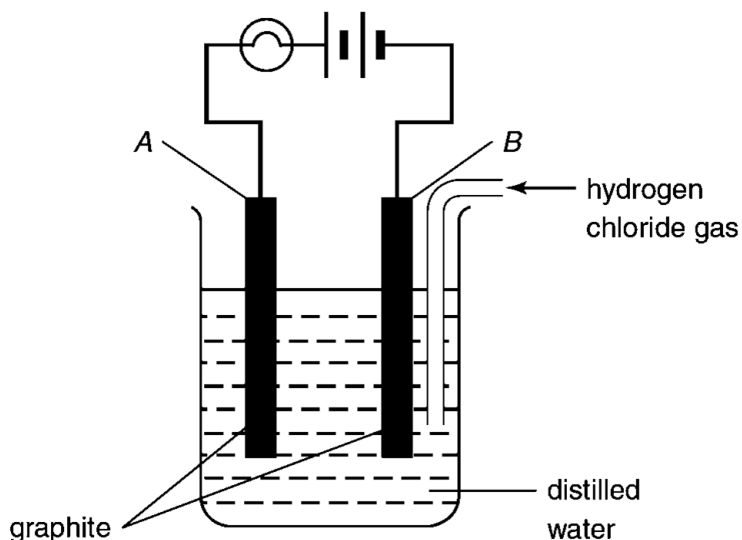
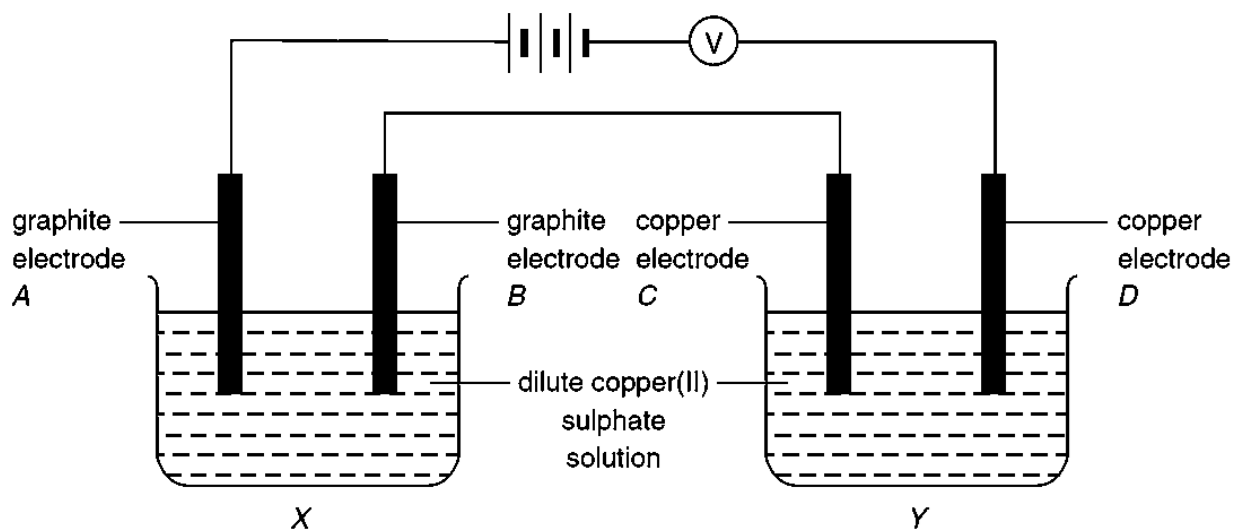


Quiz (More Complicated Questions)

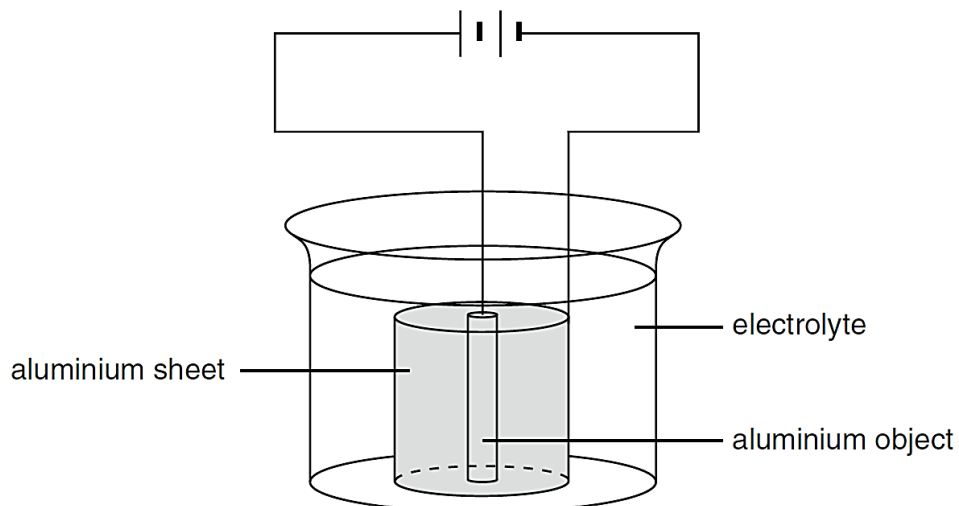
1. The diagram below shows an electrolytic cell using graphite electrodes.



