REDOX REACTIONS, IONIC EQUATIONS AND ELECTROLYSIS DETAILED NOTES AND SAMPLE QUESTIONS

(I regret for any mistake if noted)

S4 TERM ONE TOPIC

NEW LOWER SECONDARY CURRICULUM

(CHEMISTRY)

BY

TR. KISULE JOSEPH
(0751339538-0786570990)
kisjo19961@gmail.com

DEDICATED TO YOU

The attached questions are almost enough for a student to have a general idea/concept about this region(**content/subtopic**) in chemistry, however, I advise a student to search for more related questions about this content area for **better results.**

CONTENT:

1. ***Redox reactions, ionic equations and electrolysis.***
2. ***Some sample questions on the above topic***
3. ***Try so hard to answer the sample questions and look for more qns.***

Don't say tomorrow, it will be too late for chemistry revision, and yesterday is gone forever, you have got today to revise your chemistry!

"Revise as if tomorrow is not there"

May god bless you

REDOX REACTIONS, IONIC EQUATIONS AND ELECTROLYSIS

These are reactions in which both reductions and oxidations take place simultaneously.

Oxidation

Can be defined as;

i) Addition of oxygen to a substance

ii) Removal of hydrogen from a substance

iii) Loss of electrons from a substance

NB. Oxidizing agents are electron acceptors. Therefore, nonmetals are oxidizing agents.

Examples of oxidizing agents

- Bromine (Br_2)
- Chlorine (Cl_2)
- Concentrated sulphuric acid (H_2SO_4)
- Nitric acid (HNO_3)
- Oxygen (O_2)
- Potassium permanganate ($KMnO_4$)
- Potassium dichromate ($K_2Cr_2O_7$)
- Hydrogen peroxide (H_2O_2)

An oxidizing agent: is a substance which accepts electrons.

An example of an oxidation reaction is conversion of metal atoms to ions by loss of electrons.

Na(s) to Na^+: $Na\ (s)$ - $e^- \longrightarrow$ Na^+ (aq) (Oxidation reaction)

Mg(s) to Mg^{2+}: Mg(s) - $2e^- \longrightarrow$ Mg^{2+} (aq) (Oxidation reaction)

Oxidation as gain of oxygen

Brunning magnesium in oxygen

Requirements

- Magnesium ribbon
- Pair of tongs
- Sandpaper
- Spirit lamp or Bunsen burner

- Watch glass
- Safety goggles.

What to do/steps taken

- Measure about 3cm long piece of magnesium ribbon.
- Using the sandpaper, clean the magnesium ribbon by scrubbing it. (To remove magnesium oxide layer from the surface of the magnesium ribbon which would slow down or prevent burning of magnesium ribbon)
- Hold the piece of the clean magnesium with a pair of tongs.
- Burn the magnesium ribbon in the flame as shown below **(figure 1.2)**, and observe what happens.

Illustration

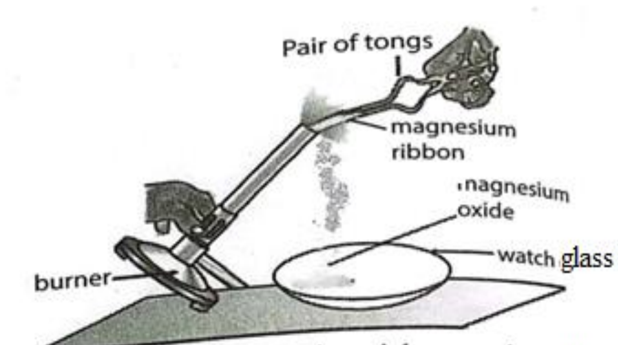


Figure 1.2: Burning Magnesium in *oxygen*

- Collect the product formed on the watch glass.

Observation

- **Magnesium** burns with a dazzling white flame.
- The grey metal solid turned into a white solid

Conclusion

- $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$
- Magnesium was oxidized to magnesium oxide (MgO)
- Oxidation is therefore the gain/addition of oxygen.

Oxidation as loss of hydrogen

Oxidation also occurs when a substance loses hydrogen.

Investigating the process of passing dry ammonia gas over heated copper (II) oxide.

Requirements

- Combustion tube
- Copper (II) oxide
- Source of heat
- Delivery tubes

- Dry ammonia gas
- Stands

Procedure/steps taken/what to do.

- The apparatus is set up as shown in figure 1.3
- Heat copper (II) oxide and pass dry ammonia gas over it.

Illustration

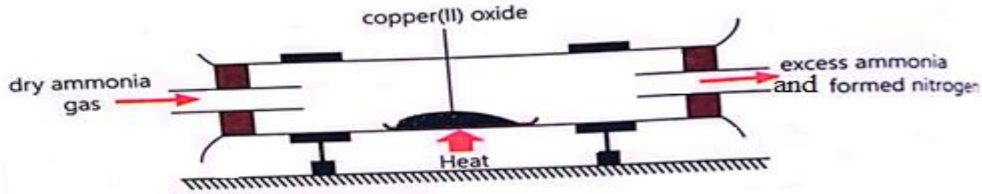


Figure 1.3: Passing dry ammonia gas over heated copper(II) oxide

Observation

- Black colour of copper (II) oxide turned to brown.
- $3CuO(s) + 2NH_3(g) \longrightarrow 3Cu(s) + N_2(g)\ 3H_2O(l)$

Conclusion

- Ammonia was oxidized. This is because it lost hydrogen to form nitrogen gas.
- Copper (II) oxide was reduced. This is because it lost oxygen to form copper metal.

Oxidation as loss of electrons

Consider the reaction between magnesium and chlorine

Magnesium loses its outer most two electrons/its valency to be gained by chlorine.

magnesium + chlorine → magnesium chloride

reduction

$$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$$

oxidation

Electrons have been transferred during this reaction. The following half equations represent the transfer of electrons.

Magnesium atoms loses electrons to form magnesium ions:

$$Mg(s) \longrightarrow Mg^{2+} + 2e$$

The electrons lost are gained by chlorine to form chloride ions:

$$Cl_2(g) + 2e^- \longrightarrow 2Cl^-$$

Discussion questions

1. What has been oxidised in the reaction between magnesium and chlorine?
2. What happens when copper is placed into silver nitrate solution?

Answers

1. **Magnesium has been oxidized because the magnesium atom lost two electrons to form magnesium ion (Mg^{2+})**
2. **The brown solid copper dissolves, and turns the colourless (silver nitrate solution) into a blue solution of copper (II) nitrate.** ***The experiment is as seen below.***

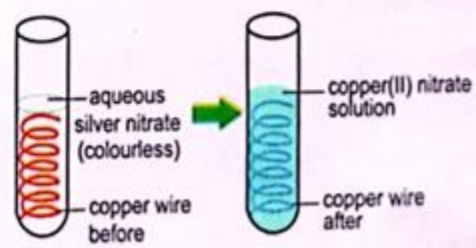
REACTION BETWEEN COPPER WIRE/METAL AND A SOLUTION OF SILVER NITRATE

What you need:

- silver nitrate solution
- strip of copper metal
- test tubes

What to do

1. Coil the strip of copper metal.
2. Place the coil of copper into a test tube containing silver nitrate solution as shown in *Figure. 1.6.*
3. Let the set-up stand for 3-4 minutes, observe and record what takes place.



Reaction between copper metal and silver nitrate solution

Observation and analysis

- The aqueous solution of silver nitrate turned to a blue solution.
- The copper wire reduces in size.
- $Cu(s) + 2AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

Conclusion

- In the reaction, each copper atom loses two electrons and becomes a copper (II) ion in aqueous solution.
- $Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$
- Copper wire loses electrons as silver ion gains electrons.

Reduction

Can be defined as;

i) Removal/loss of oxygen from a substance
ii) The addition/gain of hydrogen to a substance
iii) Gain of electrons by a substance

NB. Reducing agents are electron donors; therefore, all metals are reducing agents.

