?

Formula 1	Formula 2

Steps to go from I & t to $m(metal)$

- **Steps:**

1. Find the **electric charge** by using $Q = It$.
2. Find the **amount of electrons** by using $n(e^-)F$.
3. Find the **amount (in mol) of the metal** by using the equation:

$$\text{E.g. } Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$$

$$n(Zn) = \frac{1}{2}n(e^-)$$

4. Find the **mass of the metal** by using molar mass $m = n \times Mr$.

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Let's have a look at a question which combines the two equations together!

Question 14 Walkthrough.

Jerry wants to plate cobalt metal onto his nickel ring. He attaches his ring at the negative terminal and attaches a sheet of cobalt metal at the positive terminal. After running the electrolytic cell for 8.0 minutes at 3.0 A, what is the expected mass increase of the ring?

Space for Personal Notes

Your turn!

Question 15

In an electrolytic cell using a **concentrated NaCl** electrolyte and inert electrodes, chlorine gas is observed to be produced.

a. Write the half-equations which occur at the:

i. Positive electrode.

____

ii. Negative electrode.

____

b. Find the theoretical mass of chlorine gas (in kg) which is produced in 1.00 hours, given that the current is 1.40×10^3 A.

____

____

____

____

____

c. Given that the energy efficiency for the process is 90%, find the actual mass of chlorine gas observed to be produced.

____

____

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REMINDERS

- $\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$
- This can be found on **page 3 of the databook**.

Question 16

In an electroplating cell, $0.0500\ F$ is passed through the cell, whereby an iron key is attached to the negative terminal, and a lead electrode is attached to the other terminal. The electrolyte contains lead (II) nitrate.

a. Write the half-equation which occurs at the iron key electrode.

____

b. Find the amount of electrons which pass through the cell.

____

____

c. Find the mass change of the iron key.

____

____

NOTE: $1.00\ F$ is the charge carried by $1.00\ mol$ of electrons $= 96500\ C$.

ALSO NOTE: $1.00\ F = 1.00\ mol$ of electrons.

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Question 17 Additional Question.

Calculate the time taken to deposit 1.00 g of copper onto an object that is placed in a solution of copper nitrate $Cu(NO_3)_2$ and has a current of 2.50 A flowing through it.

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Sub-Section: Deriving the Charge on Metal Ion

Context

- **Transition metals** and **other metals** typically have charges which are different to their group.

Periodic table of the elements

Key: Atomic number: 79; Symbol of element: Au; Relative atomic mass: 197.0; Name of element: Gold