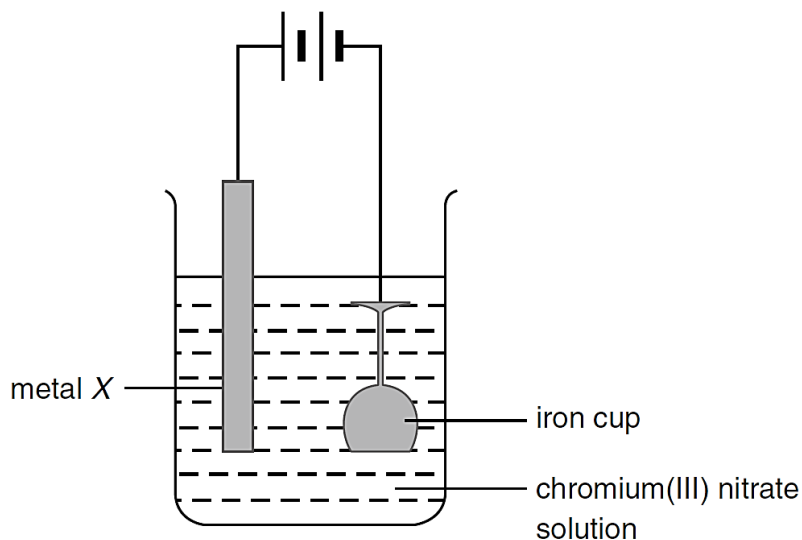
- Explain why the bulb does not light up when distilled water is used as electrolyte.
 - Explain why the bulb lights up when hydrogen chloride gas is passed into distilled water.
 - During electrolysis, colourless gas bubbles are produced at both electrodes A and B. What is the gas produced at each electrode?
 - Write half equations for the reactions taking place at electrodes A and B respectively.
 - Explain why the rates of formation of gas at the two electrodes are different.
2. The following set-up shows the electrolysis of dilute copper(II) sulphate solution with different electrodes in two electrolytic cells, X and Y, connected in series.



