REDOX CHEMISTRY

OXIDATION STATE

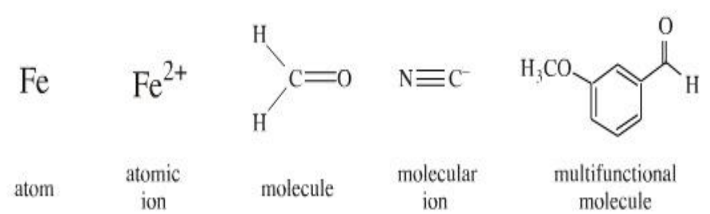
Learning Outcomes:

- Define :
  - ***a)*** oxidation as loss of electrons (1)
  - ***b)*** reduction as gain of electrons (1)
  - ***c)*** oxidation number (1)
- Give examples of species undergoing oxidation or reduction (1 or 2)
- Determine the oxidation state (number) of each atom in a given :
  - ***a)*** Element (1)
  - ***b)*** Molecule (1) or (2)
  - ***c)*** Monatomic ion (1)
  - ***d)*** Diatomic ion (2)
- Explain oxidation and reduction reactions in terms of oxidation states (numbers) (3)
- Discuss oxidation and reduction reactions in terms of oxidation states (numbers) (4)

Introduction to Redox Chemistry

- Redox chemistry is concerned with net electron flow to and from a *defined centre* during a chemical reaction.

- A defined centre may be:



- The word **"redox"** describes the chemical reactions where one substance is **reduced** and another is **oxidised**.
- Redox reactions initially were named after the gain of oxygen or loss of hydrogen (oxidation) and the loss of oxygen or gain of hydrogen (reduction).
- Later it was found that many redox reactions do not involve oxygen or hydrogen at all; all redox reactions, however, involve the transfer of electrons.

- Substances which are **oxidized**:
  (i) Gain of oxygen or
  (ii) Loss of hydrogen or
  (iii) Loss of electrons

- Substances are said to be **reduced** if it:
  (i) Lose oxygen
  (ii) Gain Hydrogen
  (iii) Gain of electrons

Oxidation	Reduction
• Gain of oxygen • Loss of hydrogen • Loss of electrons	• Loss of oxygen • Gain of hydrogen • Gain of electrons

- A useful tool in recognising redox reactions, and for determining what is oxidised and what is reduced in a reaction, involves the use of **oxidation numbers**.

OXIDATION NUMBER

- The oxidation number (symbol **ON**) describes the "degree" to which an element has been oxidised or reduced.
- The oxidation number is usually written above the atom for which the number is assigned.

  Example: Oxidation number of Permangante ion, MnO_4^-
  For the permanganate ion, MnO_4^-, the oxidation number of manganese, Mn is +7 and the oxidation number of oxygen, O is -2.

  ON (Mn) = +7 and ON (O) = -2 or $\overset{+7}{Mn}\overset{-2}{O_4^-}$

Note: Oxidation number and the charge on a polyatomic ion are not the same thing and should not be confused.

- Oxidation numbers are assigned by a set of rules: (Refer to the table on the next page for the rules to follow when assigning oxidation number).

RULES FOR ASSIGNING OXIDATION NUMBERS

Rules	Examples
1. When atoms exist as elements, they have an oxidation number of zero.	Na, Cl_2, Ne, C and H_2 all have an oxidation number of zero.
2. The oxidation number of a monoatomic (one atom) ion is the same as the charge on the ion.	Cu^{+2} has oxidation number +2, Cl^- has oxidation number of -1.
3. Hydrogen in ion has an oxidation number of +1, except in metal hydrides where it is -1.	Hydrogen in H_2O, CH_4 and NH_3 has oxidation number +1. In NaH, a metal hydride, the oxidation number is -1.
4. Oxygen in compounds has an oxidation number of -2, except in hydrogen peroxide, H_2O_2, where its oxidation number is -1.	Oxygen in MgO, H_2SO_4, H_2O and $KMnO_4$ has oxidation number -2. In hydrogen peroxide, H_2O_2, the oxidation number of oxygen is -1.
5. For polyatomic ions (ions containing more than one atom), the sum of the oxidation numbers equals the charge of the ion.	For NH_4^+, the sum of the oxidation number is +1. For SO_4^{-2}, the sum of the oxidation number is -2.
6. The sum of the oxidation numbers of atoms in a molecule is zero.	The sum of the oxidation numbers for the atoms in H_2SO_4, C_4H_{10} and H_2O is zero.

