Intensive notes (Topic 7: Redox)

Chemical cells in daily life

- Chemical cell converts **chemical energy into electrical energy**
- Primary cells are not rechargeable while secondary cells are rechargeable.

	Class	Max. voltage	Advantage	Disadvantage
Zinc-carbon cell	Primary	1.5 V	- Low cost	- Low energy density - Poor performance in high-drained electrical devices and at low temperatures
Alkaline manganese cell	Primary	1.5 V	- Long shelf life - Wide operating temperature range - Steady voltage over time during discharge	- Low energy density - Relatively expensive for use in low-drained electrical devices
Silver oxide cell	Primary	1.6 V	- Lightweight - Small in size - Wide operating temperature range - Steady voltage over time during discharge - High energy density	- Very expensive in terms of price per unit energy stored
Lithium ion cell	Secondary	3.7 V	- High voltage - High energy density - Steady voltage over time during discharge - Can be recharged over 1200 times	- High initial cost for the cells and the battery charger - More safety concerns (e.g. may be overheated and explode if a short circuit occurs)
Nickel metal hydride cell	Secondary	1.2 V	- High energy density - Wide operating temperature range - Steady voltage over time during discharge - Can be recharged over 500 times	- High initial cost for the cells and the battery charger - High self-discharge rate
Lead-acid accumulator	Secondary	12 V	- Relatively low cost of electricity - Provide a very large current - Wide operating temperature range - Can be recharged over 500 times	- Heavy - Lead metal and lead compounds are toxic

Intensive notes (Topic 7: Redox)

Oxidation and reduction

	Oxidation (As reducing agent)	Reduction (As oxidizing agent)
Oxidation number (♂)	Increase	Decrease
Transfer of electron (OIL RIG)	Lose electrons (➔ e^-)	Gain electrons (e^- ➔)
No. of O and H atom	+O / –H	–O / +H

e.g. $Cr_2O_7^{2-} + 3SO_3^{2-} + 8H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-} + 4H_2O$

Reducing agent: SO_3^{2-}. Oxidation number of S increases from +4 to +6. SO_3^{2-} is oxidized.

- Oxidation number of S in SO_3^{2-} = +4 ➔ S exists as S^{4+} ➔ There are 16 protons and 12 electrons in S^{4+}
- Oxidation number of S in SO_4^{2-} = +6 ➔ S exists as S^{6+} ➔ There are 16 protons and 10 electrons in S^{6+}
- When SO_3^{2-} is changed to SO_4^{2-}, S loses 2 electrons. When a substance is oxidized, there must be an increase in O.N.

Oxidizing agent: $Cr_2O_7^{2-}$. Oxidation number of Cr decreases from +6 to +3. $Cr_2O_7^{2-}$ is reduced.

Oxidation number calculation

Rule 1: Oxidation number of free elements = 0

Rule 2: In compound, F = -1, Group I elements (Na, K etc.) = +1, Group II elements (Mg, Ca etc.) = +2

H = +1 (except hydride), O = -2 (except peroxide and bonded to F), Cl, Br, I = -1 (except bonded to O/F)

Rule 3: Sum of O.N. = charge of the species

Balancing ionic half equation

$Cr_2O_7^{2-}$ is reduced to Cr^{3+} (under acidic medium)	SO_3^{2-} is oxidized to SO_4^{2-} (under alkaline medium)
Step 1: Write down the reactant and product	**Step 1:** Write down the reactant and product
$Cr_2O_7^{2-} \rightarrow Cr^{3+}$	**$SO_3^{2-} \rightarrow SO_4^{2-}$**
Step 2: Balance the atoms other than O / H (i.e. Cr)	**Step 2:** Balance the atoms other than O / H (i.e. S)
$Cr_2O_7^{2-} \rightarrow \boxed{2}Cr^{3+}$	**$SO_3^{2-} \rightarrow SO_4^{2-}$**
Step 3: Balance O atom using H_2O	**Step 3:** Balance O atom using H_2O
$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + \boxed{7H_2O}$	**$SO_3^{2-} + \boxed{H_2O} \rightarrow SO_4^{2-}$**
Step 4: Balance H atom using H^+	**Step 4:** Balance H atom using H^+
$Cr_2O_7^{2-} + \boxed{14H^+} \rightarrow 2Cr^{3+} + 7H_2O$	**$SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + \boxed{2H^+}$**
Step 5: Balance charge using e^- (e^- should be added to the side with higher charge.) (Usually H^+ and e^- are placed in the same side)	**Step 5:** Balance charge using e^- (e^- should be added to the side with higher charge.) (Usually H^+ and e^- are placed in the same side)
$Cr_2O_7^{2-} + 14H^+ + \boxed{6\ e^-} \rightarrow 2Cr^{3+} + 7H_2O$	**$SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + \boxed{2e^-}$**
	Step 6: For n H^+(aq), add n OH^-(aq) in both sides.
	$SO_3^{2-} + H_2O + \boxed{2OH^-} \rightarrow SO_4^{2-} + 2H^+ + 2e^- + \boxed{2OH^-}$
	Step 7: H^+(aq) should combine with OH^-(aq) to give H_2O
	$SO_3^{2-} + H_2O + 2OH^- \rightarrow SO_4^{2-} + \boxed{2H_2O} + 2e^-$
	Step 8: Simplify the equation.
	$SO_3^{2-} + 2OH^- \rightarrow SO_4^{2-} + \boxed{H_2O} + 2e^-$

Balancing overall ionic equation (method 1)

SO_2 reacts with MnO_4^- to give SO_4^{2-} and Mn^{2+} (under acidic medium)	SO_2 reacts with MnO_4^- to give SO_4^{2-} and MnO_2 (under alkaline medium)
Step 1: Write down the ionic half equations for the changes of oxidizing agent and reducing agent	**Step 1:** Write down the ionic half equations for the changes of oxidizing agent and reducing agent
$SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$ $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	$SO_2 + 4OH^- \rightarrow SO_4^{2-} + 2H_2O + 2e^-$ $MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$
Step 2: Multiply both equations (when necessary) so that the number of electrons in both sides are equal	**Step 2:** Multiply both equations (when necessary) so that the number of electrons in both sides are equal
$(SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-) \times 5$ $\rightarrow 5SO_2 + 10H_2O \rightarrow 5SO_4^{2-} + 20H^+ + \boxed{10e^-}$ $(MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O) \times 2$ $\rightarrow 2MnO_4^- + 16H^+ + \boxed{10e^-} \rightarrow 2Mn^{2+} + 8H_2O$	$(SO_2 + 4OH^- \rightarrow SO_4^{2-} + 2H_2O + 2e^-) \times 3$ $\rightarrow 3SO_2 + 12OH^- \rightarrow 3SO_4^{2-} + 6H_2O + \boxed{6e^-}$ $(MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-) \times 2$ $\rightarrow 2MnO_4^- + 4H_2O + \boxed{6e^-} \rightarrow 2MnO_2 + 8OH^-$
Step 3: Add up the chemical species and simplify the equation (when necessary). Make sure all electrons has been cancelled out	**Step 3:** Add up the chemical species and simplify the equation (when necessary). Make sure all electrons has been cancelled out
$5SO_2 + 2MnO_4^- + 2H_2O \rightarrow 5SO_4^{2-} + 2Mn^{2+} + 4H^+$	$3SO_2 + 2MnO_4^- + 4OH^- \rightarrow 3SO_4^{2-} + 2MnO_2 + 2H_2O$