1 H 1.0 Hydrogen																	2 He 4.0 Helium
3 Li 6.9 Lithium	4 Be 9.0 Beryllium											5 B 10.8 Boron	6 C 12.0 Carbon	7 N 14.0 Nitrogen	8 O 16.0 Oxygen	9 F 19.0 Fluorine	10 Ne 20.2 Neon
11 Na 23.0 Sodium	12 Mg 24.3 Magnesium											13 Al 27.0 Aluminium	14 Si 28.1 Silicon	15 P 31.0 Phosphorus	16 S 32.1 Sulfur	17 Cl 35.5 Chlorine	18 Ar 39.9 Argon
19 K 39.1 Potassium	20 Ca 40.1 Calcium	21 Sc 45.0 Scandium	22 Ti 47.9 Titanium	23 V 50.9 Vanadium	24 Cr 52.0 Chromium	25 Mn 54.9 Manganese	26 Fe 55.8 Iron	27 Co 58.9 Cobalt	28 Ni 58.7 Nickel	29 Cu 63.5 Copper	30 Zn 65.4 Zinc	31 Ga 69.7 Gallium	32 Ge 72.6 Germanium	33 As 74.9 Arsenic	34 Se 79.0 Selenium	35 Br 79.9 Bromine	36 Kr 83.8 Krypton
37 Rb 85.5 Rubidium	38 Sr 87.6 Strontium	39 Y 88.9 Yttrium	40 Zr 91.2 Zirconium	41 Nb 92.9 Niobium	42 Mo 96.0 Molybdenum	43 Tc (98) Technetium	44 Ru 101.1 Ruthenium	45 Rh 102.9 Rhodium	46 Pd 106.4 Palladium	47 Ag 107.9 Silver	48 Cd 112.4 Cadmium	49 In 114.8 Indium	50 Sn 118.7 Tin	51 Sb 121.8 Antimony	52 Te 127.6 Tellurium	53 I 126.9 Iodine	54 Xe 131.3 Xenon
55 Cs 132.9 Caesium	56 Ba 137.3 Barium	57–71 Lanthanoids	72 Hf 178.5 Hafnium	73 Ta 180.9 Tantalum	74 W 183.8 Tungsten	75 Re 186.2 Rhenium	76 Os 190.2 Osmium	77 Ir 192.2 Iridium	78 Pt 195.1 Platinum	79 Au 197.1 Gold	80 Hg 200.6 Mercury	81 Tl 204.4 Thallium	82 Pb 207.2 Lead	83 Bi 209.0 Bismuth	84 Po (210) Polonium	85 At (210) Astatine	86 Rn (222) Radon
87 Fr (223) Francium	88 Ra (226) Radium	89–103 Actinoids	104 Rf (261) Rutherfordium	105 Db (262) Dubnium	106 Sg (266) Seaborgium	107 Bh (264) Bohrium	108 (267) Hassium	109 Mt (268) Meitnerium	110 Ds (271) Darmstadtium	111 Rg (272) Roentgenium	112 Cn (285) Copernicium	113 Nh (280) Nihonium	114 Fl (289) Flerovium	115 Mc (289) Moscovium	116 Lv (292) Livermorium	117 Ts (294) Tennessine	118 Og (294) Oganesson

57 La 138.9 Lanthanum	58 Ce 140.1 Cerium	59 Pr 140.9 Praseodymium	60 Nd 144.2 Neodymium	61 Pm (145) Promethium	62 Sm 150.4 Samarium	63 Eu 152.0 Europium	64 Gd 157.3 Gadolinium	65 Tb 158.9 Terbium	66 Dy 162.5 Dysprosium	67 Ho 164.9 Holmium	68 Er 167.3 Erbium	69 Tm 168.9 Thulium	70 Yb 173.1 Ytterbium	71 Lu 175.0 Lutetium
89 Ac (227) Actinium	90 Th 232.0 Thorium	91 Pa 231.0 Protactinium	92 U 238.0 Uranium	93 Np (237) Neptunium	94 Pu (244) Plutonium	95 Am (243) Americium	96 Cm (247) Curium	97 Bk (247) Berkelium	98 Cf (251) Californium	99 Es (252) Einsteinium	100 Fm (257) Fermium	101 Md (258) Mendelevium	102 No (259) Nobelium	103 Lr (262) Lawrencium

The value in brackets indicates the mass number of the longest-lived isotope

Exploration: Unknown charge of metal ion I

- Consider the following:

$$1\ mol\ of\ X, 5\ mol\ of\ e^-.$$

- Stoichiometric ratio between the X and e^-:

$$n(X) : n(e^-)$$

Balanced Reduction Half-Equation	Charge of X^{y+}

<u>Exploration</u>: Unknown charge of metal ion II

- Consider the following:

$$0.124\ mol\ of\ P, 0.374\ mol\ of\ e^{-}.$$

- Stoichiometric ratio between the P and e^{-}:

$$n(P) : n(e^{-})$$

- Charge of the P^{x+}: ____________________

NOTE: To find an integer ratio, divide both sides of the equation by the ________________ number present!

<u>Steps to Find Charge of Ion</u>

1. Find n(metal) by $n(\text{metal}) = \frac{m}{Mr}$.
2. Find $n(e^{-})$ by $n(e^{-}) = \frac{Q}{F}$.
3. Find the stoichiometric ratio by dividing by the smallest number present.