Examples of reducing agents

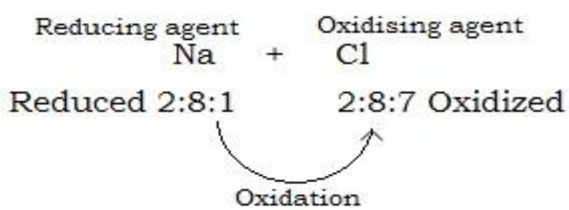
- Carbon (C)
- Carbon monoxide (CO)
- Hydrogen (H_2)
- Hydrogen sulphide (H_2S)
- Metals
- Potassium (KI)
- Sulphur dioxide (SO_2)
- Ammonia (NH_3)

A reducing agent: is a substance which donates electrons.

An example of reduction reaction is conversion of non-metal atoms to ions by gain of electrons.

Example. $Cl + e \longrightarrow Cl^-$

$Na - e \longrightarrow Na^+$



$Cl + e^- \longrightarrow Cl^-$ (Reduction)

Electron acceptor

$Na - e \longrightarrow Na^+$ (Oxidation)

Electron donor

Reduction as loss of oxygen

When a mixture of zinc powder and copper (II) oxide is heated, the following reaction occurs.

Word equation

Zinc + copper (II) oxide $\longrightarrow$ Zinc oxide + copper

Symbol equation

$Zn(s) + CuO(s) \longrightarrow ZnO(s) + Cu(s)$

In this reaction, copper (II) oxide has lost oxygen, so it is reduced to copper metal.

Reduced

$Zn(s) + CuO(s) \longrightarrow ZnO(s) + Cu(s)$

Oxidised

Reduction as gain of hydrogen

When a mixture of hydrogen and chlorine gases is exposed to sunlight, it explodes and produces white fumes of hydrogen chloride.

Word equation

Hydrogen + chlorine ⟶ hydrogen chloride

Symbol equation

Reduction as gain of electrons

In terms of electrons, reduction is gain of electrons by a substance. When hydrogen sulph gas is passed through iron(III) chloride solution, a green solution of iron(II) chloride an light yellow precipitate of sulphur are produced.

Overall equation:

$$2FeCl_3(aq) + H_2S\ (g) \rightarrow 2FeCl_2(aq) + 2HCl\ (aq) + S(s)$$

In this reaction, the half equation for iron is:

$$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$$

Observe that iron has gained an electron and so has been reduced from iron (iii) to iron

Discussion Questions

Identify the species which have been reduced in the following reactions

- a $2KI + H_2O_2 \rightarrow I_2 + 2KOH$
- b $Fe + S \rightarrow FeS$
- c $2HgCl_2 + SnCl_2 \rightarrow Hg_2Cl_2 + SnCl_4$
- d $H_2S + Br_2 \rightarrow 2HBr + S$
- e $H_2S + Cl_2 \rightarrow 2HCl + S$

Suggested responses

a) H_2O_2
b) S
c) $HgCl_2$
d) Br_2
e) Cl_2

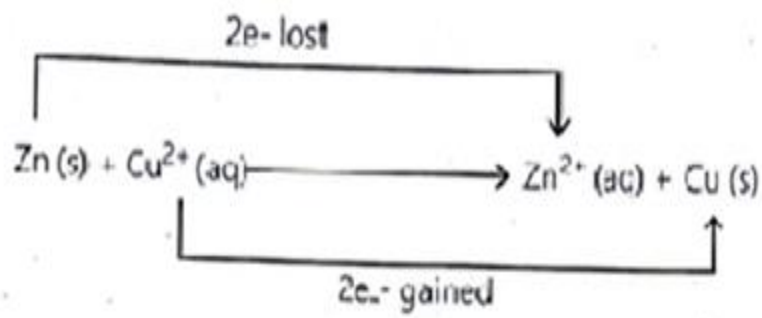
Reduced

$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

Oxidised

Chlorine is reduced, as it has gained hydrogen

Sample questions



Tasks

1. Identify the oxidising and reducing agents in the above reaction.
2. Explain the reasons for your answers in (1).
3. What has happened to the oxidation numbers of zinc and copper in the reaction?
4. Write half ionic equations to illustrate your answers.

Responses

1.
   - Reducing agent: Zn
   - Oxidising agent: Cu^{2+}
2. Zn is a reducing agent because it donates electrons to copper(II) ions. Cu^{2+} ion is an oxidising agent because it accepts electrons from zinc
3. The oxidation number of zinc has increased while that of copper(II) ions has decreased
4.
   - $Zn(s) \rightarrow Zn^{2+}(aq) + 2e$
   - $Cu^{2+}(aq) + 2e \rightarrow Cu(s)$

CTF TASKS

1. Define the following term "reducing agent".
2. In each of the following reactions, identify the reducing and oxidising agents.

a) $Mg (s) + Br_2 (l) \longrightarrow MgBr_2 (s)$
b) $2Fe3+ (aq) + Zn (s) \longrightarrow 2Fe2+ (aq) + Zn2+ (aq)$
c) $2H+ (aq) + Zn (s) \longrightarrow H2 (g) + Zn2+ (aq)$
d) $3CuO (s) + 2NH_3 (g) \longrightarrow 3Cu (s) + 3H_2O (l) + N_2 (g)$
e) $Fe_2O_3 (s) + 3CO (g) \longrightarrow 2Fe (s) + 3CO_2 (g)$
f) $MnO_4- (aq) + 8H+ (aq) + 5Fe2+ (aq) \longrightarrow Mn2+ (aq) + 5Fe3+ (aq) + 4H_2O (l)$

CTF RESPONSES

1. Define the term "reducing agent"
   - Is a substance that donates/loses electrons to an electron recipient.

In each of the following equations, identify the reducing and oxidising agent or species.

a) $Mg (s) + Br_2 (l) \rightarrow MgBr_2(s)$
- Reducing agent: Mg • Oxidising agent: Br_2

b) $2Fe3+ (aq) + Zn (s) \rightarrow 2Fe2+(aq) + Zn2+ (aq)$
- Reducing agent: Zn • Oxidising agent: Fe3+

c) $2H+ (aq) + Zn (s) \rightarrow H2+(g) + Zn2+ (aq)$
- Reducing agent: Zn • Oxidising agent: H+

d) $3CuO(s) + 2NH_3 (g) \rightarrow N_2 (g) + 3H_2O (l) + 3 Cu (s)$
- Reducing agent: NH_3 •Oxidising agent: CuO

e) $Fe_2O_3 (s) + 3CO (g) \rightarrow 2Fe (s) + 3CO_2 (g)$
- Reducing agent: CO • Oxidising agent: Fe_2O_3

f) $MnO_4- (aq) + 8H+ (aq) + 5Fe2+ (aq) \rightarrow Mn2+ (aq) + 5Fe3+ (aq) + 3H_2O (l)$
- Reducing agent: Fe2+ • Oxidising agent: MnO_4-

Importance of oxidation and reduction (redox) in everyday life

- Extraction of some metals like iron and zinc involves reduction of their ores.
- Helps in the breakdown of food by cells to produce energy, this reaction is an oxidation.

- Electrochemical cells like car batteries, dry cells use redox reactions to produce electrical energy.

Some metals and the ores from which they are extracted.

Element	Ore	Importance
Iron	Magnetite, haematite	Catalyst in the Haber process, used to make jewellery.
Zinc	Zinc blende (zinc sulphide)	Chief ore for extracting zinc
Aluminium	Bauxite	Main raw material, used in the manufacture of aluminium chemicals.
Copper	Copper pyrites	Main raw material

Questions and solutions

1. *Identify three redox reactions and identify the oxidized or reduced species in the three different reactions.*

Reaction	**Oxidized species**	**Reduced species**
a) Copper (II) ions and zinc metal	Zinc metal	Copper (II) ions
b) Chlorine and iron (II) ions	Iron (II) ions	Chlorine
c) Hydrogen peroxide and manganate (VII) ions in acidic medium	Hydrogen peroxide	Manganate (VII) ions

2. **Write the ionic equations for each of the reactions, clearly showing the electron transfer.**

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

Oxidation: $Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

Reduction: $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

b) $Cl_2(g) + 2Fe^{2+}(aq) \longrightarrow 2Cl^-(aq) + 2Fe^{3+}(aq)$

Oxidation: $2Fe^{2+}(aq) \longrightarrow 2Fe^{3+}(aq) + 2e^-$

Reduction: $Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$

c) $2MnO_4^-(aq) + 6H^+(aq) + 5H_2O_2(l) \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5O_2(g)$

Oxidation: $5H_2O_2(l) \longrightarrow 5O_2(g) + 10H^+(aq) + 10e^-$

Reduction: $2MnO_4^-(aq) + 16H^+(aq) + 10e^- \longrightarrow 2Mn^{2+}(aq) + 8H_2O(l)$

Oxidation numbers

Oxidation numbers are combining powers of substances with oxygen. The state of oxidation in an element in a given compound is indicated by its oxidation number

Simple rules when calculating oxidation numbers

1. All elements in a free state [uncombined state] have an oxidation number of zero.

2. In the case of simple ions, the element has an oxidation number with the same size and sign of the charge on the ion. For example, Cu^{2+} has an oxidation number of +2 and S^{2-} has an oxidation number of -2.