Balancing overall ionic equation (method 2)

SO_2 reacts with MnO_4^- to give SO_4^{2-} and Mn^{2+} (under acidic medium)	SO_2 reacts with MnO_4^- to give SO_4^{2-} and MnO_2 (under alkaline medium)
Step 1: Find the mole ratio of O.A. to R. A. base on changes in oxidation number	**Step 1:** Find the mole ratio of O.A. to R. A. base on changes in oxidation number
O.N. of S increases from +4 to +6 (increases by 2) O.N. of Mn decreases from +7 to +2 (decreases by 5) $\boxed{5}SO_2 + \boxed{2}MnO_4^- \rightarrow \boxed{5}SO_4^{2-} + \boxed{2}Mn^{2+}$	O.N. of S increases from +4 to +6 (increases by 2) O.N. of Mn decreases from +7 to +4 (decreases by 3) $\boxed{3}SO_2 + \boxed{2}MnO_4^- \rightarrow \boxed{3}SO_4^{2-} + \boxed{2}MnO_2$
Step 2: Balance the charges using H^+	**Step 2:** Balance the charges using OH^-
Charge on LHS = 2- Charge on RHS = 6- $5SO_2 + 2MnO_4^- \rightarrow 5SO_4^{2-} + 2Mn^{2+} + \boxed{4H^+}$	Charge on LHS = 2- Charge on RHS = 6- $3SO_2 + 2MnO_4^- + \boxed{4OH^-} \rightarrow 3SO_4^{2-} + 2MnO_2$
Step 3: Balance O and H using H_2O	**Step 3:** Balance O and H using H_2O
$5SO_2 + 2MnO_4^- + \boxed{2H_2O} \rightarrow 5SO_4^{2-} + 2Mn^{2+} + 4H^+$	$3SO_2 + 2MnO_4^- + 4OH^- \rightarrow 3SO_4^{2-} + 2MnO_2 + \boxed{2H_2O}$

Common oxidizing agents and reducing agents

Common oxidizing agent	Observation	Half-equation
$Acidified MnO_4^-(aq)	Purple solution turns colorless	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$
Acidified $Cr_2O_7^{2-}$(aq)	Orange solution turns green	$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
*Conc. H_2SO_4(l) **[Sulphuric acid]**	Colorless choking smell gas forms	$2H_2SO_4 + 2e^- \rightarrow SO_2 + SO_4^{2-} + 2H_2O$
*Conc. HNO_3(aq) **[Nitric acid]**	Brown gas forms	$NO_3^- + 2H^+ + e^- \rightarrow NO_2 + H_2O$
*Dil. HNO_3(aq) **[Nitric acid]**	#Colorless gas forms	$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$
Cl_2(aq) **[Chlorine]**	Pale green solution turns colorless	$Cl_2 + 2e^- \rightarrow 2Cl^-$
Br_2(aq) **[Bromine]**	Brown solution turns colorless	$Br_2 + 2e^- \rightarrow 2Br^-$
I_2(aq) **[Iodine]**	Brown solution turns colorless	$I_2 + 2e^- \rightarrow 2I^-$
H^+(aq)	Colorless gas forms	$2H^+ + 2e^- \rightarrow H_2$
Ag^+(aq)	Silvery solid forms	$Ag^+ + e^- \rightarrow Ag$
Cu^{2+}(aq)	Reddish brown solid forms, Blue solution turns colorless	$Cu^{2+} + 2e^- \rightarrow Cu$
^Fe^{3+}(aq)	Yellow solution turns green	$Fe^{3+} + e^- \rightarrow Fe^{2+}$
O_2(g)	---	$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$ $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

$Under neutral / alkaline conditions, $KMnO_4$ will be reduced to MnO_2 which is a brown solid.

*Oxidizing acids do NOT undergo redox reaction when they react with base / carbonate / hydrogencarbonate

^Fe^{2+} will be converted to Fe(s) in displacement reaction only

#Colorless gas will turn brown in air because NO(g) will be oxidized by O_2(g) in air to form NO_2(g).

$$2NO + O_2 \rightarrow 2NO_2$$

Common reducing agent	Observation	Half-equation
Metals	---	$M \rightarrow M^{n+} + n\ e^-$
SO_2(g) **[Sulphur dioxide]**	No observable change	$SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$
#SO_3^{2-}(aq) **[Sulphite ion]**	No observable change	$SO_3^{2-} + H_2O \rightarrow SO_4^{2-} + 2H^+ + 2e^-$
Fe^{2+}(aq)	Green solution turns yellow	$Fe^{2+} \rightarrow Fe^{3+} + e^-$
I^-(aq) **[Iodide ion]**	*Colorless solution turns brown	$2I^- \rightarrow I_2 + 2e^-$
Br^-(aq) **[Bromide ion]**	Colorless solution turns brown	$2Br^- \rightarrow Br_2 + 2e^-$
^Cl^-(aq) **[Chloride ion]**	Greenish yellow gas forms	$2Cl^- \rightarrow Cl_2 + 2e^-$
H_2(g)	---	$H_2 \rightarrow 2H^+ + 2e^-$

#SO_3^{2-} also reacts with acid to give SO_2(g). This is NOT a redox reaction.

$$SO_3^{2-} + 2H^+ \rightarrow SO_2 + H_2O$$

^Cl^- is a weak reducing agent. However, it can still be oxidized by acidified $KMnO_4$(aq).

Oxidizing acids

	Dil. HCl	Dil. HNO_3	Dil. H_2SO_4	Conc. HCl	Conc. HNO_3	Conc. H_2SO_4
Zn	*H_2	#NO	*H_2	*H_2	#NO_2	#SO_2
Cu	^---	#NO	^---	^---	#NO_2	#SO_2
$ZnCO_3$	$$CO_2$	$$CO_2$	$$CO_2$	$$CO_2$	$$CO_2$	$$CO_2$
$CuCO_3$	$$CO_2$	$$CO_2$	$$CO_2$	$$CO_2$	$$CO_2$	$$CO_2$
C	---	---	---	---	NO_2 and CO_2	SO_2 and CO_2
S	---	---	---	---	NO_2 and SO_2	SO_2

* Dilute HCl, dilute H_2SO_4 and concentrated HCl are typical acids. They react with Zn to give $H_2(g)$.

