

Redox reactions

Chemical cells

| Primary Cell | Secondary Cell |
|---|---|
| CANNOT be recharged | CAN be recharged |
| <ul style="list-style-type: none">• Zinc Carbon cell<ul style="list-style-type: none">◦ Terminal V decreases with use◦ DO NOT dispose in fire because there is H_2 gas in it• Alkaline Manganese Cell<ul style="list-style-type: none">◦ Better Zinc Carbon• Silver Oxide Cell<ul style="list-style-type: none">◦ Small in size; fit in watches | <ul style="list-style-type: none">• Lithium ion cell• Nickel metal hydride cell• Lead acid accumulator<ul style="list-style-type: none">◦ Used in car batteries |

Nothing interesting here, just remember the few useful cases. This topic is not assessed much.

Half equation and oxidation states

What is redox?

| Oxidation | Reduction |
|---|--|
| Gaining oxygen $Zn \rightarrow ZnO$ | Losing oxygen $ZnO \rightarrow Zn$ |
| Losing hydrogen $H_2S \rightarrow S$ | Gaining hydrogen $S \rightarrow H_2S$ |
| Losing electrons $Fe \rightarrow Fe^{2+} + 2e^-$ | Gaining electrons $Na^+ + e^- \rightarrow Na$ |
| Increasing oxidation number | Decreasing oxidation number |

Oxidation number

An analogy of charge on an element if it was an ion

Rules when determining oxidation numbers:

- An element by itself (no charge) has an oxidation number of 0
- All atoms of an element inside a compound has the same oxidation number
- Sum of oxidation number of all elements in the compound/compound ion must be equal to its charge
- An ion has an oxidation number same as its charge (e.g. Fe^{2+} has O.N. +2)
- Most importantly: some elements have FIXED oxidation numbers in compounds

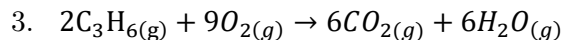
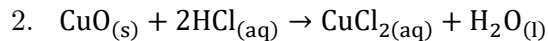
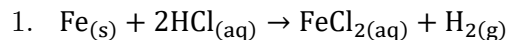
| | |
|---|----|
| All alkali metals (Group I) in their compounds | +1 |
| H in (most) compounds | +1 |
| All group II metals in their compounds | +2 |
| F in its compounds | -1 |
| Cl, Br, I in most of its compounds (see example for an exception) | +1 |
| O in most of its compounds | -2 |

Example: find the oxidation number of the underlined element in the compounds:

| | | | |
|---------------------------|-------------------------------|--------------------------|---------------------------------------|
| <u>Fe</u> SO ₄ | Zn <u>O</u> | Cl ₂ | Mg <u>S</u> O ₄ |
| H ₂ O | H ₂ O ₂ | CaBr ₂ | O ₃ |
| <u>C</u> H ₄ | NaO <u>Cl</u> | <u>N</u> Cl ₃ | <u>K</u> ₂ SO ₃ |

| Reducing agent (<i>RA</i>) | Oxidizing agent (<i>OA</i>) |
|---|--|
| Reduces others by undergoing oxidation itself. Gives electrons to others; loses electrons itself. | Oxidizes others while undergoing reduction itself. Gains electrons from others. |
| Common RA to remember: SO _{2(g)} , SO _{3²⁻(aq)} , I _(aq) , Fe ²⁺ _(aq) , H _{2(g)} , C _(s) reactive metals in elemental state (e.g. Na) | Common OA to remember: O _{2(g)} , Non reactive metal ions (e.g. Ag ⁺), Cr ₂ O _{7²⁻(aq)} , MnO _{4⁻(aq)} , Cl ₂ , F ₂ , Br ₂ , conc. H ₂ SO _{4(l)} , conc. HNO ₃ , dilute HNO ₃ , Fe ³⁺ _(aq) , H ⁺ _(aq) , OCl ⁻ |

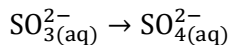
Example: Determine and explain whether the following reactions are redox reactions in reference to change in oxidation number. Identify the reducing and oxidizing agents if applicable



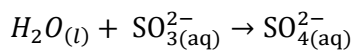
Writing and balancing redox reaction formulas

Half equations (e.g. $\text{SO}_3^{2-} \rightarrow \text{SO}_4^{2-}$)

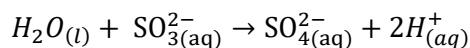
P Step 1: Balance products (ignore O, H and charge first)



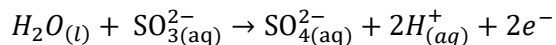
O Step 2: Balance O by adding H₂O



H Step 3: Balance H by adding H^+



e⁻ Step 4: Add electron to make charge balanced

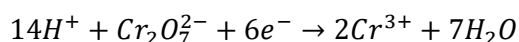


Step 5: Eliminate H^+ if redox is done in an alkaline environment (usually not required)

Examples: Write the half equation for the redox reaction for common OA and RAs.

1. Na (RA)

2. $Cr_2O_7^{2-}$ (OA)

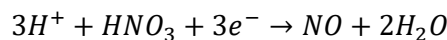


3. $Cl_{2(g)}$ (OA)

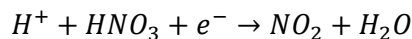
4. I⁻ (RA)

5. Conc. H_2SO_4 (OA)

6. **Dilute** HNO_3 (OA)

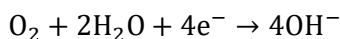


7. **Conc.** HNO_3 (OA)



8. MnO_4^- (OA/RA)

9. ** $O_{2(g)}$ (OA)



Writing FULL redox reactions

e.g. Reaction between $KMnO_4$ and $SO_{2(g)}$ (most annoying example)