4. The set-up below is used for the anodization of aluminium.

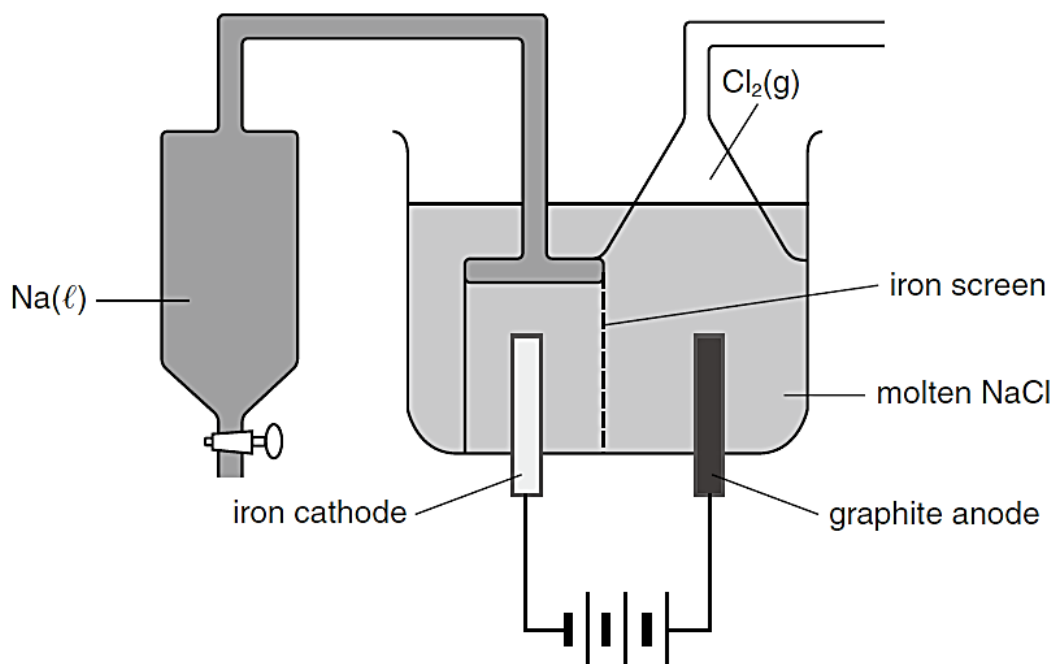


- What is the electrolyte used in the anodization process?
 - What is the polarity of the aluminium object to be anodized?
 - Identify the anode and the cathode in the anodization set-up.
 - Explain how the coated aluminium oxide protects the aluminium underneath.
5. A student carried out a chromium-plating experiment in the laboratory using the set-up shown below.



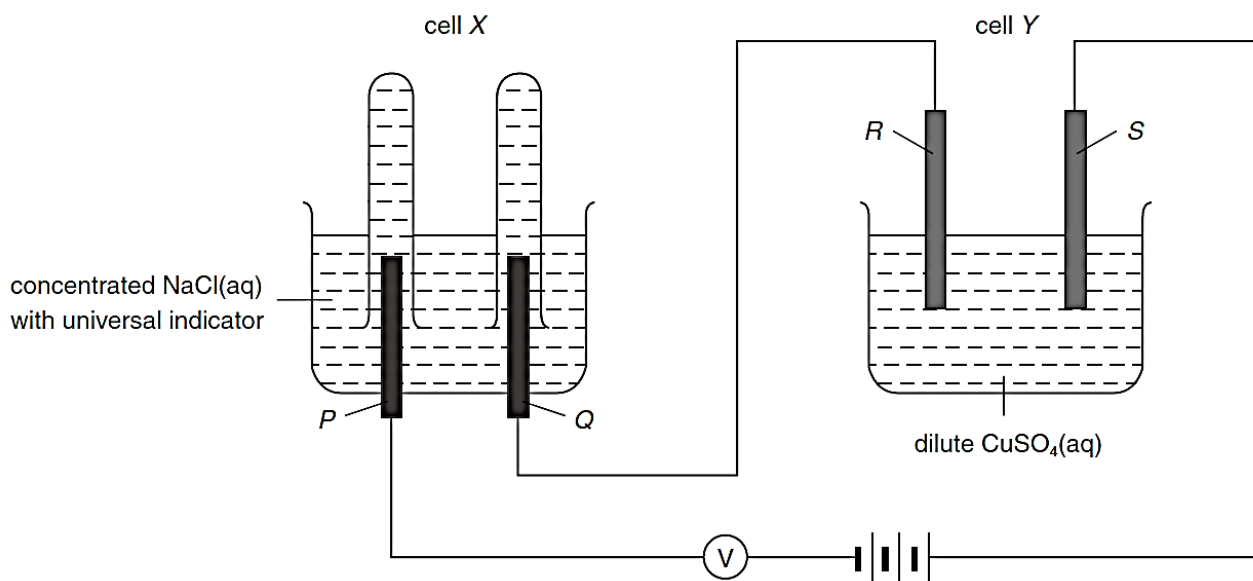
- What is metal X in the set-up?
- Write a half equation for the reaction taking place at
 - metal X, and
 - the iron cup.
- What would be observed at metal X?
- Suggest ONE advantage of electroplating an iron cup with chromium.
- Suggest how the student should treat the electrolyte before discharge.

6. The following diagram shows an electrolytic cell called Downs cell, in which molten sodium chloride is electrolysed to give sodium metal and chlorine gas.



- (a) Explain why the sodium chloride used in the Downs cell is in molten state but not in solid state.
- (b) Write the half equations for the reactions taking place at the iron cathode and the graphite anode respectively.
- (c) Write an overall equation for the electrolytic process.
- (d) Sodium and chlorine can also be obtained by the electrolysis of concentrated sodium chloride solution using mercury cathode and graphite anode.
- Write an overall equation for the electrolytic process.
 - State ONE advantage of this process over Downs cell.

7. An electric circuit is set up as shown in the following diagram. Electrodes *P* and *Q* are made of graphite while electrodes *R* and *S* are made of copper.