^ Dilute HCl, dilute H_2SO_4 and concentrated HCl are typical acids. They do not react with metal with reactivity ≤ Cu.

\# Dilute HNO_3, concentrated HNO_3 and concentrated H_2SO_4 are oxidizing acids. They react with reducing agents (metals) to give NO(g), $NO_2(g)$ and $SO_2(g)$ respectively.

$ Metal carbonates are NOT reducing agent. When they are added to oxidizing acids, they react with hydrogen ion in these acids only to give $CO_2(g)$.

Chlorine

*1. $Cl_2(g)$ dissolves in cold dilute NaOH(aq) to give NaCl(aq) and NaOCl(aq) (a main component in chlorine bleach).

$$\mathbf{Cl_2(g) + 2NaOH(aq) \rightarrow NaCl(aq) + NaOCl(aq) + H_2O(l)}$$

*2. $Cl_2(g)$ dissolves in hot concentrated NaOH(aq) to give NaCl(aq) and $NaClO_3(aq)$

$$\mathbf{3Cl_2(g) + 6NaOH(aq) \rightarrow 5NaCl(aq) + NaClO_3(aq) + 3H_2O(l)}$$

*3. $Cl_2(g)$ dissolves in water to give HCl(aq) and HOCl(aq)

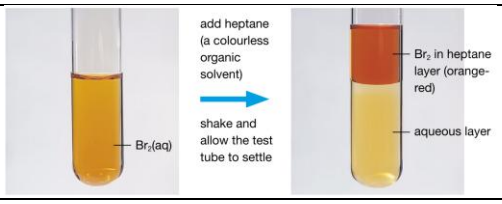
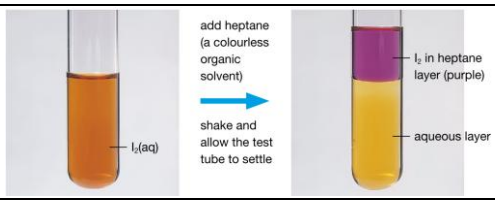
$$\mathbf{Cl_2(g) + H_2O(l) \rightleftharpoons HCl(aq) + HOCl(aq)}$$

*These reactions are the examples of disproportionation as Cl_2 is simultaneously oxidized and reduced.

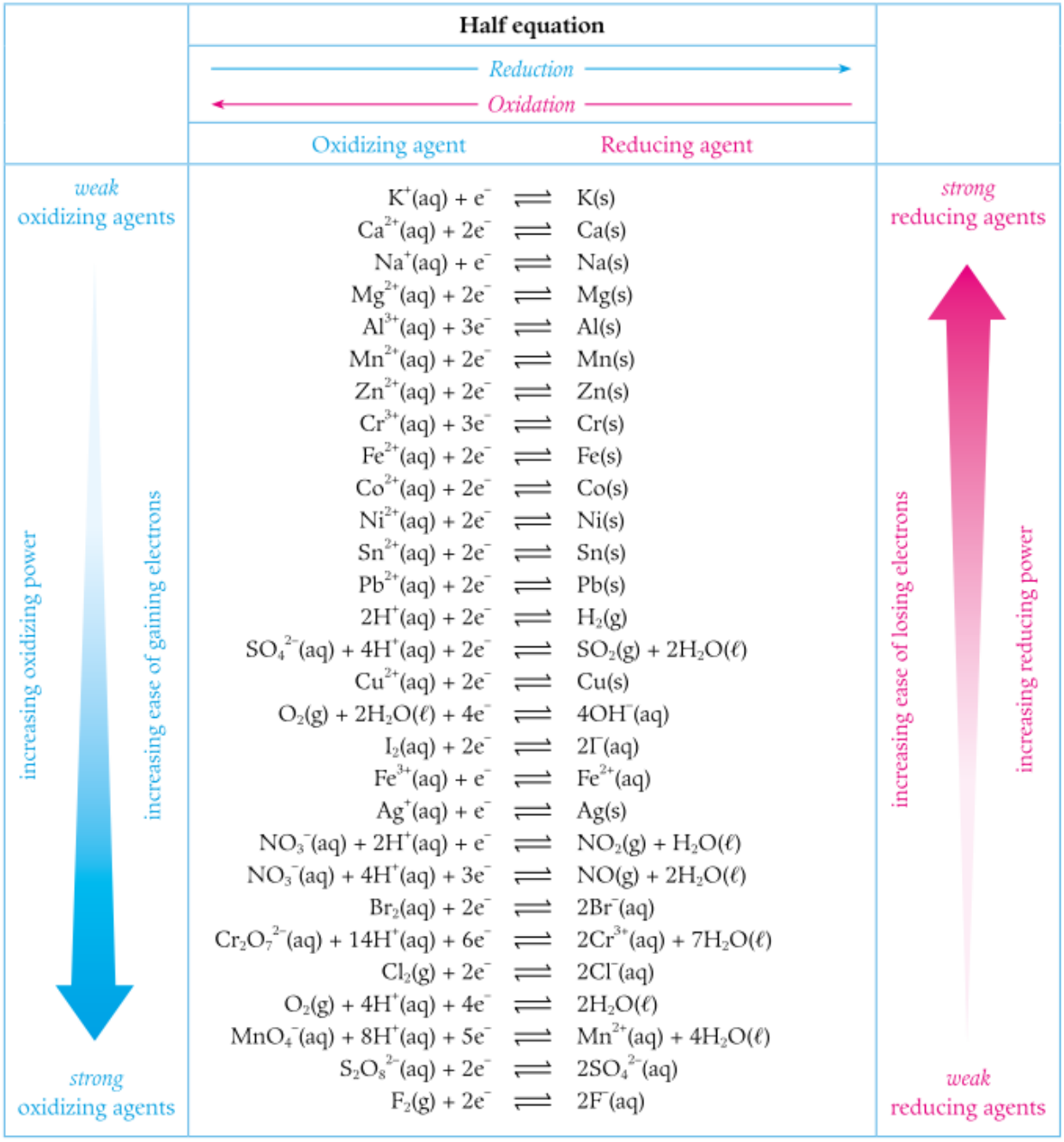
4. When acid is added to bleach, $Cl_2(g)$ will be produced. (Don't try to mix bleach and acid as this will produce toxic $Cl_2(g)$!)

$$\mathbf{Cl^-(aq) + OCl^-(aq) + 2H^+(aq) \rightarrow Cl_2(g) + H_2O(l)}$$

5. $Cl_2(g)$ can react with $Br^-(aq)$ to give a $Br_2(aq)$ which is a brown solution. Br_2 can further dissolve in heptane to give an orange-red solution. On the other hand, $Cl_2(g)$ can also react with $I^-(aq)$ to give a $I_2(aq)$ which is also a brown solution. I_2 can further dissolve in heptane to give a purple solution.

$Cl_2(aq) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$	$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$
	

Electrochemical series



- $KMnO_4$ should be acidified by dilute $H_2SO_4(aq)$, but NOT dilute HCl(aq) as MnO_4^- is a stronger oxidizing agent than Cl_2 but a weaker oxidizing agent than $S_2O_8^{2-}$.