Step 1: Identify your OA and RA

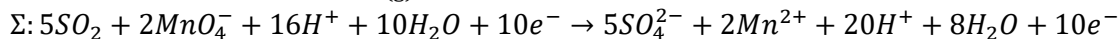
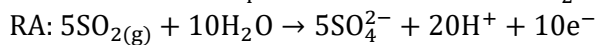
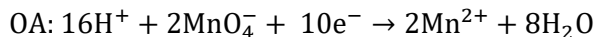
OA: MnO_4^- RA: $SO_{2(g)}$

Step 2: Write out the *half equation* for OA and RA

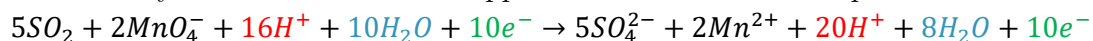
| | |
|---|-----|
| OA: $8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$ | (1) |
| RA: $SO_{2(g)} + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$ | (2) |

Step 3: Add the two equations by considering their number of electrons

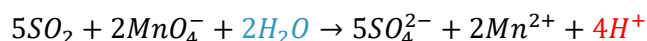
(1) * 2 + (2) * 5 → Both equations now have 10 electrons (i.e. multiply to LCM)



Step 4: Cancel out any extra reactants which appear on both sides of the equation



Final equation:



Review: Strongly recommend leaving in ionic equation

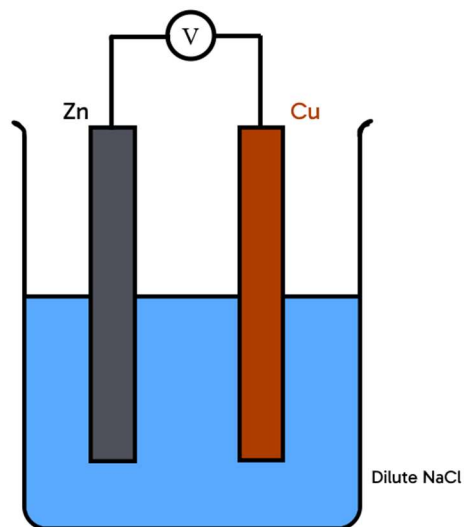
Checking your redox reaction equation:

- * Are the elements balanced?
- * Is charge balanced?

Example: Write a chemical equation for the reaction between sodium sulphite and chlorine

Chemical cells

Now that you have your fundamentals its time to do some *real* redox. The goal of a chemical cell is turn ultimately generate electricity from the chemical energy stored in the RA and OA. Consider the following typical cell:



Questions of interest:

1. Will there be a voltage?
2. How do the electrons flow
3. What is going on at the two electrodes
4. Is there anything going on in the solution

Q1: Is there a voltage:

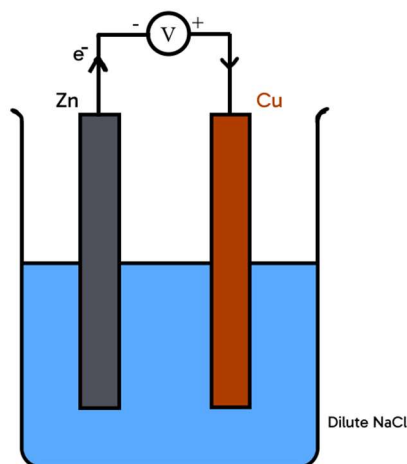
If one electrode loses electrons more readily than the other, then electrons will flow from one electrode to the other. In the figure, Zn is more reactive than Cu, so Zn loses electrons to Cu. If the **negative terminal of the voltmeter is connected with the side losing electrons** (i.e. Zn), the voltage will be **positive**. The voltage is negative otherwise.

The magnitude of the voltage depends on the difference in the positions on ECS of the anode and cathode.

Q2: Will there be a flow of electrons:

1. There must be a closed circuit
 - a. The electrolyte solution MUST conduct electricity
 - b. An external circuit must exist

Then, electrons will flow from electrode LOSING electrons through the external circuit to the electrode GAINING electrons. In this example, electrons flow from Zn to Cu through the external circuit.



Q3: what's going on at the electrodes?

First we need to identify and classify the two electrodes:

| Anode (Zn) | Cathode (Cu) |
|---|--|
| Where oxidation takes place, i.e. losing electrons by definition | Where reduction takes place, i.e. gaining electrons by definition |

Second, identify the chemical species at the cathode & anode

| | Zn Anode | Cu Cathode |
|------------------|---|---|
| Polarity | Positive | Negative |
| Chemical Species | Zn, OH ⁻ , Cl ⁻ | Cu, Na ⁺ , H ⁺ |
| | Find strongest reducing agent which preferentially undergoes oxidation | find strongest oxidizing agent which will preferentially undergo reduction |

Third, write the half equation for the species which are preferentially *oxidized/reduced*

Anode: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^{-}$

Cathode: $2\text{H}^{+} + 2\text{e}^{-} \rightarrow \text{H}_2$

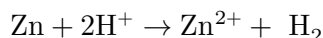
Finally, make suitable observations:

At **anode**, Zn is preferentially oxidized to form Zn²⁺. Hence, the size of the electrode decreases.

At **cathode**, H⁺ is preferentially discharged to form H₂, colourless gas bubbles are formed.

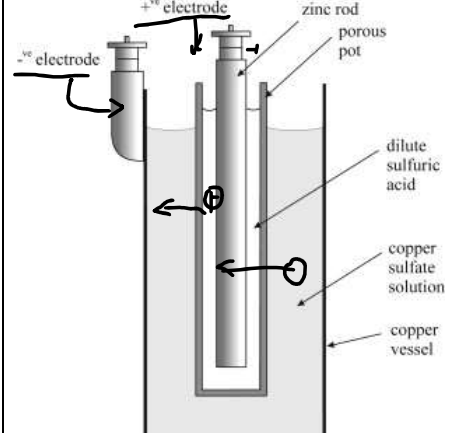
Q4: What happens in the solution?