3. The sum of all oxidation numbers of the elements in the compound is zero.

Example

Calculate or determine the oxidation numbers of stated elements in the following i) S in SO_2 ii) S in SO_3 iii) H in H_2O

iv) Mn in $MnO_4{}^-$

v) N in $NO_3{}^-$

vi) Mn in MnO_2

Solutions

i) Sulphur in Sulphur dioxide

$SO_2 = 0$

$S + -2X2 = 0$

$S\_4 = 0$

S = +4

ii) Sulphur in Sulphur trioxide

SO3=0

S + (-2X3) = 0

S – 6 = 0

S = +6

iii) Hydrogen in water

2H + (-2) =0

2H – 2 = 0

2H = +2

H = +1

iv) Manganese in Permanganate ion

M + (-2X4) = -1

$M - 8 = -1$

$M = -1 + 8$

$M = +7$

v) Nitrogen in Nitrate

$N + (-2 \times 3) = -1$

$N - 6 = -1$

$N = -1 + 6$

$N = +5$

vi) Manganese in Manganese (IV) Oxide

$M + (-2 \times 2) = 0$

$M - 4 = 0$

$M = +4$

Exercise

Calculate or determine the oxidation numbers of stated elements in the following;

i) Sulphur in Sulphite ion

ii) Hydrogen in Hydrochloric acid

iv) Phosphorous in phosphate

v) Carbon in Carbon dioxide

vi) Carbon in carbon monoxide

vii) Sulphur in Sulphur dioxide vii) Oxygen in Hydroxyl ion

viii) Chromium in Dichromate ion

ix) Copper in Copper (II) oxide

x) Manganese in Permanganate (VII) ions

Examples

State what is taking place and write half reaction equation in each case.

NB. Oxidation is an increase in oxidation number and reduction is a decrease in the oxidation number.

i) The conversion of hydrogen ions (H+) to hydrogen molecules (H2)

$2H^+(aq) + 2e \rightarrow H_2(g)$

This is a reduction because hydrogen ions have gained electrons

ii) Conversion of Iron (II) ions (Fe^{2+}) to Iron (III) ions (Fe^{3+})

$Fe^{2+}(aq) - e \rightarrow Fe^{3+}(aq)$

This is oxidation because iron (II) ions have lost electrons

iii) Conversion of Copper (II) ions (Cu^{2+}) to Copper, Cu

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

This is a reduction because copper (II) ions have gained electrons.

iv) Conversion of Zinc-to-Zinc ions, Zn^{2+}

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

This is oxidation because zinc has lost electrons

Other redox reactions

1. Reactions of metals with acids

Example

$Zn(s) + 2HCl\ (aq) \rightarrow ZnCl_2\ (aq) + H_2\ (g)$

$Zn(s) + 2H^+, \cancel{2Cl^-} \rightarrow Zn^{2+}, \cancel{2Cl^-} + H_2(g)$

Ionic equation

$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

$Zn(s) - 2e \rightarrow Zn^{2+}(aq)$ oxidation half reaction……….(i)

$2H^+(aq) + 2e \rightarrow H_2(g)$ reduction half reaction……….(ii)

Redox overall equations: $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq)\ H_2(aq)$ ie (i) + (ii)

2. Displacement reactions as redox reactions

(a) Metals

A displacement reaction is one in which one element takes the place of another element in a chemical reaction. Example, metals high up in the reactivity series [more reactive metals] tend to display those below them from their solutions [compounds].

In this case, the more reactive metals act as reducing agents [electron donors] and their reducing power decreases as reactivity decreases.

Example

1. When zinc dust is added to a solution of copper [II] sulphate, the copper [II] ions are displaced from a solution to form a brown precipitate[solid] and the blue colour of the solution gradually fades

$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$
(Blue solution) (Colourless) (Brown)

$Zn(s) + Cu^{2+}, SO_4^{2-}(aq) \longrightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$

Overall equation

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

i) $Zn(s) - 2e \longrightarrow Zn^{2+}(aq)$ (oxidation half reaction) ………………(i) + (ii) ii)

$Cu^{2+}(aq) + 2e \longrightarrow Cu(s)$ (Reduction half reaction)………………(ii)

Overall redox equation: $Zn(s) + Cu^{2+} \longrightarrow Zn^{2+}(aq) + Cu(s)$

Example

Show that the following reactions are redox reactions. In each case write half reactions and overall redox reaction and state the observation(s) made if any.

i) Iron in Copper (II) sulphate
ii) Magnesium and Zinc (II) nitrate
iii) Silver and Copper (II) Sulphate
iv) Zinc and Lead (II) nitrate

b) Displacement reactions of halogens

Halogens are displaced from their solutions by other halogens which are higher in the electro chemical series [more reactive halogens].

The order of reactivity of halogens is Fluorine>Chlorine>Bromine>Iodine>Astatine.

For example

When chlorine gas is passed through an aqueous colourless solution of potassium iodide, a brown colouration develops immediately.

The brown colouration is due to iodine liberated as iodide ions are displaced by chlorine from solution example;

$Cl_2(g) + 2KI(aq) \longrightarrow 2KCl(aq) + I_2(l)$

$Cl_2(g) + 2K^+ + 2I^-(aq) \longrightarrow 2K^+(aq) + 2Cl^-(aq) + I_2(l)$

$Cl_2(g) + 2I^-(aq) \longrightarrow 2Cl^-(aq) + I_2(l)$

$Cl_2(g) + 2e \longrightarrow 2Cl^-(aq)$ (Reduction half reaction) – (i)

$2I^-(aq) - 2e \longrightarrow I_2(l)$ (Oxidation half reaction) – (ii)

Overall Redox equation:

$Cl_2(g) + 2I^-(aq) \longrightarrow 2Cl^-(aq) + I_2(aq)$

Chlorine in the above case acts as the oxidizing agent and the oxidizing power of the halogens decreases with their decreasing reactivity.

IONIC EQUATIONS

The ionic equation is used to describe the chemical reaction while also clearly indicating which of the reactants and/or products exist primarily as ions in aqueous solution.

Rules in writing ionic equations

1. Start with a balanced molecular equation.

A molecular equation is one that shows the chemical formulas of all reactants and products but does not expressly indicate their ionic nature.

For example,

$2\ Na_3PO_4(aq) + 3\ CaCl_2(aq) \longrightarrow 6\ NaCl(aq) + Ca_3(PO_4)_2(s)$

2. Break all soluble strong electrolytes (compounds with (aq) beside them) into their ions

- indicate the correct formula and charge of each ion
- indicate the correct number of each ion
- write (aq) after each ion

3. Bring down all compounds with (s), (l), or (g) unchanged

$6\ Na^+(aq) + 2PO_4^{3-}(aq) + 3Ca^{2+}(aq) + 6Cl^-(aq) \longrightarrow 6Na^+(aq) + 6Cl^-(aq) + Ca_3(PO_4)_2(s)$

How did I get this equation banange?