- Zn can undergo displacement with $CuSO_4$ while Cu cannot undergo displacement with $ZnSO_4$ as Zn is a stronger reducing agent than Cu (Or Cu^{2+} is a stronger oxidizing agent than Zn^{2+}.)

- Metals with reactivity ≤ Cu does not react with typical acids. As these metals are weaker reducing agent than H_2.

- Fe^{2+} is very difficult to be reduced to Fe (e.g. displacement) as Fe^{2+} is a very weak oxidizing agent.

Intensive notes (Topic 7: Redox)

Chemical cell

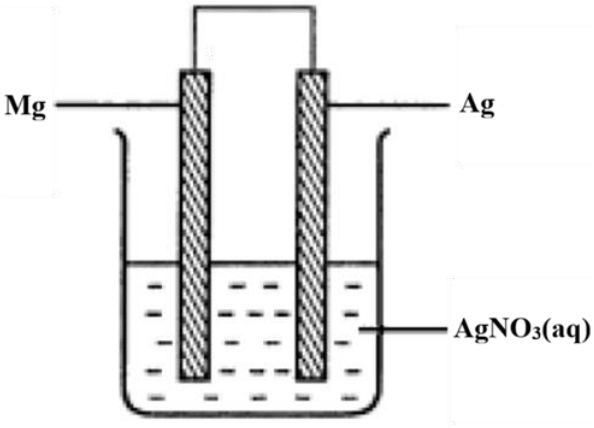
Electrolyte is a substance which conducts electricity in molten / aqueous state

Rules to be remembered:

1. Electrons ALWAYS flow from **negative** to **positive**.
   (**negative** electrode **losing** electrons, **positive** electrode **gaining** electrons)
2. OIL RIG
   (**O**xidation **I**s **L**osing electrons, **R**eduction **I**s **G**aining electrons)
3. Red cat an ox
   (**Red**uction = **Cat**hode, **An**ode = **Ox**idation)

Negative	Positive
Losing electrons	Gaining electrons
[➔ e^-]	[e^- ➔]
Oxidation	Reduction
Anode	Cathode
Reducing agent	Oxidizing agent
O.N. increases	O.N. decreases

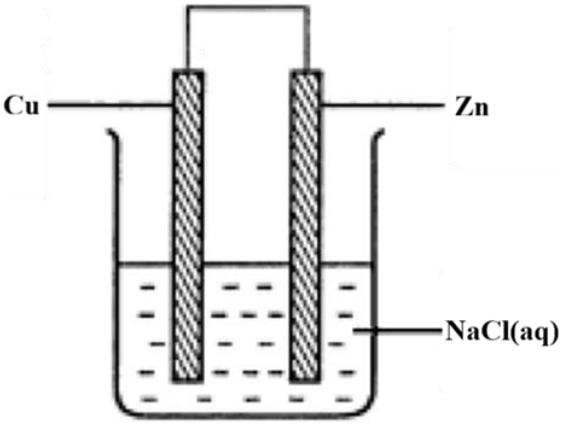
Example 1



	Mg electrode	Ag electrode
Half-equation	$Mg \rightarrow Mg^{2+} + 2e^-$	$Ag^+ + e^- \rightarrow Ag$
Polarity	Negative	Positive
	(e^- flows from – to +)	
Anode / cathode	Anode (Mg is oxidized)	Cathode (Ag^+ is reduced)
Observable change	Size of Mg electrode decreases	Size of Ag electrode increases

- Mg will undergo displacement reaction with $AgNO_3$ directly. A layer of Ag forms on the Mg electrode. The potential difference between two electrodes decreases. Hence, the voltage of the cell drops rapidly.

Example 2

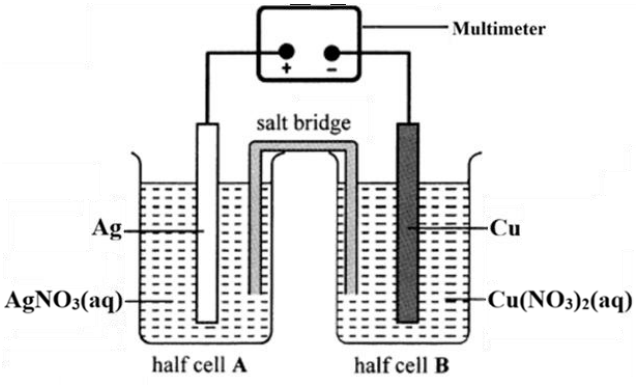


	Cu electrode	Zn electrode
Half-equation	$*2H^+ + 2e^- \rightarrow H_2$	$Zn \rightarrow Zn^{2+} + 2e^-$
Polarity	Positive	Negative
	(e^- flows from – to +)	
Anode / cathode	Cathode (H^+ is reduced)	Anode (Zn is oxidized)
Observable change	Colorless gas bubbles evolve	Size of Zn electrode decreases

* In NaCl(aq), H^+(aq) and Na^+(aq) are cations that can accept electrons. Since Na^+(aq) is a very weak oxidizing agent, H^+(aq) ions will gain electrons to form H_2(g).

- H_2(g) produced at Cu electrode hinders the conduction of electricity (as hydrogen gas forms a non-conductive layer on the surface of the Cu electrode). Hence, the voltage of the cell drops rapidly.

Example 3



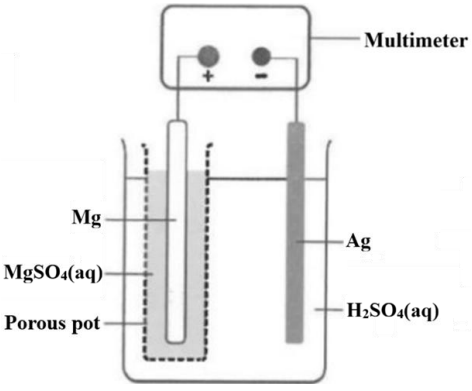
	Ag electrode	Cu electrode
Half-equation	$Ag^+ + e^- \rightarrow Ag$	$Cu \rightarrow Cu^{2+} + 2e^-$
Polarity	Positive	Negative
	(e^- flows from – to +)	
Anode / cathode	Cathode (Ag^+ is reduced)	Anode (Cu is oxidized)
Observable change	Size of Ag electrode increases	Size of Cu electrode decreases

- Since Ag (a **positive** electrode) is connected to the **positive** terminal of the multimeter, it should show a **positive** reading.

- TWO functions of salt bridge*:
  1. It completes the circuit by allowing ions to move from one half cell into another.
  2. It provides ions to balance the charges in the solutions of the two half cells.
  * Salt bridge in the above set-up can be prepared by soaking a filter paper with saturated $KNO_3(aq)$ (As K^+ and NO_3^- will NOT undergo any redox reaction / precipitation)

- In half cell A, the conversion of Ag^+ to Ag creates excess negative charge in the solution. K^+ from the salt bridge will migrate to half cell A to balance the excess negative charge. In half cell B, the conversion of Cu to Cu^{2+} creates excess positive charge in the solution. NO_3^- from the salt bridge will migrate to half cell B to balance the excess positive charge.