With the 2 half equations of the reactions happening at cathode and anode, write the *full equation*



Make suitable observations. Here H^+ is consumed, so OH^- becomes in excess, so the solution turns alkaline. (if an indicator exists)

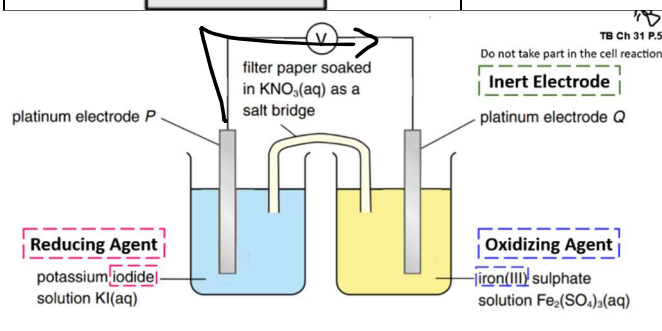
Other forms of the chemical cell:



Porous/Daniel Cell:

- Anode: Zn Rod
- Cathode: Cu container vessel
- Function of porous device: allow mobile ions to flow through it to prevent build up of excess charges. Separates the two electrolytic solutions.

Problem solving principle is still the same. Identify anode and cathode by determining where electrons are lost more readily. Then apply the remaining steps.



Half-cell, a very frequently assessed cell setup

Half cell

- Inert electrodes may be used, though they do not take part in the reaction
- Instead, cathode and anode are determined by the electrolyte solution in both of the cells
- Anode: Cell containing RA. Cathode: Cell containing OA.
- Electrons flow from anode cell through the external circuit to the cathode cell
- Salt bridge
 - Completing/Closing the circuit so mobile ions can flow
 - Must be soaked with an electrolyte
 - The electrolyte should not react with substances found in either of the cell

More regarding choice of inert electrodes and electrolyte in salt bridge.

| | When NOT to use it |
|---------------|---|
| Pt (Platinum) | When Cl_2 is formed. This is because Cl_2 will attack Pt and corrode it. |
| Graphite | When O_2 is formed. C reacts with O_2 to form CO_2 , so the electrode is corroded. |

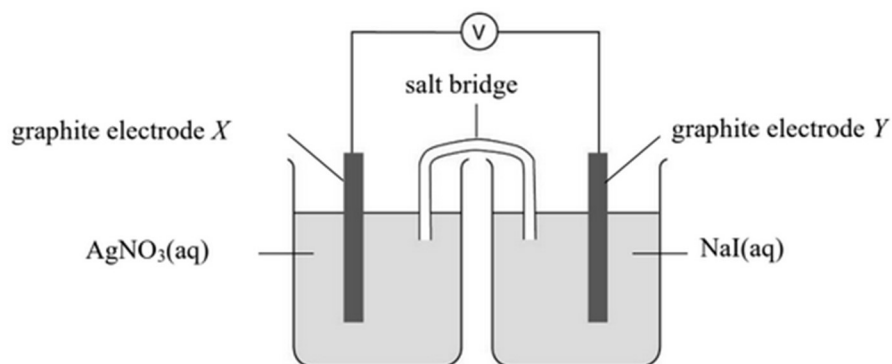
Salt bridge requirements:

The electrolyte used should not react with the substances in either half cell.

e.g. NaCl should not be used if one half cell contains Ag^+ ions as Cl^- would undergo precipitation reaction with Ag^+ to form $\text{AgCl}_{(\text{s})}$. Instead, something like $\text{NaNO}_{3(\text{aq})}$ should be used.

Example: Half cell

A student prepares a simple chemical cell as shown below.



(a) Write half equations for the reactions taking place at electrodes X and Y respectively.

X: $\text{Ag}^+ + e^- \rightarrow \text{Ag}$

Y: $2\text{I}^- \rightarrow \text{I}_2 + 2e^-$

(b) What would be observed at electrodes X and Y respectively?

X: Silvery solids are formed on the electrode.

Y: The solution turns from colourless to brown

(c) State and explain the change in the voltage measured when silver nitrate solution is replaced by acidified potassium permanganate solution.

The voltage increases. MnO_4^- is a stronger oxidizing agent Ag^+ .

(d) (i) Explain why concentrated sodium chloride solution should NOT be used as the electrolyte in the salt bridge.

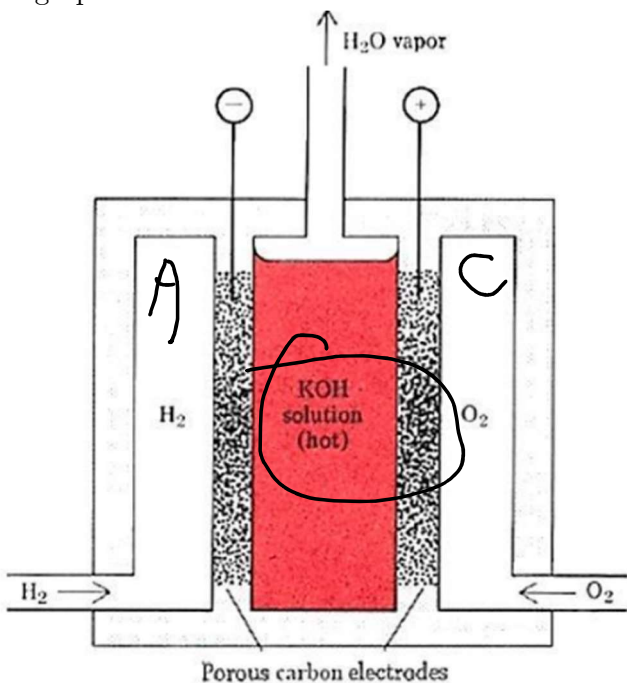
Cl^- will form precipitate with Ag^+ . $\text{Ag}_{(\text{aq})}^+ + \text{Cl}_{(\text{aq})}^- \rightarrow \text{AgCl}_{(\text{s})}$

(ii) Suggest a suitable electrolyte used in the salt bridge.