Consider each reactant or product separately:

1 mole of Na_3PO_4 contains 3 moles of Na^+ and 1 mole of PO_4^{3-}. Since the balanced equation shows that two moles of sodium phosphate are involved in the reaction, a total of 6 moles (2 x 3) of Na^+ and 2 moles (2 x 1) of PO_4^{3-} are formed. Notice that the subscript "4" in the formula for the phosphate ion is not used when determining

the number of phosphate ions present. That particular subscript is part of the formula for the phosphate ion itself.

$$2\ Na_3PO_4\ (aq) \longrightarrow 6\ Na^+\ (aq) + 2\ PO_4^{3-}\ (aq)$$

1 mole of $CaCl_2$ contains 1 mole of Ca^{2+} and 2 moles of Cl^-. Remember, the subscript "2" indicates the number of chloride ions. Cl_2 is elemental chlorine. You will never have a diatomic chlorine ION (i.e. Cl_2^- or Cl_2^{2-}) in aqueous solution. Since the balanced equation shows that 3 moles of calcium chloride are involved in the reaction a total of 3 moles (3 x 1) of Ca^{2+} and 6 moles (3 x 2) of Cl^- are formed.

$$3\ CaCl_2\ (aq) \longrightarrow 3\ Ca^{2+}\ (aq) + 6\ Cl^-\ (aq)$$

1 mole of NaCl contains 1 mole of Na^+ and 1 mole of Cl^- . Since the balanced equation shows that 6 moles of NaCl are produced by the reaction, 6 moles (6 x 1) of Na^+ and 6 moles (6 x 1)of Cl^- will be formed.

$$6\ NaCl\ (aq) \longrightarrow 6\ Na^+\ (aq) + 6\ Cl^-\ (aq)$$

Since calcium phosphate is an insoluble solid (indicated by the (s) beside its formula), it will not form appreciable amounts of ions in water. It is brought down unchanged into the complete ionic equation.

4. Cross out the spectator ions that are present

Spectator ions are ions that are present in the reaction mixture but do not participate in it. They "sit around and watch the reaction take place" just like a spectator at a football game watches the players in the game but doesn't play the game himself. You can recognize spectator ions by looking for ions that are present on both sides of the equation. They will always have the same exact formula, charge, and physical state. They will also be present in exactly the same number on both sides of the equation.

$$\cancel{6\ Na^+(aq)} + 2PO_4^{3-}(aq) + 3Ca^{2+}(aq) + \cancel{6Cl^-(aq} \longrightarrow \cancel{6\ Na^+(aq)} + \cancel{6\ Cl^-(aq)} + Ca_3(PO_4)_2(s)$$

5. Write the "leftovers" as the net ionic equation

$$2\ PO_4^{3-}\ (aq) + 3\ Ca^{2+}\ (aq) \longrightarrow Ca_3(PO_4)_2\ (s)$$

The ions that appear in the final net ionic equations are called participants as they are the ones that are actively involved in the reaction.

Example 1:

Write the ionic equation for the word equation

Sodium chloride(aq) + silver nitrate(aq) → silver chloride(s) + sodium nitrate(aq)

Solution:

Step 1: Write the equation and balance it.

$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Step 2: Split the ions. (Only compounds that are aqueous are split into ions.)

$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-$ **Step**

3: Cancel out spectator ions.

$$\cancel{Na^+}(aq) + Cl^-(aq) + Ag^+(aq) + \cancel{NO_3^-(aq)} \rightarrow AgCl(s) + \cancel{Na^+(aq)} + \cancel{NO_3^-}$$

Step 4: Write a balanced ionic equation

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Example 2:

Write the ionic equation for the word equation

Sodium(s) + hydrochloric acid(aq) -> sodium chloride(aq) + hydrogen(g) Solution:

Step 1: Write the equation and balance it.

$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$

Step 2: Split the ions. (Only compounds that are aqueous are split into ions.)

$2Na(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow 2Na^+(aq) + 2Cl^-(aq) + H_2(g)$

Step 3: Cancel out spectator ions.

$$2Na(s) + 2H^+(aq) + \cancel{2Cl^-(aq)} \rightarrow 2Na^+(aq) + \cancel{2Cl^-(aq)} + H_2(g)$$

Step 4: Write a balanced ionic equation

$2Na(s) + 2H^+(aq) \rightarrow 2Na^+(aq) + H_2(g)$

Exercise

Write the complete ionic and net ionic equations for each of the following reactions:

1. $3\ (NH_4)_2CO_3\ (aq) + 2\ Al(NO_3)_3\ (aq) \longrightarrow 6\ NH_4NO_3\ (aq) + Al_2(CO_3)_3\ (s)$
2. $2\ NaOH\ (aq) + H_2SO_4\ (aq) \longrightarrow Na_2SO_4\ (aq) + 2\ H_2O\ (l)$
3. $Mg\ (s) + 2\ HCl\ (aq) \longrightarrow MgCl_2\ (aq) + H_2\ (g)$
4. $Zn\ (s) + CuSO_4\ (aq) \longrightarrow ZnSO_4\ (aq) + Cu\ (s)$
5. $Na_2CO_3\ (aq) + 2\ HNO_3\ (aq) \longrightarrow 2\ NaNO_3\ (aq) + H_2O\ (l) + CO_2\ (g)$

Electrolysis

Electrolysis is the decomposition of an electrolyte in aqueous solution or molten state by passing an electric current through it.

Conductors and nonconductors

Conductors:

These are substances that allow electricity to pass through them.

In electrolytes, the conducting particles are called ions while in metals the conducting particles are electrons.

Some substances do not conduct electricity in solid state e.g. Solid sodium chloride or gaseous state e.g. hydrogen chloride gas but conduct well in aqueous (solution) or molten form.

This is because in solid, the compounds consist of ions held together by strong forces of attraction but ions separate in molten or solution form and can move freely.

Non- conductors/Insulators:

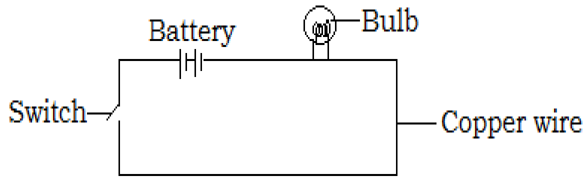
These are substances that do not allow the passage of electricity through them

Investigation of electrical conductivity through solid materials

Examples of such materials include: copper wire, zinc wire, plastic, graphite, rubber.

Procedure

Connect the copper wire to the batteries through the bulb and the switch as shown below.



Results

The bulb produced light on complete connection with copper and zinc.

Conclusion

Copper and zinc wires conduct electricity and they are called **conductors**.

A conductor;

Is therefore a substance in solid form which can conduct electricity.

Examples include;

- all metals and graphite (the only nonmetal that can conduct electricity.

When the above experiment was repeated using rubber and plastics, the bulb did not light which indicate that they do not conduct electricity and are referred to as **insulators** or **nonconductors**.

A nonconductor

is therefore a substance in solid form that does not conduct electricity. **Examples;** all nonmetal except graphite.

N.B

Metals conduct electricity because they have delocalized, free or mobile electrons but nonmetals do not have these delocalized electrons as they are all locked up in bond formation

Electrolytes and non-electrolytes

An electrolyte is an ionic compound which conducts an electric current in aqueous solution or in molten state and is decomposed by it.

Electrolytes are composed of ions. In the solid state, the ions are rigidly held in regular positions and are not able to move freely.

Melting the solid breaks, the forces between the ions and therefore the ions are free to move in a molten electrolyte.

Dissolving a solid in water or any other polar solvent, causes the breakdown of the lattice setting the ions free in aqueous state.

Investigation of electrical conductivity through liquid substances in solution

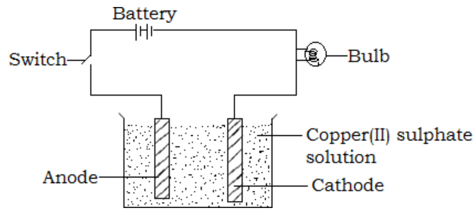
Examples:

Using *ethanol, urea, hydrochloric acid, copper (II) sulphate, ethanoic acid, water, ammonium hydroxide.*

Procedures

1. Put the liquid under investigation in an electrolytic cell.
2. Dip two rods in the liquid which can either be a metal or carbon(graphite) called electrodes

3. Connect the electrodes using a conductor to a bulb via a switch to the source of power (the batteries) as shown below.



4. Close the switch.
5. Repeat the experiment with hydrochloric acid, ethanol, ethanoic acid, water, urea, ammonium, ammonium hydroxide.