Example 4



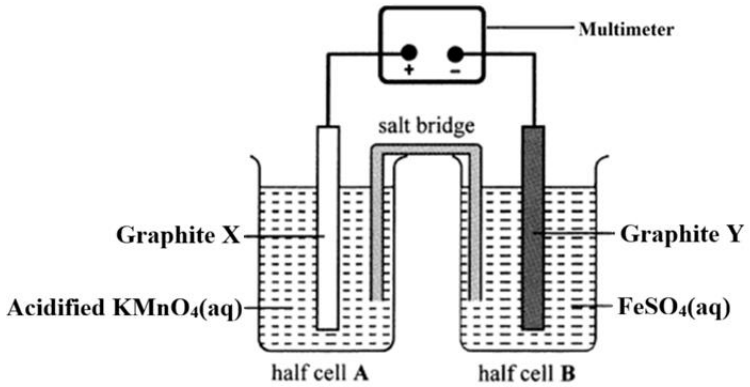
	Mg electrode	Ag electrode
Half-equation	$Mg \rightarrow Mg^{2+} + 2e^-$	$2H^+ + 2e^- \rightarrow H_2$
Polarity	Negative	Positive
	(e^- flows from – to +)	
Anode / cathode	Anode (Mg is oxidized)	Cathode (H^+ is reduced)
Observable change	Size of Mg electrode decreases	Colorless gas bubbles evolve

- Since Mg (a **negative** electrode) is connected to the **positive** terminal of the multimeter, it should show a **negative** reading.

- TWO functions of porous pot:
  1. It prevents direct mixing of the two electrolytes
  2. It completes the circuit by allowing ions to move from one electrolyte into another.

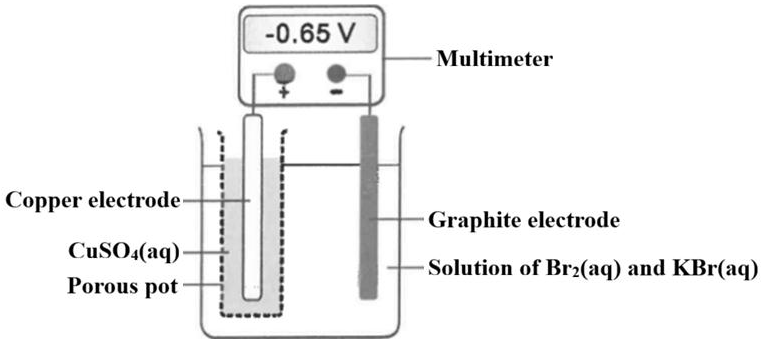
- Small holes in porous pot allow excess Mg^{2+} to move out of the device and SO_4^{2-} to move into the device. Thus, excess positive and negative charges will not accumulate in the two half cells.

Example 5



	Half-equation + Observation	Cathode / Anode	Positive / Negative
Graphite X	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ Observation: Solution turns from purple to colorless	Cathode (MnO_4^- is reduced)	Positive (e^- flows from – to +)
Graphite Y	$Fe^{2+} \rightarrow Fe^{3+} + e^-$ Observation: Solution turns from green to yellow	Anode (Fe^{2+} is oxidized)	Negative (e^- flows from – to +)

Example 6



- A negative multimeter reading indicates that the positive electrode of the chemical cell is connected to the negative pole of multimeter. Therefore copper is the negative electrode and graphite is the positive electrode.

  ➔ Electron flows from copper electrode (-) to graphite electrode (+)

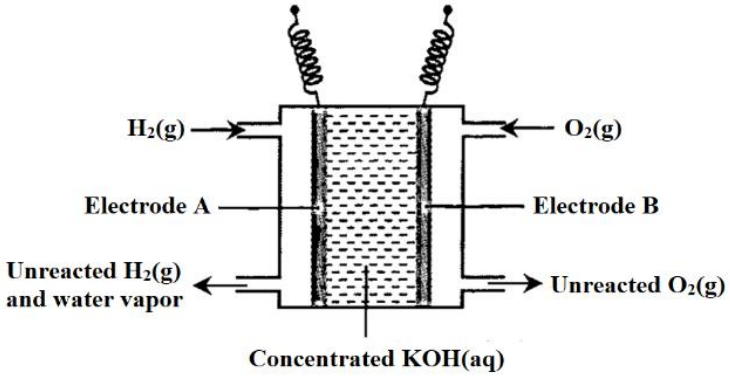
	Half-equation + Observation	Cathode / Anode	Positive / Negative
Copper	$Cu \rightarrow Cu^{2+} + 2e^-$ (NOT $Cu^{2+} + 2e^- \rightarrow$ Cu as copper electrode loses electrons) Observations 1: Color intensity of blue solution increases Observations 2: Size of copper electrode decreases	Anode (Cu is oxidized)	Negative (e^- flows from – to +)
Graphite	$Br_2 + 2e^- \rightarrow 2Br^-$ (NOT $2Br^- \rightarrow Br_2 + 2e^-$ as graphite electrode gains electrons) Observation: Color intensity of brown solution decreases	Cathode (Br_2 is reduced)	Positive (e^- flows from – to +)

- If the solution of Br_2(aq) and KBr(aq) is replaced by a solution of I_2(aq) and KI(aq) (other conditions remain unchanged):

  ➔ The multimeter reading becomes less negative

  ➔ As I_2 is less reactive / weaking oxidizing agent than Br_2. (Reactivity of Group VII elements decreases down the group)

Fuel cell



- Fuel cells convert chemical energy (in H_2 and O_2) to electrical energy

- Half equation at electrode A: $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$
  Half equation at electrode B: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$
  (Since the electrolyte used is alkaline KOH, H^+ should NOT appear in the half equations)
  Overall equation: $2H_2 + O_2 \rightarrow 2H_2O$

- If concentrated KOH(aq) is replaced by concentrated H_3PO_4(aq):
  Half equation at electrode A: $H_2 \rightarrow 2H^+ + 2e^-$
  Half equation at electrode B: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

- TWO functions of porous electrode (made of Pt / Ni) A and B.
  1. Allow the flow of $H_2(g)$, $O_2(g)$ and $H_2O(g)$ into and out of the compartments
  2. Catalyze the reaction between hydrogen and oxygen

- TWO advantages and TWO disadvantages of using hydrogen-oxygen fuel cell
  Advantage 1: It has high energy conversion efficiency
  Advantage 2: The only product of this cell is water which is non-polluting
  Advantage 3: A hydrogen-oxygen fuel cell can operate continuously when there is a continuous supply of H_2 and O_2
  Advantage 4: It has a higher energy-to-mass ratio (i.e. energy density)
  Disadvantage 1: It is not easy to store and transport gaseous hydrogen and oxygen
  Disadvantage 2: Using platinum and liquefaction of hydrogen and oxygen are expensive
  Disadvantage 3: Hydrogen is flammable and explosive
  Disadvantage 4: Hydrogen gas is mainly produced from fossil fuels. The use of fuel cells still relies on fossil fuels.