$\text{KNO}_{3(\text{aq})}$

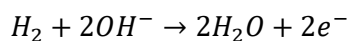
Example 2: Hydrogen oxygen fuel cell

A hydrogen-oxygen fuel cell uses concentrated potassium hydroxide solution as the electrolyte and porous graphite as electrodes.

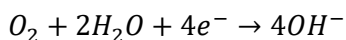


(a) Write an ionic half-equation for the reaction taking place at the,

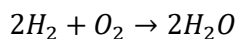
i. Anode



ii. Cathode



(b) Write a chemical equation for the overall cell reaction.



(c) Suggest ONE advantage and ONE disadvantage of using hydrogen-oxygen fuel cell.

Advantage: No air pollutants are produced

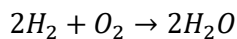
Disadvantage: It is expensive to liquefy hydrogen and oxygen gas.

(d) Hydrochloric acid can be used to replace potassium hydroxide solution. Write an ionic half-equation for the reaction taking place at the anode and cathode

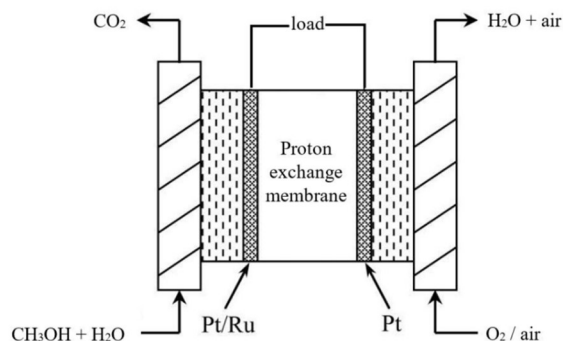
Anode: $H_2 \rightarrow 2H^+ + 2e^-$

Cathode: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ (acidic medium)

(e) Write a chemical equation for the overall cell reaction for (d).



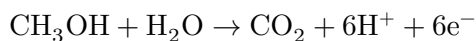
Example 3: Some electronic products run on fuel cells which use substances other than hydrogen as a fuel. One of these fuel cells is reformed methanol fuel cell using methanol (CH_3OH) as fuel.



(a) Complete the following table of the electrodes when the cell is in function.

| Electrode | Pt/Ru | Pt |
|----------------|-------|---------|
| Polarity (+/-) | - | + |
| Anode/Cathode | Anode | Cathode |

(b) Determine the mass of CH_3OH consumed when 2.55×10^{24} electrons have passes through the external circuit with the aid of a half equation.



$$\text{CH}_3\text{OH} : \text{e}^- = 1:6$$

$$\text{Mole of CH}_3\text{OH} = (1/6) * (2.55 \times 10^{24} / 6.02 \times 10^{23})$$

$$\text{Mass of CH}_3\text{OH} = 0.706 * (12+3+16+1) = 22.6\text{g}$$

(c) The compartment in the middle is known as the proton exchange membrane. **State the chemical formula of proton.**



(d) Write the overall equation of this fuel cell.

(e) Give an advantage of the reformed methanol fuel cell over the hydrogen oxygen fuel cell.

(f) Given that the percentage of volume of the methanol used is 67.0% and the density of methanol is 0.792 g cm^{-3} , calculate the concentration of methanol used in mol dm^{-3} (Relative molecular mass of $\text{CH}_3\text{OH} = 32.0$)

$$m = \rho V$$

Consider 1 dm^3 of the solution

$$\text{Volume of methanol} = 670 \text{ cm}^3$$

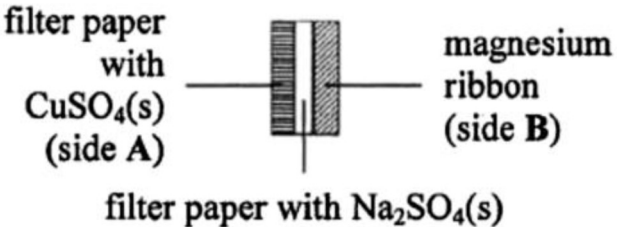
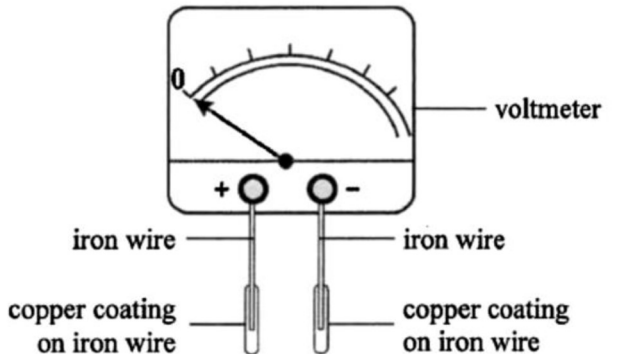
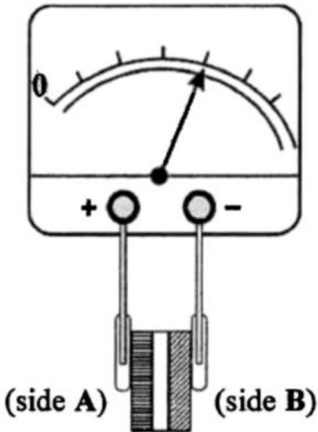
$$\text{Mass of methanol in } 1 \text{ dm}^3 = 0.792 * 670 = 530\text{g}$$

$$\text{Mole of methanol in } 1 \text{ dm}^3 = 530 / 32 = 16.6 \text{ mol}$$

$$\text{Concentration of methanol} = 16.6 \text{ mol dm}^{-3}$$

Example 4: DSE 2020 Q6

The diagrams below show the component of a chemical cell, an experimental set-up and how the pointer of the voltmeter deflects when the set-up is connected to the component.

| | |
|---|---|
| <p>Component of a chemical cell</p>  | <p>Experimental Set-up</p>  |
| <p>The pointer of the voltmeter deflects to a positive reading when a few drops of water are added to the component.</p>  <p style="text-align: center;">Diagram (1)</p> | |

- (a) Why does the pointer of the voltmeter deflect as shown when a few drops of water are added to the component? (2 marks)
- (b) Write the half equation for the change that occurs at each of the following electrodes when the pointer of the voltmeter deflects:
- Anode
 - Cathode

- (c) Consider the following design modified from Diagram (1) by only removing the copper coating at side A.