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Let's have a look at a question together!

Question 18 Walkthrough.

A solution of Vanadium ions are reduced at a cathode, whereby it is found that a current of 6.00 A is produced over 40 minutes. It is found that 1.91 g of Vanadium metal is deposited at the cathode. Find the charge of the Vanadium ions.

Question 19

The electrolysis of a solution of chromium ions using a current of 2.2 A for 25 minutes produced 0.60 g of chromium. Calculate the charge on the chromium ion.

Question 20 Additional Question.

It is found that 3500 C of electric charge is passed through a galvanic cell. As this takes place, it is found that 0.398 g of manganese metal is deposited. Find the charge on the Manganese metal ions.

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Section D: Mass of Metal Deposited at the Cathode

Context

- Sometimes, questions ask to **compare the mass of metal deposited** at the electrodes when the **same amount of charge** is passed through.

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Sub-Section: Same Electric Charge

Let's first compare two metals!

Exploration: Difference in mass

- Two electrolytic cells are setup, one which results in $Cu^{2+}(aq)$ reducing, and the other with $Ag^{+}(aq)$ reducing.
- **Copper reduction**: $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$
- **Silver reduction**: $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$
- **Scenario**: When the **same amount of electric charge** is passed through both cells, copper has a **different change in mass** to silver.
- **Why?**
  - ____
  - ____

How do we figure out which cell will deposit more mass?

Exploration: Comparing Mass Deposited

- **Scenario**: Two cells are setup with inert electrodes and both initialised for $\mathbf{3.00\ A}$ **of current** for $\mathbf{5.00}$ **minutes.**
  - **Cell A**: copper (II) nitrate.
  - **Cell B**: silver (I) nitrate.
- **Question**: Which cell has the greatest change in mass?

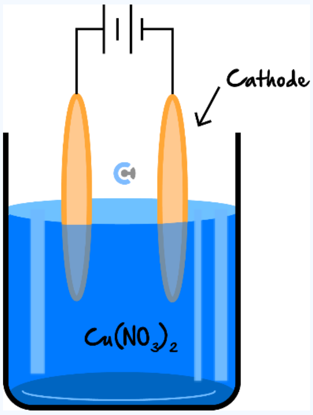
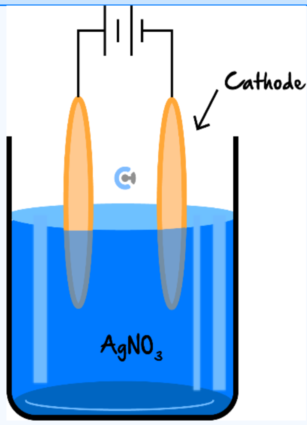
Method #1: Brute Force

- Amount of electrons (in mol) which pass through each cell:

____

____

____

Metal	Copper (II) ions (Cu^{2+})	Silver ions (Ag^{+})
Electrolytic Cell		
Cathode half-equation		
Mole of metal deposited		
Mass of metal deposited		

- Conclusion:

$$m(\mathrm{Cu}) \qquad m(\mathrm{Ag})$$

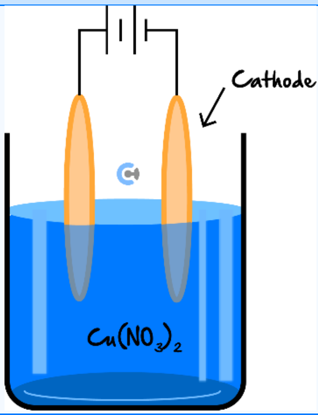
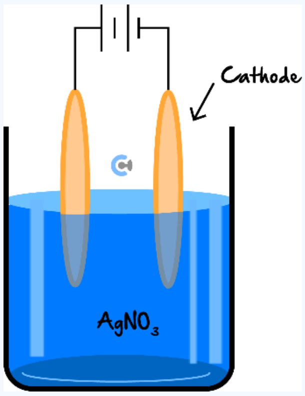
Is there a quicker method without finding the mole of electrons?