- THREE applications of hydrogen-oxygen fuel cell.
  1. Remote power sources of spacecraft
  2. Backup power sources
  3. Generating electricity for electric vehicles.

Electrolysis

- Electrolytic cells convert **electrical energy to chemical energy**

Rules to be remembered:

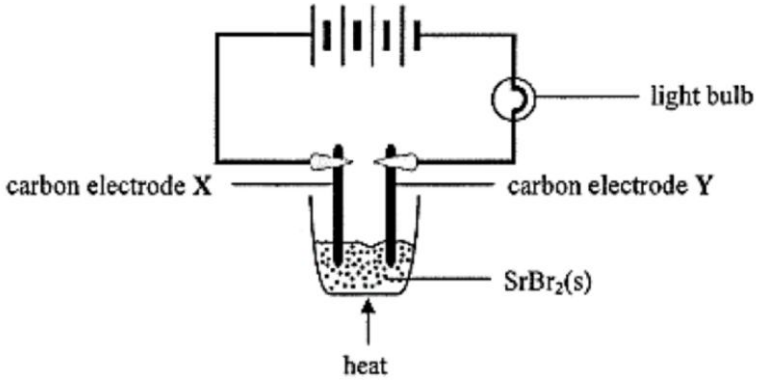
1. Electrons ALWAYS flow from **negative pole of BATTERY** to **positive pole of BATTERY.**
2. OIL RIG (**O**xidation **I**s **L**osing electrons, **R**eduction **I**s **G**aining electrons)
3. Red cat an ox (**Red**uction = **Cat**hode, **An**ode = **Ox**idation)
4. Electrode connected to positive pole of battery
   ➔ Positive electrode in electrolytic cell
   Electrode connected to negative pole of battery
   ➔ Negative electrode in electrolytic cell

Positive	Negative
Losing electrons	Gaining electrons
[➔ e^-]	[e^- ➔]
Oxidation	Reduction
Anode	Cathode
Reducing agent	Oxidizing agent
O.N. increases	O.N. decreases

Common products obtained in electrolysis

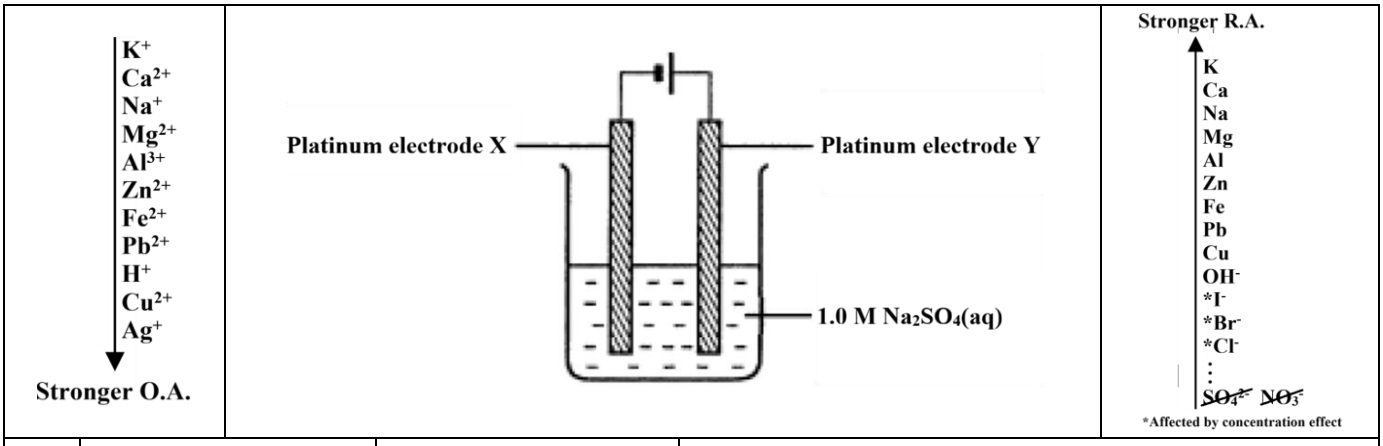
Product	Observable change	Product	Observable change
$H_2(g)$	Colorless gas	$O_2(g)$	Colorless gas
Metals (except Cu(s), Hg(l))	Silvery grey solid	Cu(s)	Reddish brown solid
$I_2(aq)$	**Brown solution**	$I_2(g)$	Violet gas
$Br_2(aq)$	Brown solution	$Br_2(g)$	Brown gas
$Cl_2(g)$	Yellowish green gas	**$NO_2(g)$ / $SO_2(g)$ / NO(g)**	**Wake up, stop dreaming**

Electrolytic cell using molten electrolytes

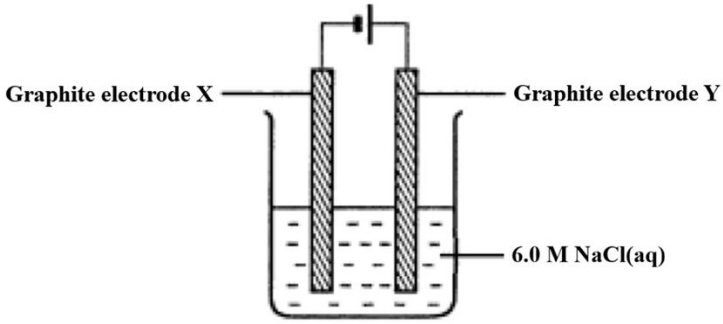


	Cathode / Anode	Positive / Negative	Half-equation + Observation
X	Anode (Lose e^-, oxidation)	Positive (Connected to +)	$2Br^- \rightarrow Br_2 + 2e^-$ (Br^- ions are attracted to positive electrode X) Observation: Brown gas evolved
Y	Cathode (Gain e^-, reduction)	Negative (Connected to -)	$Sr^{2+} + 2e^- \rightarrow Sr$ (Sr^{2+} ions are attracted to negative electrode Y) Observation: Silvery grey solid forms