Results

When ethanol and urea were used there was no light produced indicating that they do not conduct electricity, they are therefore called **non electrolytes**.

When ammonium hydroxide and ethanoic acid were used, the bulb produced a dim light indicating that they weakly conduct and are therefore **weak electrolytes**.

When copper sulphate solution and hydrochloric acid were used, the bulb produced bright light indicating that they strongly conduct electricity and are **strong electrolytes**.

Types of electrolytes

Strong electrolyte:

This is a compound which is completely ionized in dilute solution and in molten state e.g. salts such as sodium chloride and mineral acids e.g. hydrochloric acid

$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

$NaCl_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$

Weak electrolyte:

This is a compound which is only slightly ionized in dilute solution and in the molten state. They have very few mobile ions and therefore slightly conduct an electric current e.g. water, carbonic acid, ethanoic acid and ammonia solution.

$H_2CO_{3(aq)} \rightleftharpoons 2H^+_{(aq)} + CO_3{}^{2-}{}_{(aq)}$

$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$

$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$

$NH_4OH_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-_{(aq)}$

Non electrolyte:

This is a solution or molten compound which cannot be decomposed by an electric current e.g., sugar, alcohols, benzene and most organic compounds

A non- electrolyte.

This is a solution or a molten compound which does not conduct electricity and therefore cannot be decomposed by an electric current e.g. Paraffin, sugar solution, ethanol etc.

Non- electrolytes exist only in the form of molecules and are incapable of ionization. The molecules have no charge and are therefore not able to carry an electric current.

Electrodes

These are poles of carbon (graphite) or metals where current enters and leaves the electrical device/electrolyte. The types of electrodes include;

Cathode:

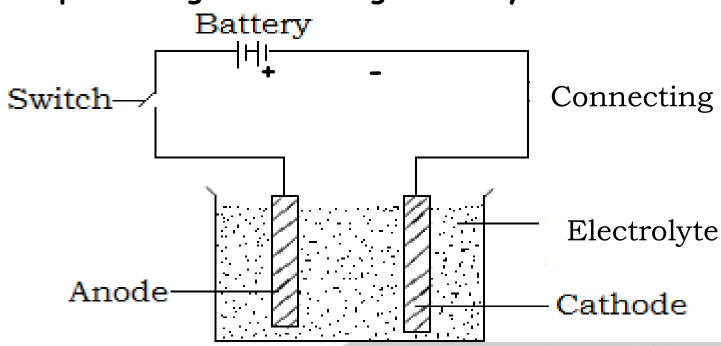
This is a negative electrode at which electrons enter the electrical device/electrolyte or leave the external circuit

Anode:

This is the positive electrode at which the electrons leave the electrical/electrolyte or enter the external circuit.

NB: An electrode must be a good conductor of electricity and should not react with the electrolyte.

A simple arrangement during electrolysis



Ions

An ion is a charged particle. Types of ions include:

Cation:

This is a positively charged ion that will move to the cathode during electrolysis e.g. all metallic ions e.g. Na^+, NH_4^+, H^+, Cu^{2+}, Pb^{2+} etc.

Anion:

This is a negatively charged ion that moves to the anode during electrolysis e.g. all non- metal ions and radicals e.g. Cl^-, SO_4^{2-}, OH^-, NO_3^-, Br^- etc.

Theory of electrolysis (Ionic theory)

This states that electrolytes consist of ions which are positively and negatively charged particles that move to different electrodes during electrolysis.

In ionic compounds, these charges are held together by electrostatic forces but in solution or molten state, these ions are free to move.

The positive ions move to the cathode and the negative to the anode.

What happens to anions at the anode?

When an electric current is applied to the electrolyte, the negatively charged ions called **anions** move to the positively charged electrode called the **anode**.

Once there, they lose electrons to become atoms and are said to be **discharged** i.e.

$2X^{n-}_{(aq\ or\ l)} \rightarrow X_{2(g)} + ne$

What happens to cations at the cathode?

The positively charged ions called **cations** move to the negatively charged electrode called cathode where they gain electrons and become atoms which are then said to be **discharged** i.e.

$M^{n+}_{(aq\ or\ l)} + ne \rightarrow M_{(s)}$

Selective discharge of ions

When two or more ions of similar charges are present under similar conditions in a solution e.g., K^+ and H^+ or $SO_4{}^{2-}$ and OH^-, one is preferentially selected for discharge.

The selective discharge depends on the following factors.

1. Position of the metal or group in the electrochemical (activity) series

If there are two ions with the same charge in the solution, the least reactive ion is discharged first.

For example, in the electrolysis of sodium chloride solution, both Na^+ and H^+(from water) are present and migrate to the cathode.

The H^+ being less reactive than the Na^+ is discharged first.

If $Cu^{2+}$and H^+ ions are both present in solution, both migrate to the cathode but the H^+ being less reactive than the Na^+ is discharged first.

The electrochemical series for ions is given below.

Cations		Anions	
K^+		$NO_3{}^-$	
Ca^{2+}		$SO_4{}^{2-}$	
Na^+		Cl^-	
Mg^{2+}		Br^-	
Al^{3+}		I^-	
Zn^{2+}	decreasing reactivity ↓	OH^-	decreasing reactivity ↓
Fe^{2+}			
Pb^{2+}			
H^+			
Cu^{2+}			
Hg^{2+}			
Ag^+			

2. The nature of electrodes

This factor sometimes influences the choice of ion discharge.

E.g., when a solution of copper (II) sulphate is electrolyzed using copper electrodes, copper ions are discharged at the cathode but neither SO_4^{2-} nor OH^- are discharged at the anode.

Instead, the anode dissolves.

Reaction at the cathode

$Cu^{2+}_{(aq)} + 2e \rightarrow Cu_{(s)}$

Reaction at the anode:

$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e$

When copper (II) sulphate solution is electrolyzed using the copper cathode and carbon anode, the copper is discharged at the cathode and OH^- is discharged at the anode.

Reaction at the cathode:

$Cu^{2+}_{(aq)} + 2e \rightarrow Cu_{(s)}$

Reaction at the anode

$4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$

E.g.,

Electrolysis of a solution of sodium chloride with mercury as a cathode and with platinum as cathode. With platinum, the hydrogen ion is discharged in accordance with the order of the activity series, sodium ion being higher in the series.

The cathode product is hydrogen gas.

If the mercury cathode is used, there is a possibility of discharging sodium ion to form sodium amalgam with mercury.

This requires less energy than the discharge of hydrogen ions to form hydrogen gas and so occurs in preference.

3. Concentration of the ions

Increase of concentration of an ion tends to promote its discharge. For example, in concentrated sodium chloride solution, there are hydroxide ions and chloride ions but the concentration of the chloride ions exceeds that of the hydroxide ions and therefore chloride ions are discharged.

The more concentrated ions are discharged in preference to ones which are less concentrated.

Electrolysis of dilute sulphuric acid

This is commonly called electrolysis of water.

Ions present:

From sulphuric acid are H^+ and SO_4^{2-}

From water are H^+ and OH^-.

Reaction at cathode:

The hydrogen ions migrate to the cathode, gain electrons and become hydrogen gas.

$2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$

Reaction at the anode:

The hydroxide ions and sulphate ions migrate to the anode.

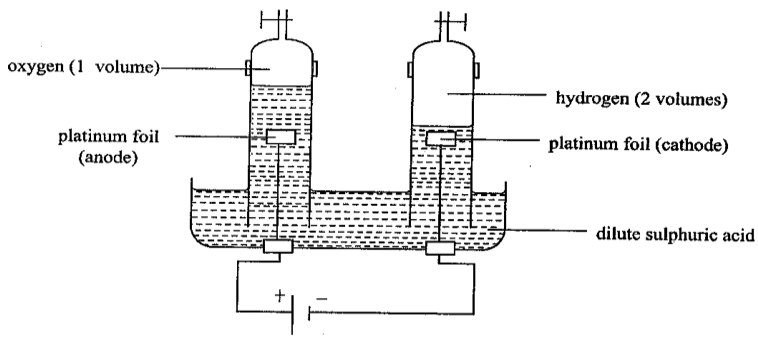
The hydroxide ions being less reactive than sulphate ions are discharged and oxygen gas is formed.