- **Consider:** If $3.00\ A$ of current for 5.00 minutes **changed to** $1.00\ A$ of current for 1.00 minute,
  - Will the answer have changed? [yes] / [no]

- **Consider**: If the **mole of electrons** $n(e^-)$ changed from $0.00933\ mol$ to $1.00\ mol$,
  - Will the answer have changed? [yes] / [no]
- **Conclusion**: can arbitrarily set the electric charge to ________________ and the mole of electrons ($n(e^-)$) to be ________________!

Method #2: Shortcut (Mr/Z Ratio)

- Assume $1.00\ mol$ **of electrons** passes through the cell.

Metal	Copper (II) ions (Cu^{2+})	Silver ions (Ag^+)
Electrolytic Cell		
Cathode half-equation	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	$Ag^+(aq) + e^- \rightarrow Ag(s)$
Mole of metal deposited		
Mass of metal deposited		

What have we actually done for each of them?

Metal	Copper (II) ions (Cu^{2+})	Silver ions (Ag^{+})
Operations		
Operations Summarised		

- **Result**: find the molar mass:charge ratio (Mr/z) for each, see which has a greater ratio!
- Mr/z ratio:

Copper (II) ions (Cu^{2+})	Silver ions (Ag^{+})

- Conclusion:

$$m(\text{Cu}) \qquad m(\text{Ag})$$

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Definition

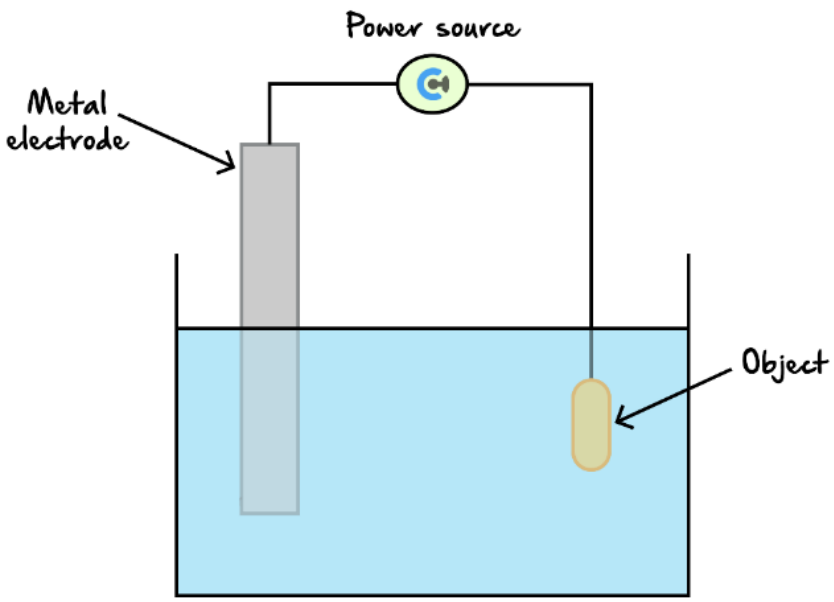
Comparison of Mass Deposited

- **Scenario**: The question asks to **compare** which substance produces the largest mass.
- **Solution**: Find molar mass: charge ratio (**Mr/z ratio**). The substance with the highest Mr/z ratio will deposit the largest mass!
- **Caveat**: If the question asks for the **numerical amount** of mass deposited, **cannot use this method**!

Let's have a look at a question together!

Question 21 Walkthrough.

The diagram below shows an electroplating cell:



The cell contains 1 L of an electroplating solution. The electroplating cell is run for one hour at 3 A.

Which one of the following electroplating solutions will deposit the largest mass of metal onto the object?