- (a) (i) What are the gases collected in the tubes above electrodes *P* and *Q* respectively?
- (ii) Write the half equations for the reactions taking place at electrodes *P* and *Q*.
- (iii) Predict the volume ratio of the two gases collected.
- (b) State the colour changes of the universal indicator around electrodes *P* and *Q* during electrolysis.
- (c) (i) In cell *Y*, identify the positive electrode and the negative electrode.
- (ii) State the expected observable changes at the two electrodes in cell *Y*.
- (d) The copper electrodes *R* and *S* are replaced by platinum electrodes.
- (i) What would be observed at platinum electrodes *R* and *S* respectively?
- (ii) State and explain the change of the electrolyte solution after electrolysis.

8. In an electroplating plant, the liquid waste often contains chromate ions. To remove chromate ions, sodium sulphite solution and sulphuric acid are added. Sulphite ions reduce chromate ions to chromium(III) ions. The chromium(III) ions are then precipitated out and the precipitate is filtered off.
- (a) Apart from chromate ions, name TWO other pollutants that can be found in the liquid waste.
- (b) What is the effect of chromate ions on the environment if they are discharged directly?
- (c) Write an equation for the reaction between sulphite ions and chromate ions.
- (d) What is the change in oxidation number of chromium when chromate ions are reduced to chromium(III) ions?
- (e) (i) Suggest a chemical that can be used to precipitate out chromium(III) ions.
- (ii) Write an ionic equation for the reaction involved.

Suggested Answer

1. (a) Pure water ionizes only very slightly to give mobile ions for conducting electricity. / It is a poor conductor of electricity.
 - (b) Hydrogen chloride gas dissolves in water and ionizes to give mobile $\text{H}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$ ions for conducting electricity.
 - (c) At electrode A: oxygen;
At electrode B: hydrogen
 - (d) At electrode A: $4\text{OH}^-(\text{aq}) \longrightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^-$
At electrode B: $2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{H}_2(\text{g})$
 - (e) The overall equation for the reaction is $2\text{H}_2\text{O}(\text{l}) \longrightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$.
The mole ratio of the gases formed at the anode and at the cathode is 1:2.
Thus, the rates of formation of the gases are different.
2. (a) (i) Anode: electrode A;
Cathode: electrode B
 - (ii) At electrode A: $4\text{OH}^-(\text{aq}) \longrightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^-$
At electrode B: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cu}(\text{s})$
 - (iii) Electrode A: colourless gas bubbles form.
Electrode B: reddish brown solid deposits on the surface of electrode.
 - (b) (i) Anode: electrode C;
Cathode: electrode D
 - (ii) At electrode C: $\text{Cu}(\text{s}) \longrightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$
At electrode D: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cu}(\text{s})$
 - (iii) At electrode C: the copper electrode becomes smaller.
At electrode D: the copper electrode gets bigger.
 - (c) When graphite anode is used, $\text{OH}^-(\text{aq})$ ions are preferentially discharged.
When copper anode is used, $\text{Cu}(\text{s})$ is preferentially discharged as $\text{Cu}(\text{s})$ is a stronger reducing agent than $\text{OH}^-(\text{aq})$.
 - (d) The solution in the left electrolytic cell becomes acidic / sulphuric acid. This is because as electrolysis goes on, $\text{Cu}^{2+}(\text{aq})$ and $\text{OH}^-(\text{aq})$ are consumed and $\text{H}^+(\text{aq})$ and $\text{SO}_4^{2-}(\text{aq})$ ions remain.

There is no change in the copper(II) sulphate solution in the right electrolytic cell. For every $\text{Cu}^{2+}(\text{aq})$ ion discharged at the cathode, one $\text{Cu}^{2+}(\text{aq})$ ion forms at the anode. Therefore, the concentration of the copper(II) sulphate solution does not change.