$4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)} + 4e^-$

Overall equation

$4H^+_{(aq)} + 4OH^-_{(aq)} \rightarrow 2H_{2(g)} + O_{2(g)} + 2H_2O_{(l)}$

Electrolysis of dilute sulphuric acid in an apparatus called a voltammeter



Note:

1. PH changes at the anode and cathode:

The acidity at the cathode decreases (pH increases) because the hydrogen ions are discharged as hydrogen gas and therefore the concentration of hydrogen ions in solution decreases. At the anode, the acidity increases (pH decreases).

The discharge of hydroxide ions disturbs the ionic equilibrium of water and therefore more water ionizes to restore it.

$H_2O_{(l)} \rightarrow H^+_{(aq)} + OH^-_{(aq)}$

Therefore, the excess hydrogen ions produced, with incoming sulphate ions, is equivalent to increased concentration of sulphuric acid. This means that the total acidity at anode and cathode together remains constant. This implies that the final change is that water is decomposed to produce hydrogen and oxygen. **That is why it is called electrolysis of water.**

2. Two volumes of hydrogen are produced at the cathode and one volume of oxygen produced at the anode i.e., Hydrogen: Oxygen =2:1

Overall equation

$4H^+_{(aq)} + 4OH^-_{(aq)} \rightarrow 2H_{2(g)} + O_{2(g)} + 2H_2O_{(l)}$

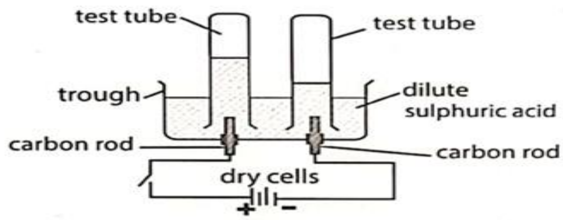
Trial experiment:

Electrolysis of dilute sulphuric acid
Using Carbon Electrodes

What is needed:
- trough
- switch
- carbon electrode with holders
- connecting wires
- dry cells
- burning splint
- test tubes
- glowing splint
- 0.5 M sulphuric acid

What to do

1. Fill the trough with the electrolyte as set up in Figure 1.13 with 0.5 M dilute sulphuric acid, until it is half-full.
2. Fill each test tube will dilute sulphuric acid and then immerse it at the anode and cathode as shown in Figure 1.13.



Electrolysis of dilute sulphuric acid solution

3. Turn on the switch and allow the current to flow for 20 minutes. Observe and record any changes at the anode, cathode, and in the electrolyte. Record the observations
4. Test the gas liberated at the anode with a glowing splint, while the one at the cathode using a burning splint. Record your observations.

Sample questions on electrolysis of sulphuric acid

1. What are the ions present in dilute sulphuric acid?
2. a) What did you observe at the each of the electrodes?
   b) Write ionic equations for the reactions that occurred at each of the electrodes.
3. What happened to the colour of the solution at the end of the experiment?

Answers
(Take note of the symbols! Some are not correct symbols and not written correctly, so do the necessary)

1. H+, SO42- from sulphuric acids , OH- and H+ from water
2. a)
- Anode: Bubbles of a colourless gas that relighted a glowing split.
- Cathode: Bubbles of a colourless gas that burnt with a pop sound

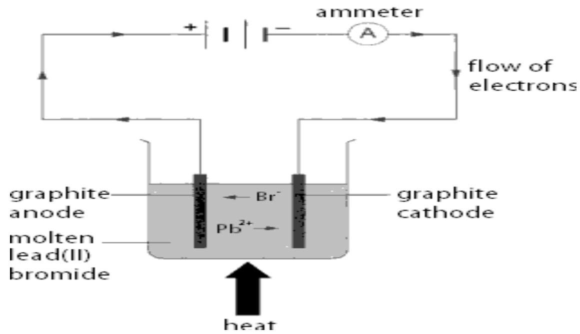
b) Anode:
- 4OH- (aq) → O2 (g) + $2H_2O$ (l) + 4e

Cathode:
- 4H+ (aq) + 4e → 2H2 (g)

3. It remains colourless because H+ and OH-ions are in constant supply and their discharges causes more H_2O molecules to ionise.

Electrolysis of molten lead (II) bromide

The bulb does not give out light while the lead (II) bromide is solid showing that no electric current passes through the solid lead (II) bromide. As the lead (II) bromide melts, the bulb gives out light. After a while, a brown coloration is observed at the anode and a shiny grey solid (lead) is deposited at the cathode.



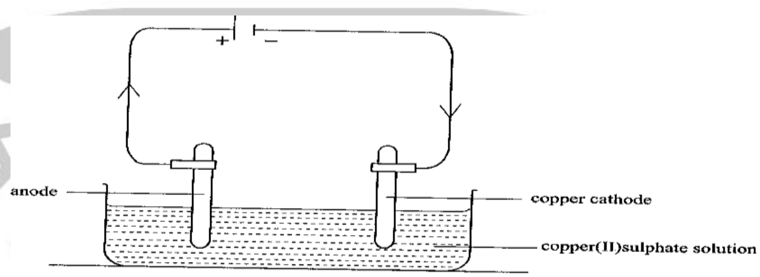
Reaction at the anode:

$2Br^-_{(1)} \rightarrow Br_{2(l)} + 2e^-$

Reaction at the cathode:

$Pb^{2+}_{(l)} + 2e^- \rightarrow Pb_{(s)}$

Electrolysis of copper (II) sulphate solution (using copper electrodes – active electrodes)



Ions present:

From copper (II) sulphate: $Cu^{2+}$and SO_4^{2-}

From water: H^+ and OH^-

Reaction at cathode:

Copper (II) ions and hydrogen ions migrate to cathode. Copper (II) ions are discharged because they are less reactive than hydrogen ions. Copper (II) ions gain electrons from the cathode and copper is deposited.

A brown layer of copper is deposited at the cathode and thus the mass of the cathode increases.

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

Reaction at anode:

Both the sulphate and hydroxide ions migrate to the anode but none loses its electrons. Instead, the copper anode itself loses electrons and as it does so, it becomes copper (II) ions which dissolves in solution.

The anode electrode dissolves and its mass decreases.

$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$

Electrolysis of copper (II) sulphate solution (using copper cathode and platinum anode)

Ions present: From copper (II) sulphate: $Cu^{2+}$and SO_4^{2-}

From water: H^+ and OH^-

Reaction at cathode:

Copper (II) ions and hydrogen ions move to the cathode. Copper (II) ions being less reactive than hydrogen ions are discharged. Copper (II) ions gain electrons and copper is deposited.

The blue color of the electrolyte (copper (II) sulphate solution) fades as copper is deposited because copper (II) ions are removed from the solution.

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

Reaction at anode:

Sulphate ions and hydroxide ions move to the anode. Hydroxide ions being less reactive than sulphate ions are discharged by giving up their electrons.

Bubbles of a colorless gas (oxygen) are formed at the anode.

$4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)} + 4e^-$

The overall equation is obtained by adding the two equations after multiplying the first equation by 2, to obtain the same number of electrons in both equations.

$2Cu^{2+}_{(aq)} + 4e- \rightarrow 2Cu_{(s)}$

$4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)} + 4e^-$

Overall equation:

$2Cu^{2+}_{(aq)} + 4OH^-_{(aq)} \rightarrow 2Cu_{(s)} + 2H_2O_{(l)} + O_{2(g)}$

Trial experiment

Electrolysis of copper(II) sulphate solution

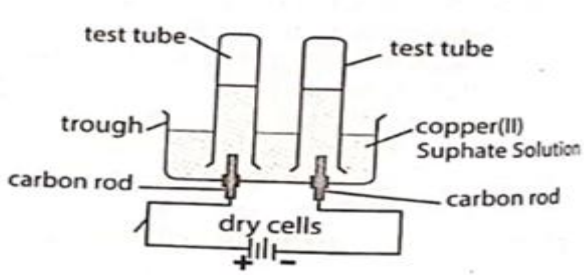
Using lead pencils/carbon electrodes

What is needed:

- dry cells
- connecting wires
- beaker
- copper(II) sulphate solution
- switch
- trough
- lead pencils of about 10 cm length each

What to do

1. Sharpen two lead pencils to expose the carbon rods
2. Arrange the apparatus as shown in Figure 1.12



Sample questions

1. Explain the observed colour change, if any.
2. State the electrode where;
   a) oxidation occurred
   b) reduction occurred
3. Write the overall equation for the reaction.
4. Why are carbon and platinum electrodes usually preferred during electrolysis?