A. 1 M $AgNO_3$

B. 1 M $Cd(NO_3)_2$

C. 1 M $Pb(NO_3)_2$

D. 1 M $Al(NO_3)_3$

Your turn!

Question 22

Two electrolytic cells are compared, whereby 5.20 A of current is passed through for 45 min for each cell.

Cell A contains cobalt (II) nitrate, while cell B contains nickel (II) nitrate. Determine which cell will have a greater change in mass at the negative electrode.

Question 23

Four different electroplating cells are setup, with a material at the anode and an appropriate electroplating solution.

Which option will result in the **smallest change in mass** when the same current and time is passed through as that of the Pb?

	Material at Anode	Electroplating Solution
A.	Chromium, Cr	1.0 M $Cr(NO_3)_3$
B.	Silver, Ag	1.0 M $AgNO_3$
C.	Gold, Au	1.0 M $AuCl_3$
D.	Tin, Sn	1.0 M $SnSO_4$

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Question 24

Inspired from VCAA Chemistry Exam 2013
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2013/2013chem-w.pdf#page=12

A student prepares 1.0 *M* aqueous solutions of $AgNO_3, Fe(NO_3)_2$ and KNO_3.

Equal volumes of each solution are placed in separate beakers, identical platinum electrodes are placed in each beaker and each solution undergoes electrolysis with the same current applied for 5.0 minutes under SLC. Each cathode is then dried and weighed to determine mass change.

Assume that the concentrations of the solutions have decreased only slightly.

In order of increasing mass, the metals deposited on the three cathodes are likely to be:

A. Potassium, Silver, Iron.

B. Silver, Iron, Potassium.

C. Iron, Potassium, Silver.

D. Potassium, Iron, Silver.

NOTE: Be sure to double check if all metal ions reduce in the first place!

TIP: For High Achievers – try to memorise all species below water!

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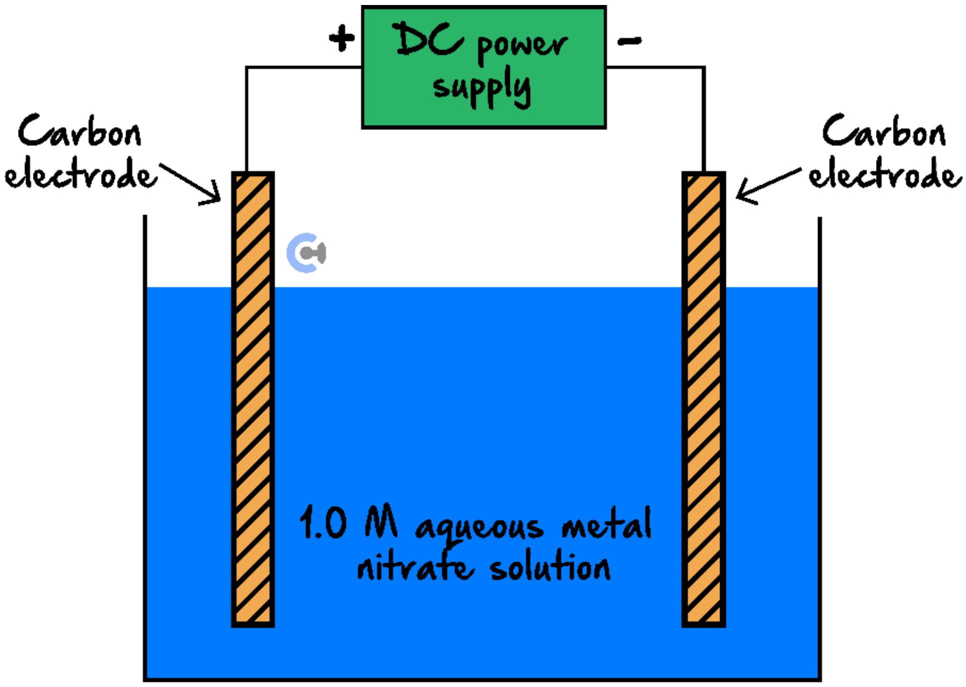
Question 25 Additional Question.