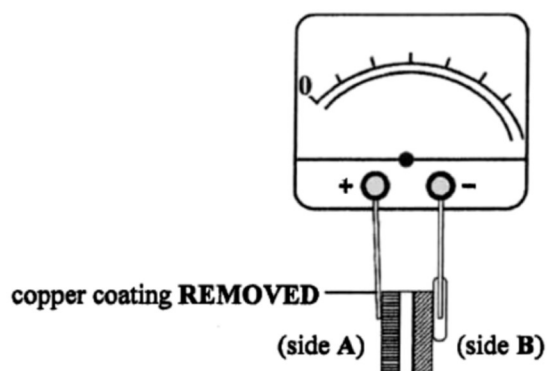


Diagram (2)

Draw on Diagram (2) the expected position of the pointer of the voltmeter when water is added to the component. (1 mark)

- (d) In the design in part (c) above, a redox reaction occurs at side A when water is added to the component. (2 marks)

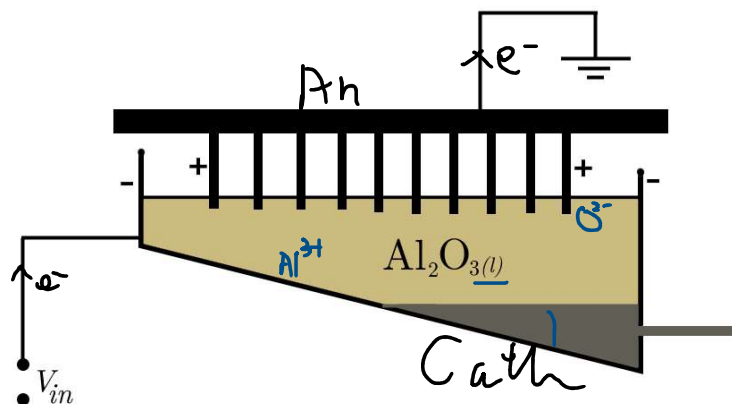
(i) Write a chemical equation for the reaction.

(ii) Name this type of reaction.

Electrolysis

General notion: Electrolysis is chemical cell, but the reactions are *forced* by the use of a battery, which acts as a source for electrons.

Example 1: Electrolysis of molten ore



Identify the anode and cathode!

| Anode (+) | Cathode (-) |
|---|--|
| Returns the electrons back to the power source/ground | Receives electrons from the power source |
| Polarity: Positive (connected to ground/positive terminal) Negative ions are attracted | Polarity: Negative (connected to negative terminal) Positive ions are attracted |
| Oxidation occurs | Reduction occurs |

Example 1: Electrolysis of molten ore

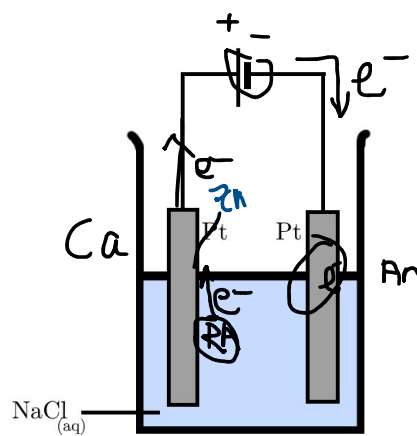
| Anode (+) | Cathode (-) |
|--|---|
| Ions present: $O^{2-}_{(l)}$ | Ions present: $Al^{3+}_{(l)}$ |
| Ions are FORCED to oxidize | Ions are FORCED to reduce |
| $2 O^{2-}_{(l)} \rightarrow O_{2(g)} + 4e^{-}$ | $Al^{3+}_{(l)} + 3e^{-} \rightarrow Al_{(s)}$ |
| Gas bubbles are evolved at the anode | Silvery solids are formed on the cathode |

Energy conversion of electrolysis:

Electrical energy \rightarrow Chemical energy

Preferential Discharging:

Consider the following example:



Step 1: Identify which one is the anode and which one is the cathode

Step 2: Identify the chemical species present at the cathode and anode

Step 3: Write the half equations for the reactions occurring at the cathode and anode