Electrolytic cell using aqueous electrolytes



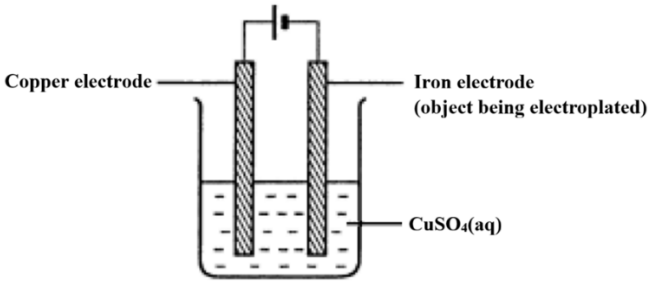
	Cathode / Anode	Positive / Negative	Half-equation + Observation
X	Cathode (Gain e^-, reduction)	Negative (Connected to -)	$2H^+ + 2e^- \rightarrow H_2$ H^+ is a stronger oxidizing agent than Na^+ H^+ ions are preferentially discharged to give H_2 Observation: Colorless gas evolved
Y	Anode (Lose e^-, oxidation)	Positive (Connected to +)	$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ OH^- is a stronger reducing agent than SO_4^{2-} OH^- ions are preferentially discharged to give O_2 Observation: Colorless gas evolved
Change in electrolyte: Concentration of Na_2SO_4(aq) increases (As H^+ and OH^- (i.e. H_2O) is electrolyzed)			



	Cathode / Anode	Positive / Negative	Half-equation + Observation
X	Cathode (Gain e^-, reduction)	Negative (Connected to -)	$2H^+ + 2e^- \rightarrow H_2$ H^+ is a stronger oxidizing agent than Na^+ H^+ ions are preferentially discharged to give H_2 Observation: Colorless gas evolved
Y	Anode (Lose e^-, oxidation)	Positive (Connected to +)	$2Cl^- \rightarrow Cl_2 + 2e^-$ $[Cl^-]$ is much higher than $[OH^-]$ Cl^- ions are preferentially discharged to give Cl_2 Observation: Yellowish green gas evolved
Change in electrolyte: Solution becomes NaOH(aq) (As H^+ and Cl^- are discharged, leaving Na^+ and OH^-)			

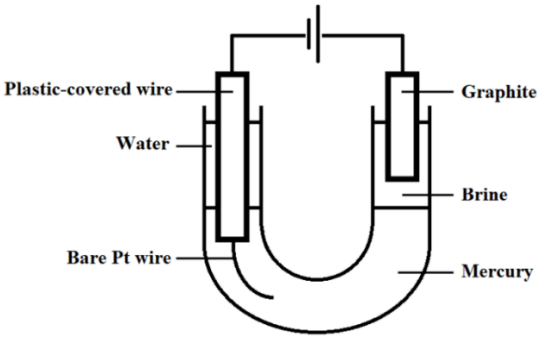
- As chlorine gas is produced, graphite CANNOT be replaced by platinum as chlorine gas will attack platinum.

Electroplating



	Cathode / Anode	Positive / Negative	Half-equation + Observation
Cu	Anode (Lose e^-, oxidation)	Positive (Connected to +)	$Cu \rightarrow Cu^{2+} + 2e^-$ Cu is a stronger reducing agent than OH^- and SO_4^{2-}. Cu oxidizes and dissolves to give Cu^{2+} Observation: Size of electrode decreases
Fe	Cathode (Gain e^-, reduction)	Negative (Connected to -)	$Cu^{2+} + 2e^- \rightarrow Cu$ Cu^{2+} is a stronger oxidizing agent than H^+. Cu^{2+} ions are preferentially discharged to give Cu. Observation: Size of electrode increases.
Change in electrolyte: No change. (As the rate of formation of Cu^{2+} at anode and the rate of consumption of Cu^{2+} at cathode are the same)			

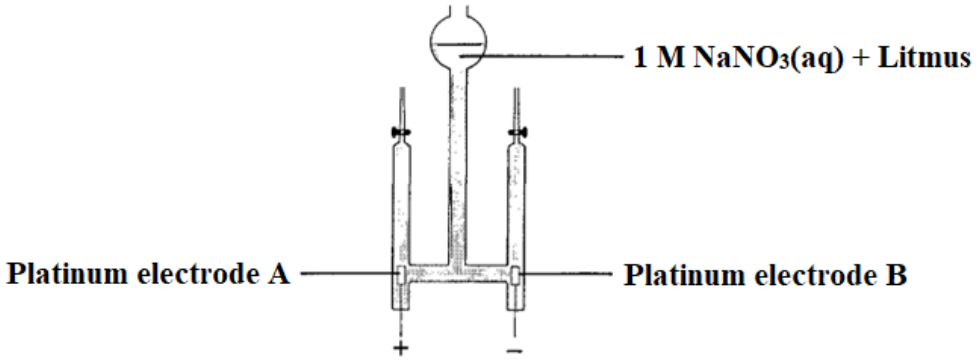
Electrolysis of brine (concentrated NaCl(aq)) using mercury cathode



	Cathode / Anode	Positive / Negative	Half-equation + Observation
Graphite	Anode (Lose e^-, oxidation)	Positive (Connected to +)	$2Cl^- \rightarrow Cl_2 + 2e^-$ $[Cl^-]$ is much higher than $[OH^-]$ Cl^- ions are preferentially discharged to give Cl_2 Observation: Yellowish green gas evolved
Mercury	Cathode (Gain e^-, reduction)	Negative (Connected to -)	$Na^+ + e^- + Hg \rightarrow Na/Hg$ The use of Hg cathode favors the discharge of Na^+. Na^+ ions are preferentially discharged to give Na metal, which dissolves in mercury immediately to give Na/Hg (sodium amalgam). Observation: No observable change

- Water is added to the left arm to prevent the evaporation of toxic mercury.

Electrolytic solution containing pH indicator



	Cathode / Anode	Positive / Negative	Half-equation + Observation
A	Anode (Lose e^-, oxidation)	Positive (Connected to +)	$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ $OH^-(aq)$ ion is a stronger reducing agent than $SO_4^{2-}(aq)$ ion. $OH^-(aq)$ ions are preferentially discharged to give $O_2(g)$. H_2O molecules ionize continuously to replace the discharged $OH^-(aq)$ ions. $H^+(aq)$ ions accumulate around electrode A and the solution around electrode A becomes acidic Observation 1: Colorless gas evolved Observation 2: Solution turns red ($[OH^-] < [H^+]$)
B	Cathode (Gain e^-, reduction)	Negative (Connected to -)	$2H^+ + 2e^- \rightarrow H_2$ $H^+(aq)$ is a stronger oxidizing agent than $Na^+(aq)$ ion. $H^+(aq)$ ions are preferentially discharged to give $H_2(g)$. H_2O molecules ionize continuously to replace the discharged $H^+(aq)$ ions. $OH^-(aq)$ ions accumulate around electrode B and the solution around electrode B becomes alkaline Observation 1: Colorless gas evolved Observation 2: Solution turns blue ($[H^+] < [OH^-]$)

- If 1 M $NaNO_3$ is replaced by 1 M H_2SO_4:

  Solution around electrode A will remain red as 1 M H_2SO_4 is acidic.

  Solution around electrode B will remain red. In 1 M H_2SO_4, $[H^+]$ is much higher than $[OH^-]$. Solution around electrode B will always be acidic despite H^+ ions are discharged.