Inspired from VCAA Chemistry Exam 2 2011
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2011chem2-w.pdf#page=9

A series of electrolysis experiments are conducted using the apparatus shown below.



An electric charge of 0.030 Faraday was passed through separate solutions of 1.0 M $Cr(NO_3)_3$.

1.0 M $Cu(NO_3)_2$ and 1.0 M $AgNO_3$. In each case, the corresponding metal was deposited on the negative electrode.

The amount, in mol, of each metal deposited is:

	Amount, in mol, of chromium deposited	**Amount, in mol, of copper deposited**	**Amount, in mol, of silver deposited**
A.	0.030	0.030	0.030
B.	0.010	0.015	0.030
C.	0.090	0.060	0.030
D.	0.030	0.020	0.010

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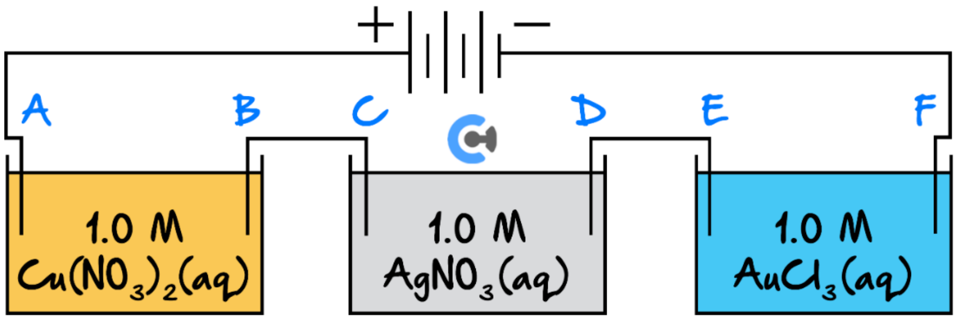
Sub-Section: Connected Cells

Discussion: When electrolytic cells are connected such as the following, and $1.00\ A$ of current is passed through for $60\ s$, what is the amount of electric charge which passes through each cell?



____

Question 26

Based on the connected cell diagram provided above,

a. Which electrodes *A*-*F* are acting as cathodes?

____

b. If the same 1.00 *A* of current is passed through for 60 *s*, find which cathode out of the ones identified in **part a.** will have the smallest increase in mass.

____

____

____

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Contour Check

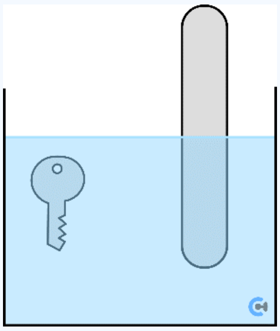
Learning Objective: [2.4.1] Identify the electroplating setup (location of object) & find the electroplating reactions

Key Takeaways

- Definition:

Object	Metal Used
[Cathode] / [Anode]	[Cathode] / [Anode]
[Positive] / Negative]	[Positive] / Negative]

- Setup:



Anode Reaction	Cathode Reaction

- Concentration of Electrolyte:
- EMF:

☐ Learning Objective: [2.4.2] Find next-order reactions during electrolysis

Key Takeaways

- ☐ Assume the current strongest oxidant runs out.
- ☐ Move to the next ______________________oxidant.
- ☐ End game scenarios:

Learning Objective: [2.4.3] Apply Faraday's laws to electroplating calculations

Study Design

"The application of Faraday's Laws and stoichiometry to determine the quantity of electrolytic reactant and product, and the current or time required to either use a particular quantity of reactant or produce a particular quantity of product."

Key Takeaways

- Equations:

- Typical Steps:

  1.

  2.

  3.

  4.

- Faraday's First Law:

- Faraday's Second Law:

- **Molar Mass: Charge Ratio** (M_r/z)

  - **Use:** ________________________ mass deposited for different metals.

  - **Formula:**

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