Step 4: Combine the two equations to find the change in the entire setup

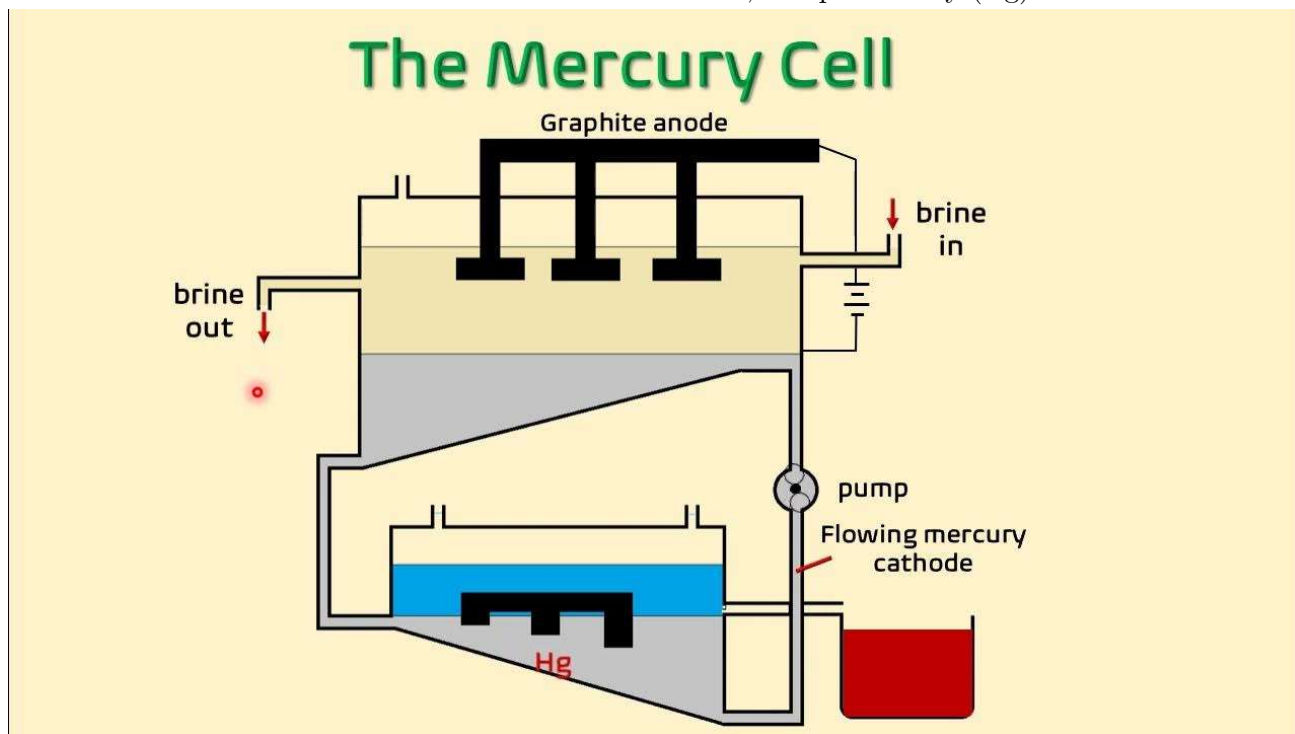
| Left (Anode (+)) | Right (Cathode (-)) |
|--|-------------------------------------|
| Species present: $Cl^{-}, OH^{-}, Zn_{(s)}$ | Species present: H^{+}, Na^{+} |

How to determine which species will be *preferentially discharged*?

- Position on ECS
 - For **oxidation (anode) reaction**, the strongest reducing agent (i.e. highest on the ECS) will be *preferentially oxidized, (discharged if applicable)*.
 - Example: OH^- will be preferentially discharged over Cl^- since OH^- is higher up on the ECS compared to Cl^- . Here I can use discharge since OH^- (charged anion) will be turned into O_2 (uncharged molecule)
 - For **reduction (cathode) reaction**, look for the strongest oxidizing agent (i.e. lowest on the ECS) to be *preferentially reduced, (or discharged if applicable)*.
 - Example: At the cathode, H^+ will be preferentially discharged over Na^+ since H^+ occupies a spot lower down on the ECS compared to Na^+
- Concentration of reactant
 - When a reactant or chemical species is sufficiently concentrated in the solution, it can preferentially undergo redox reaction EVEN WHEN a stronger OA or RA exists. This is called the concentration effect.
 - Typical species to be affected by concentration effect:
 - Halogen ions (F^- Cl^- Br^- I^-) can all undergo preferential oxidation over OH^- in water if their concentrations are sufficiently high.
 - Metal ions (up to Zn^{2+}) can undergo preferential reduction over H^+ in water, again with sufficiently high concentration.
 - Example: Suppose NaCl is concentrated in our above example
 - As $[\text{Cl}^-_{(\text{aq})}]$ is very high and much higher than that of $[\text{OH}^-_{(\text{aq})}]$, by the concentration effect, Cl^- is preferentially discharged over OH^- .
 - Note: if the solution is dilute, then OH^- is preferentially discharged as it occupies a spot higher up on the ECS than Cl^- ; OH^- is a stronger reducing agent than Cl^- .
- Nature of electrodes
 - This rarely has an effect on the products of electrolysis if the electrodes are inert (i.e. Pt or graphite)
 - If the electrodes are made of metals, especially the anode, consider the elemental metal as another species present.
 - Example: Suppose a Zn electrode is used at the anode instead of Pt.
 - At the anode, Zn **will preferentially oxidize** over OH^- and Cl^- since Zn higher up on the ECS than OH^- or Cl^- .
 - Note: I cannot use preferentially **discharge** here since Zn is gaining charge in the above redox reaction, so it is not very accurate to use that term.

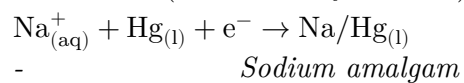
Nature of electrodes: Special case

If you could only take away one thing from “nature of electrodes” it would be this one. Normally all metal electrodes work about the same for the cathodic case, except mercury (Hg).



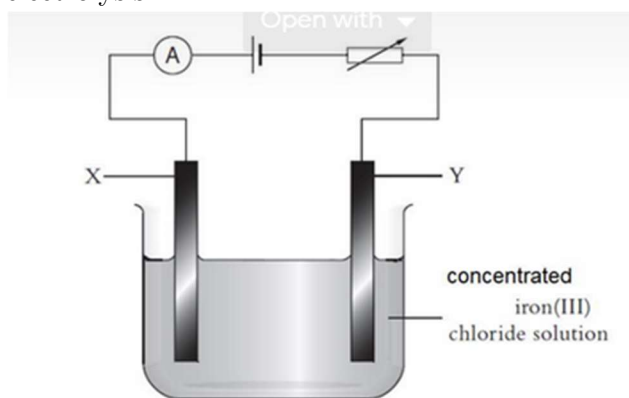
The biggest change that a mercury cathode brings is that Na^+ now preferentially discharge over H^+ despite being a much weaker oxidising agent. Due to the Hg cathode, Na^+ preferentially discharges over H^+ .