Answers

(Take note of the symbols! Some are not correct symbols and not written correctly, so do the necessary)

1. The blue colour of the copper(II) solution turns colourless
   - Bubbles of a colourless gas (oxygen) are formed at the anode and the brown solid deposited at the cathode

     At the anode: 4OH-(aq) → O2(g) + $2H_2O$(l) +4e-

     At the cathode: Cu_2+(aq) + 2e- → Cu(s)
2. a) at the anode

   b) at the cathode
3. $2Cu_2$+(aq) + 4OH- (aq)→ 2Cu(s) +O2+2H2O(l)
4. These electrodes are generally inert and hence do not participate in the reaction.

Electrolysis of concentrated copper (II) chloride solution *(using carbon anode and copper cathode)*

Copper (II) ions from copper (II) chloride and hydrogen ions from water migrate to the cathode. Copper (II) ions are discharged because they are more concentrated than hydroxide ions, thus chlorine gas (greenish yellow gas) is liberated.

Reaction at cathode:

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

Chloride ions from copper (II) chloride and hydroxide ions from water move to the anode but chloride ions are discharged because they are more concentrated than hydroxide ions, thus chlorine gas (greenish yellow gas) is liberated.

Reaction at anode:

$2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$

The overall equation is obtained by adding the two equations.

$Cu^{2+}_{(aq)} \rightarrow Cl_{2(g)} \rightarrow Cu_{(s)} + 2e^-$

However, if the copper (II) chloride solution is very dilute, some discharge of hydroxide ions will also occur. As the copper (II) chloride solution is diluted, there will not be a point at which chlorine ceases to be produced and oxygen replaces it. Instead, a mixture of the two gases will come off, with the proportion of oxygen gradually increasing. The same case arises in the electrolysis of sodium chloride solution and hydrochloric acid, because the same anions are involved.

Application of electrolysis

(i) Electroplating

This is the process of coating layer of a metal onto another metal by the process of electrolysis.

Importance of electroplating

- Electroplating is done to protect metals from corrosion
- To improve the appearance of metals (for decoration purposes).

The metal to be plated is made the cathode in a suitable electrolyte containing ions of the plating material.

For example, during silver plating, the metal to be plated is made the cathode in a silver salt solution as an electrolyte and pure silver is made the anode.

The silver salt solution contains positively charged silver ions which are attracted to the cathode (metal to be plated). Once there, they gain electrons to form silver atoms.

$Ag^{+}_{(aq)} + e^{+} \rightarrow Ag_{(s)}$

The anode (plate of pure silver) loses its electrons and forms silver ions which dissolve in the solution to replace the ones moving to the cathode.

The process continues until an adequate layer of silver has been deposited on the metal being plated.

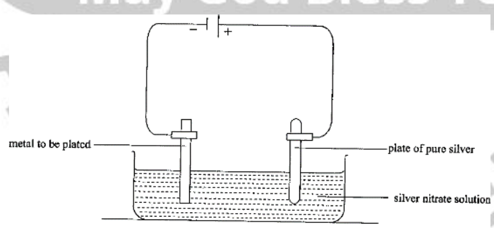


Fig 7.6 Silver plating

Other metals which can be used to coat other metals include chromium, nickel, copper and gold. When iron is to be chromium plated, it is first electroplated with nickel to prevent corrosion and then with chromium.

For successful electroplating, the material to be plated should be clean and the electric current, temperature and concentration of the electrolyte should be exactly right.

When a very low current is used, electrolysis proceeds very slowly and a very smooth deposit can be obtained.

(ii) Purification of metals

Metals such as copper and zinc may be refined, that is purified by electrolysis. The impure metal is made the anode and the pure metal the cathode.

The electrolyte is a solution containing the metal ions.

(iii) Extraction of metals

Reactive metals such as aluminium and sodium are extracted by electrolysis of the fused electrolytes

(iv) Manufacture of chemicals

The most important example is the manufacture of sodium hydroxide, chlorine and hydrogen using the flowing mercury cathode cell.

In the manufacture of sodium hydroxide, concentrated chloride solution (brine) is electrolyzed using a graphite (carbon) anode and a flowing mercury cathode.

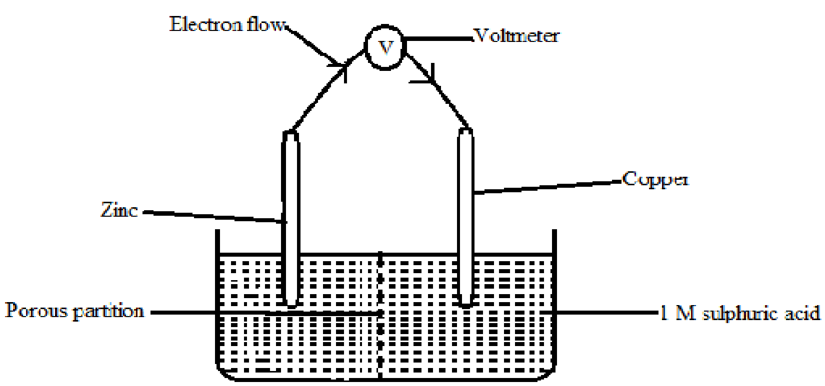
ELECTROCHEMICAL CELLS

Electrochemical cells are cells that produce electricity from chemical reactions i.e., chemical energy is converted to electric energy.

Any cell that generates an electric current from a chemical reaction is called **a galvanic or voltaic cell.**

The common simple galvanic or voltaic cell is that of zinc strip and copper strip dipped into 1M sulphuric acid.

The strips are connected through a voltimeter as shown below.



When the connection is complete, the voltmeter indicates a voltage of about 1.0v, implying that electricity has been generated.

The zinc strip dissolves to form zinc ions by losing electrons.

$Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e$

The electrons lost by the zinc electrode move through the wire to the copper strip and they are gained by the positively charged hydrogen ions around the copper strip hence forming hydrogen gas.

$2H^{+}_{(aq)} + 2e \longrightarrow H_{2(g)}$

The cell reaction:

This is the overall reaction of the cell. It is obtained by adding the two half-cell equations. i.e

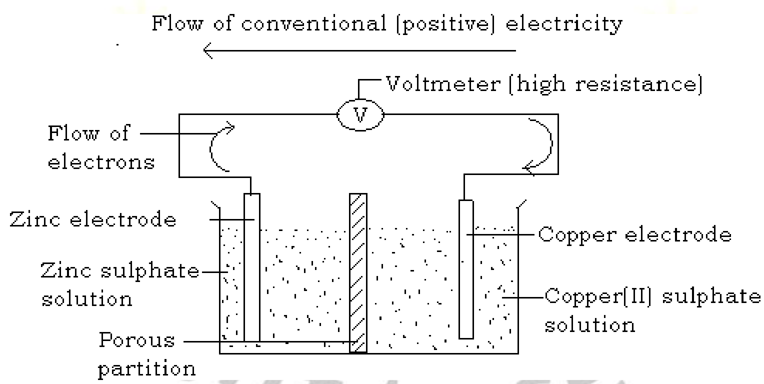
$Zn_{(s)} + 2H^{+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + H_{2(g)}$

Note:

1) The electrode containing a strip of a more reactive element becomes the anode whereas that containing a strip of a less reactive element becomes the cathode.
2) At the cathode there is reduction reaction whereas at the anode there is oxidation reaction.
3) Reduction is the addition of electrons to the substance whereas oxidation is the removal of electrons from the substance.
4) Electrons move from the electrode that produces them during the half-cell reaction (the anode) through the wire to the other electrode that does not produce them (cathode) and electric current flows in a direction opposite to that of the electrons.

The Daniell cell:

The Daniell cell (a primary cell) consists of zinc half-cell and copper half-cell separated by a porous partition a shown in the diagram below.