3. (a) (i) The solution gradually turns red.
OH⁻(aq) ions are preferentially discharged and oxidized to O₂(g).
The concentration of H⁺(aq) ions is higher than that of OH⁻(aq) ions around carbon rod X as OH⁻(aq) ions are consumed.
- (ii) The solution gradually turns blue.
H⁺(aq) ions are preferentially discharged and reduced to H₂(g).
The concentration of OH⁻(aq) ions is higher than that of H⁺(aq) ions around carbon rod Y as H⁺(aq) ions are consumed.
- (b) H⁺(aq) ions OH⁻(aq) ions are discharged during electrolysis.
Water is decomposed to replace the ions consumed.
As a result, the amount of water in the electrolyte solution decreases.
- (c) No. This is because the same kind of ions, OH⁻(aq) ions, will be discharged at carbon rod X even if the electrolyte is replaced by very dilute NaBr(aq).
4. (a) Dilute sulphuric acid
- (b) Positive electrode
- (c) Anode: the aluminium object;
Cathode: the aluminium sheet
- (d) The coated aluminium oxide layer is impermeable to air and water.
Thus, the aluminium underneath can be protected from further corrosion.
5. (a) Chromium
- (b) (i) $\text{Cr(s)} \longrightarrow \text{Cr}^{3+}(\text{aq}) + 3\text{e}^{-}$
- (ii) $\text{Cr}^{3+}(\text{aq}) + 3\text{e}^{-} \longrightarrow \text{Cr(s)}$
- (c) Metal X becomes smaller.
- (d) The cup becomes more corrosion resistant. / The appearance of the cup is improved.
- (e) Add sodium hydroxide solution to the electrolyte. Chromium(III) ions would be precipitated out as chromium(III) hydroxide. The precipitate is then filtered off before discharge.

6. (a) The ions in solid sodium chloride are not mobile. Thus, it does not conduct electricity.
But in molten state, the ions in sodium chloride become mobile.
- (b) At graphite anode: $2\text{Cl}^{-}(\text{l}) \longrightarrow \text{Cl}_2(\text{g}) + 2\text{e}^{-}$
At iron cathode: $\text{Na}^{+}(\text{l}) + \text{e}^{-} \longrightarrow \text{Na}(\text{l})$
- (c) $2\text{Na}^{+}(\text{l}) + 2\text{Cl}^{-}(\text{l}) \longrightarrow 2\text{Na}(\text{l}) + \text{Cl}_2(\text{g})$
- (d) (i) $2\text{Hg}(\text{l}) + 2\text{Na}^{+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq}) \longrightarrow 2\text{Na/Hg}(\text{l}) + \text{Cl}_2(\text{g})$
- (ii) In Downs cell, a large amount of heat is needed to melt the solid sodium chloride. Electrolysis of concentrated sodium chloride solution saves much energy.
7. (a) (i) At electrode P: hydrogen gas
At electrode Q: chlorine gas
- (ii) At electrode P: $2\text{H}^{+}(\text{aq}) + 2\text{e}^{-} \longrightarrow \text{H}_2(\text{g})$
At electrode Q: $2\text{Cl}^{-}(\text{aq}) \longrightarrow \text{Cl}_2(\text{g}) + 2\text{e}^{-}$
- (iii) The volume ratio of the H_2 to Cl_2 collected is 1 : 1.
- (b) The universal indicator turns blue around electrode P during electrolysis.
The universal indicator turns red and then colourless around electrode Q during electrolysis.
- (c) (i) Positive electrode: electrode S
Negative electrode: electrode R
- (ii) Electrode R gets bigger while electrode S becomes smaller.
- (d) (i) Reddish brown solid deposits on the platinum electrode R.
Colourless gas bubbles evolve at platinum electrode S.
- (ii) As electrolysis continues, $\text{Cu}^{2+}(\text{aq})$ and $\text{OH}^{-}(\text{aq})$ ions are consumed and $\text{H}^{+}(\text{aq})$ and $\text{SO}_4^{2-}(\text{aq})$ ions remain. Hence, the electrolyte solution eventually becomes sulphuric acid.

8. (a) Acids / alkalis / compounds of other heavy metals / cyanides (Any TWO)
- (b) Chromate ions are highly toxic. They may kill the aquatic life if the effluents containing chromate ions are discharged directly.
- (c) $3\text{SO}_3^{2-}(\text{aq}) + 2\text{CrO}_4^{2-}(\text{aq}) + 10\text{H}^+(\text{aq}) \longrightarrow 3\text{SO}_4^{2-}(\text{aq}) + 2\text{Cr}^{3+}(\text{aq}) + 5\text{H}_2\text{O}(\text{l})$
- (d) The oxidation number of Cr changes from +6 to +3.
- (e) (i) Sodium hydroxide solution / sodium carbonate solution (Any ONE)
- (ii) $\text{Cr}^{3+}(\text{aq}) + 3\text{OH}^-(\text{aq}) \longrightarrow \text{Cr}(\text{OH})_3(\text{s})$ (if sodium hydroxide solution is used)
 $2\text{Cr}^{3+}(\text{aq}) + 3\text{CO}_3^{2-}(\text{aq}) \longrightarrow \text{Cr}_2(\text{CO}_3)_3(\text{s})$ (if sodium carbonate solution is used)