Special half equation for electrolysis in this case (with mercury cathode):



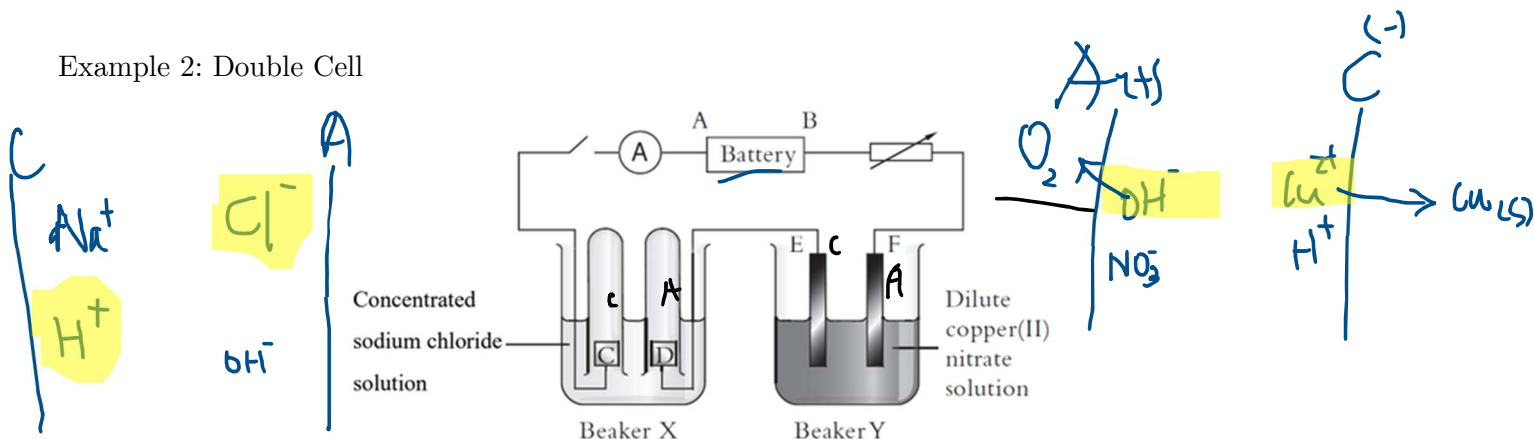
If you understood everything here, great job! Redox reactions is one of the harder topics in DSE chemistry, now you must do exercise questions to familiarize yourself with the process, as well as know more niche answers towards some specific questions.

Example 1: Electric circuit electrolysis



- Identify the cathode and anode of the setup.
Cathode: Y
Anode: X
- What are the observable changes at electrodes X and Y? (Check the electrochemical series)
X: Greenish yellow gas is observed
Y: The colour of the solution changes from yellow to pale green.
- Write half equations for the reactions occurring at X and at Y.
X: $2\text{Cl}^-_{(\text{aq})} \rightarrow \text{Cl}_{2(\text{g})} + 2\text{e}^-$
Y: $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$
- Write a chemical equation, with state symbols, for the overall reaction.
$$2\text{Fe}^{3+} + 2\text{Cl}^- \rightarrow 2\text{Fe}^{2+} + \text{Cl}_{2(\text{g})}$$

Example 2: Double Cell



An electric circuit was set up as below. Electrodes C and D are made of platinum while electrodes E and F are made of carbon. After completing the circuit, a colourless gas was formed at electrode F.

- Identify the polarities of the battery at A and B
A: Negative (-)
B: Positive (+)

b) Identify whether C, D, E, F are cathodes or anodes

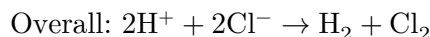
C: Cathode

D: Anode

E: Cathode

F: Anode

c) Write an overall chemical equation, with state symbols, to represent the reaction occurring in beaker X (conc. NaCl).



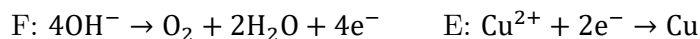
d) State the theoretical mole ratio of gases produced at electrodes C and D.

1:1

e) Suggest why platinum electrode D should be replaced by graphite.

When Cl_2 is formed. This is because Cl_2 will attack Pt and corrode it.

f) Write ionic half equations, with state symbols, for the reactions occurring at electrodes E and F.



g) State any observable changes at electrodes E and F.

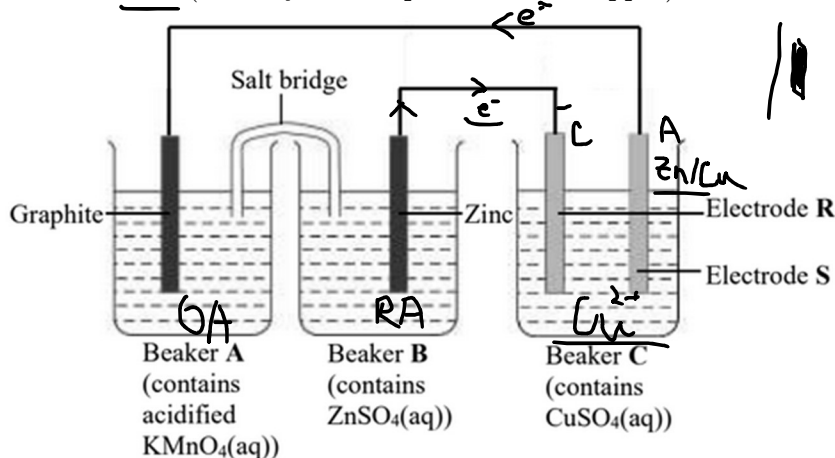
E: Brownish red solids are formed on the electrode.

The solution around electrode E becomes paler in colour.