Reaction at the zinc electrode

Zinc is more reactive than copper. This means that zinc becomes the anode and since at the anode there is oxidation, the zinc strip losses electrons hence dissolve in solution forming zinc ions.

$Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e$ **(this is an oxidation reaction)**

Note:

- The zinc half-cell can be represented as: $\mathbf{Zn_{(s)} / Zn^{2+}_{(aq)}}$
- Electrons will flow from the zinc electrode to the copper electrode through the wire since the zinc electrode is the one that produces them during the oxidation reaction as shown above.

Reaction at the copper electrode

Copper is less reactive than Zinc. This means that copper becomes the cathode and since at the cathode there is reduction, Copper ions gain electrons and hence copper is deposited on the electrode.

$Cu^{2+}_{(aq)}$ + 2e $\longrightarrow$ $Cu_{(s)}$ **(this is a reduction reaction)**

Note:

- The copper half-cell can be represented as: **$Cu^{2+}_{(aq)}$ / $Cu_{(s)}$**
- Current will flow from the copper electrode to the zinc electrode since the electrons are flowing in the opposite direction as shown in the diagram above.

The cell reaction (Overall reaction)

It is obtained by adding the two half-cell equations

$Zn_{(s)}$ + $Cu^{2+}_{(aq)}$ $\longrightarrow$ $Zn^{2+}_{(aq)}$ + $Cu_{(s)}$ **(this is a redox reaction)**

The cell notation:

This is written by combining the two half-cells with the electrode of the more reactive element (anode) placed on the left while that of the less reactive element (cathode) placed on the right and the two separated by symbol //.

Therefore, the cell notation of the reaction is; $Zn_{(s)}$ / $Zn^{2+}_{(aq)}$ // $Cu^{2+}_{(aq)}$ / $Cu_{(s)}$

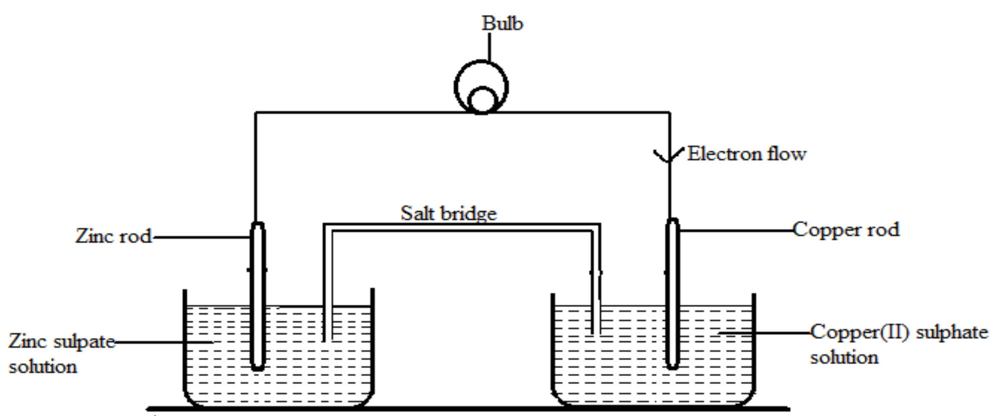
The symbol // represents the separation of the two half-cells by a porous medium.

Note:

Since at the cathode, copper ions (responsible for the blue color) are being removed from the solution and converted to copper solid, the blue color of the solution changes to colorless.

Experiment to show the electron transfer in a redox reaction:

Zinc sulphate and copper (II) sulphate solutions are places in separate beakers into which zinc rod and copper rod are dipped respectively. The apparatus is set up as shown below.



When the circuit is complete the bulb shows light to indicate that the electrons are flowing through the wire.

At the zinc electrode, zinc atoms lose electrons in the process called oxidation.

$Zn_{(s)}$ $\longrightarrow$ $Zn^{2+}_{(aq)}$ + 2e **(this is an oxidation reaction)**

These electrons lost from the zinc rod flow through the external circuit to the copper rod where they combine with copper (II) ions to form copper atoms in a process called reduction.

$Cu^{2+}_{(aq)}$ + 2e $\longrightarrow$ $Cu_{(s)}$ **(this is a reduction reaction)**

The two half-cells can be connected using a salt bridge.
A salt bridge is a bent glass tube containing sodium or potassium salts from strong acids. For example, potassium chloride and sodium nitrate or pieces of filter paper wetted with solutions of these salts.
The salt bridge is used not only to complete the circuit but also to minimize excessive mixing of the solutions contained in the two half-cells.

Uses of electrochemical cells

- They are used making dry cell batteries
- They are also used in making rechargeable cells such as the lead-acid batteries.

End of Chapter Questions

1. a) Copy and complete the following sentences.
   i) Oxidation is the loss of ___________ , or the ________ of oxygen.
   ii) Reduction is the gain of ___________ , or the ________ of oxygen.
   b) State, with reasons, if the following reactions are redox reactions:
   i) $2I^-(aq) + Br_2(aq) \longrightarrow I_2(aq) + 2Br^-(aq)$
   ii) $2Ca(s) + O_2(g) \longrightarrow 2CaO(s)$
   iii) $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
2. State whether the underlined substances have been oxidised or reduced.
   a) $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$
   b) $2Ag(s) + Pb^{2+}(aq) \longrightarrow 2Ag^+(aq) + Pb(s)$
   c) $2Ni(s) + Sn^{2+}(aq) \longrightarrow Ni^{2+}(aq) + Sn(s)$
3. When copper(II) oxide dissolves in dilute sulphuric acid, copper(II) sulphate and water are produced.
   a) Write the balanced chemical equation for the reaction.
   b) Discuss the changes in the oxidation numbers of the species involved in this reaction.
   c) Is copper oxidised or reduced in the reaction?

4. Study *Figure 1.1:* and answer the questions that follow.

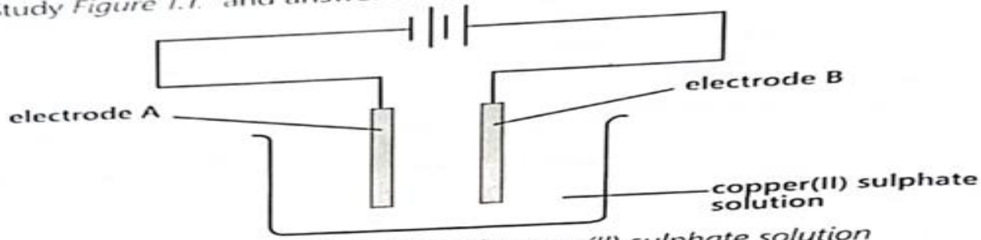


Figure 1.1 *: Electrolysis of copper(II) sulphate solution*

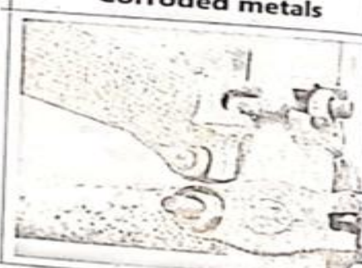
a) Give the names of the electrodes A and B.
b) Which electrode is the oxidising electrode?

5. a) What is an electrolyte?
b) Classify the following substances as strong electrolytes, or weak electrolytes, or non-electrolytes: acetic acid, ammonium hydroxide, ammonium chloride, carbon tetrachloride, dilute hydrochloric acid, sodium acetate, dilute sulphuric acid.

Compare different textbooks to see different sample Activities of integration

Sample Activity of Integration

Rancidity (spoilage of food), corrosion and change in colour of sliced fruits and other food stuffs are some of the causes of losses in both industries and homesteads.

Rancidity in meat	Aerial oxidation in fruits	Corroded metals
	one side ripe, one side not 	

Unsaturated fats from, for example, fish and chicken, are more prone to rancidity than saturated fats.

Rancidity lowers the nutrient content of food and brings about unpleasant smell to those food stuffs, thus rendering them unfit for human consumption.

Corrosion affects metallic objects, causing them to lose efficiency.

As a chemistry learner, write a report to sensitise people in the following businesses about the causes and ways of minimising randicity, change in colour of food stuffs and corrosion of metals.

a) Butchery b) Construction hardware

c) Hotel management