- If 1 M $NaNO_3$ is replaced by 1 M NaOH:

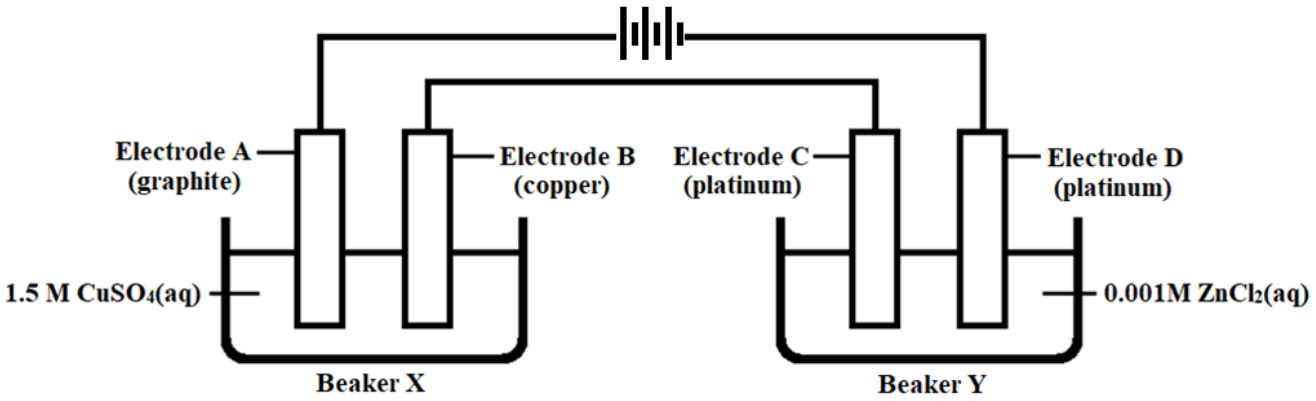
  Solution around electrode A will remain blue. In 1 M NaOH, $[OH^-]$ is much higher than $[H^+]$. Solution around electrode B will always be alkaline despite OH^- ions are discharged.

  Solution around electrode A will remain blue as 1 M NaOH is alkaline.

- If 1 M $NaNO_3$ is replaced by 6 M NaCl

  Solution around electrode A will turn red and then colorless quickly. $[Cl^-]$ is much higher than $[OH^-]$. Cl^- ions are preferentially discharged to give Cl_2 which dissolves in water immediately to give HCl and HOCl. HCl turns litmus red as it is acidic. HOCl turns litmus colorless as it is a bleaching agent.

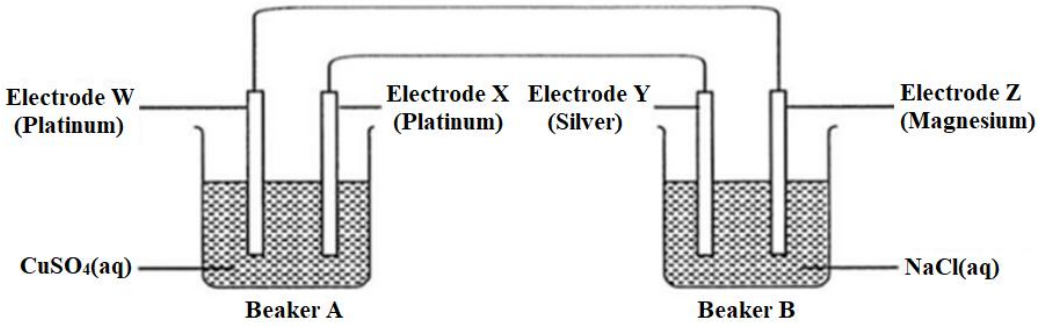
Multiple electrolytic cells



- In the above set-up, both beaker X and beaker Y are electrolytic cell.

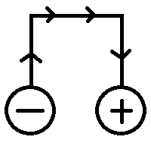
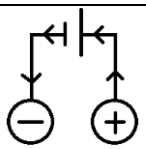
	Cathode / Anode	Positive / Negative	Half-equation + Observation
A	Anode (Lose e^-, oxidation)	Positive (Connected to +)	$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ OH^- is a stronger reducing agent than SO_4^{2-} OH^- ions are preferentially discharged to give O_2 Observation: Colorless gas evolved
B	Cathode (Gain e^-, reduction)	Negative	$Cu^{2+} + 2e^- \rightarrow Cu$ Cu^{2+} is a stronger oxidizing agent than H^+. Cu^{2+} ions are preferentially discharged to give Cu. Observation: Size of electrode increases.
C	Anode (Lose e^-, oxidation)	Positive	$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ OH^- is a stronger reducing agent than Cl^- OH^- ions are preferentially discharged to give O_2 Observation: Colorless gas evolved
D	Cathode (Gain e^-, reduction)	Negative (Connected to -)	$2H^+ + 2e^- \rightarrow H_2$ H^+ is a stronger oxidizing agent than Na^+ H^+ ions are preferentially discharged to give H_2 Observation: Colorless gas evolved

Cell x electrolysis



Reactions in beaker B (A chemical cell)			
	Half-equation + Observation	**Cathode / Anode**	**Positive / Negative**
Y	$2H^+ + 2e^- \rightarrow H_2$ (Ag is less reactive, electrons flow from Mg to Ag) Observation: Colorless gas bubbles produced	Cathode (H^+ is reduced)	Positive (e^- flows from – to +)
Z	$Mg \rightarrow Mg^{2+} + 2e^-$ (Mg is more reactive, it loses electrons more readily) Observation: Size of Mg electrode decreases	Anode (Mg is oxidized)	Negative (e^- flows from – to +)

Reactions in beaker A (An electrolytic cell)			
	Cathode / Anode	**Positive / Negative**	**Half-equation + Observation**
W	Cathode (Gain e^-, reduction)	Negative (Connected to -)	$Cu^{2+} + 2e^- \rightarrow Cu$ Cu^{2+} is a stronger oxidizing agent than H^+. Cu^{2+} ions are preferentially discharged to give Cu. Observation: Size of electrode increases.
X	Anode (Lose e^-, oxidation)	Positive (Connected to +)	$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ OH^- is a stronger reducing agent than SO_4^{2-} OH^- ions are preferentially discharged to give O_2 Observation: Colorless gas evolved

Chemical cell (chemical energy to electrical energy)		**Electrolysis (Electrical energy to chemical energy)**	
			
Negative	**Positive**	**Negative**	**Positive**
Losing electrons	Gaining electrons	Gaining electrons	Losing electrons
[$\rightarrow e^-$]	[$e^- \rightarrow$]	[$e^- \rightarrow$]	[$\rightarrow e^-$]
Oxidation	Reduction	Reduction	Oxidation
Anode	Cathode	Cathode	Anode
Reducing agent	Oxidizing agent	Oxidizing agent	Reducing agent
O.N. increases	O.N. decreases	O.N. decreases	O.N. increases