F: Colourless gas bubbles are evolved

Example 3: Copper purification (out of scope but not the knowledge underlying it)

The diagram below shows a set-up which can purify copper. One of the electrodes in beaker C is pure copper while another one is brass (an alloy made up of zinc and copper).



a) What is the role of beakers A and B?

They act as a chemical cell to provide electricity

b) State the observations in both beakers A and B:

Beaker A: The colour of the solution turns from purple to colourless

Beaker B: The size of the Zn electrode decreases.

c) Is the electrode R pure copper electrode or brass electrode? State and explain the expected observation in electrode R. (out of scope but you should be able to do it)

R is pure copper.

At electrode R, Cu^{2+} is preferentially discharged to form $\text{Cu}_{(s)}$ over H^+ and Zn^{2+} because Cu^{2+} is lower down in the ECS than H^+ and Zn^{2+} . Hence, brownish red solids are formed on R.

d) 'The concentration of copper(II) ions in copper(II) sulphate solution remains unchanged in the above electrolysis.' Is this statement correct? Explain your answer.

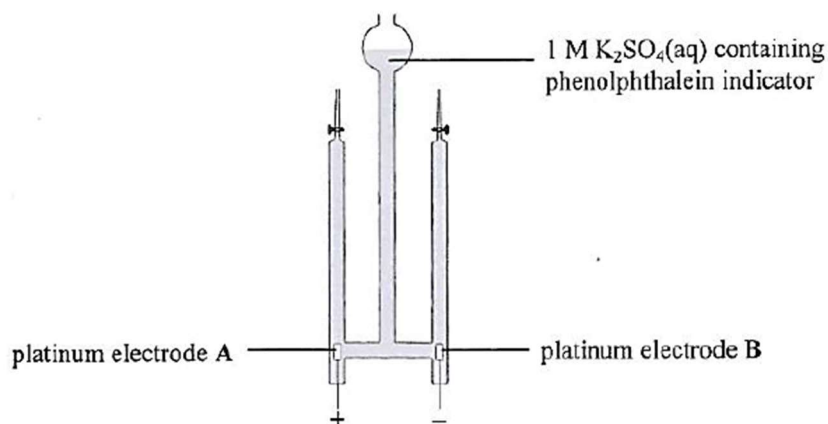
At electrode S, Zn is preferentially oxidized to form $\text{Zn}^{2+}_{(aq)}$ over Cu, OH^- , and $\text{SO}_4^{2-}_{(aq)}$ since Zn is higher up on the ECS than the other species.

Overall reaction: $\text{Zn}_{(s)} + \text{Cu}^{2+}_{(aq)} \rightarrow \text{Zn}^{2+}_{(aq)} + \text{Cu}_{(s)}$

Here $\text{Cu}^{2+}_{(aq)}$ is being consumed to form $\text{Cu}_{(s)}$ without being replenished, so $[\text{Cu}^{2+}_{(aq)}]$ decreases during electrolysis.

Example – DSE 2017 Q4, Electrolysis of ???

The diagram below shows a set-up for the electrolysis of a colourless solution of 1 M $\text{K}_2\text{SO}_4(aq)$ containing phenolphthalein indicator.



a) State, with explanation, the expected observation around the following electrodes during the electrolysis:

i) Electrode A

- ii) Electrode B
- b) Write the equation of the overall reaction in the electrolysis.
- c) Explain whether there are any changes in the expected observation around the following electrodes during the electrolysis if the 1 M $\text{K}_2\text{SO}_4(\text{aq})$ is replaced with 1 M $\text{H}_2\text{SO}_4(\text{aq})$:
 - i) Electrode A
 - ii) Electrode B

Appendix: Electrochemical series (ECS)

| | Half equation | | |
|---|--|---|--|
| | <div><div>Reduction</div><div>Oxidation</div></div> | | |
| | Oxidizing agent | Reducing agent | |
| <div>very weak oxidizing agents</div> <div>increasing oxidizing power</div> <div>increasing ease of gaining electrons</div> <div>very strong oxidizing agents</div> | <div>Ca</div> <div>$\text{K}(\text{aq}) + \text{e}^- \rightleftharpoons \text{K}(\text{s})$</div> <div>$\text{Ca}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Ca}(\text{s})$</div> <div>$\text{Na}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Na}(\text{s})$</div> <div>$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Mg}(\text{s})$</div> <div>$\text{Al}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Al}(\text{s})$</div> <div>$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Zn}(\text{s})$</div> <div>$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Fe}(\text{s})$</div> <div>$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Pb}(\text{s})$</div> <div>$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{H}_2(\text{g})$</div> <div>$\text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{SO}_2(\text{g}) + 2\text{H}_2\text{O}(\ell)$</div> <div>$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$</div> <div>$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\ell) + 4\text{e}^- \rightleftharpoons 4\text{OH}^-(\text{aq})$</div> <div>$\text{I}_2(\text{aq}) + 2\text{e}^- \rightleftharpoons 2\text{I}^-(\text{aq})$</div> <div>$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$</div> <div>$\text{NO}_3^-(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{NO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$</div> <div>$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$</div> <div>$\text{Br}_2(\text{aq}) + 2\text{e}^- \rightleftharpoons 2\text{Br}^-(\text{aq})$</div> <div>$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightleftharpoons 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\ell)$</div> <div>$\text{Cl}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-(\text{aq})$</div> <div>$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\ell)$</div> <div>$\text{S}_2\text{O}_8^{2-}(\text{aq}) + 2\text{e}^- \rightleftharpoons 2\text{SO}_4^{2-}(\text{aq})$</div> <div>$\text{F}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{F}^-(\text{aq})$</div> | <div>very strong reducing agents</div> <div>increasing ease of losing electrons</div> <div>increasing reducing power</div> <div>very weak reducing agents</div> | |

Table 30.7 The Electrochemical Series.