***Note:** *It is important to use the '+' sign and the '-' sign when writing ONs: use '+1' not '1', '+2' not '2',etc.*

Example 1: Finding Oxidation Numbers

Find the oxidation number of carbon in the polyatomic ion, CO_3^{-2}.

Solution:

1. The charge on the ion CO_3^{-2} is -2, so the sum of the oxidation numbers is -2 (according to Rule #5), so:

   ON(C) + 3 x ON(O) = -2

2. The oxidation number of oxygen is -2 (according to Rule #4).

3. Total contribution from oxygen is 3 x (-2) = - 6, since there are three oxygen.

4. The oxidation number of carbon is +4, since
   ON (C) + (-6) = -2
   $\therefore$ ON (C) = -2 – (-6)
   ON (C) = +4

   Carbonate ion can also be written as: $\overset{+4\,-2}{\mathbf{CO_3}}{}^{-2}$

Example 2: Finding Harder Oxidation Numbers

Calculate the oxidation number of sulphur in the molecule H_2SO_4.

Solution:

1. The sum of the oxidation number is zero, since H_2SO_4 is a neutral molecule (according to Rule #6).

   $\therefore$ 2 x ON **(H)** + ON **(S)** + 4 x ON **(O)**= 0

2. The oxidation number of hydrogen is (+1) (Rule #3). Total contribution from hydrogen is +2, since there are two Hydrogen.

3. The oxidation number of oxygen is -2 (Rule #4). Total combination from oxygen is 2 x (-4) = -8, since there are 4 O's.

4. The oxidation number of sulphur is H_2SO_4 is thus +6, since:
   (+2) + **ON (S)** + (-8) = 0
   ON (S) - 6 = 0
   $\therefore$ **ON (S) = +6**

   $\overset{+1\;+6\;-2}{\mathbf{H_2SO_4}}$

- Some atoms take a variety of oxidation states, depending upon the compounds they are found in.
- This in turn means their oxidation numbers can vary.
- The properties of some ions vary greatly depending on which oxidation state they are in.

Example: Oxidation states of iron ions
The properties of the two oxidation states of iron ions are quite different.

Ion	Oxidation state/number	Colour	Reactivity
Fe^{+2}	+2	Pale blue/green	Unstable (easily oxidised to Fe^{+3})
Fe^{+3}	+3	Orange-red/brown	stable

USING OXIDATION NUMBER TO DETERMINE OXIDATION AND REDUCTION

- In all redox reactions, the oxidation number(s) of one or some of the atoms involved will change.
- Calculating changes in oxidation number is useful to determine which reactant is oxidized and which is reduced.

- If the oxidation number of a reactant or an atom in that reactant *increases* during a reaction – the reactant has been **oxidized**.
- If the oxidation number of a reactant or an atom in that reactant *decreases* during a reaction – **reduction** has occurred.

Example:

a) *Magnesium metal forms magnesium ions:*

$$Mg \longrightarrow Mg^{+2} + 2e^-$$

ON is zero (Rule #1)   *ON is +2 (Rule #2)*

The oxidation number of magnesium increases from 0 to +2, so magnesium metal is said to be oxidized to magnesium ions.

b) *Oxygen gas forming oxide ions:*

$$O_2 + 2e^- \longrightarrow 2O^{-2}$$

ON is zero (Rule #1)   *ON is -2 (Rule #2)*

The oxygen gas is <u>reduced</u> to oxide ions, since the oxidation number has decreased from 0 to -2.

- Oxidation and reduction reactions always occur together, so if one reactant is oxidized then another reactant must be reduced.

<u>Example: Magnesium and Oxygen</u>

When magnesium burns in oxygen, the magnesium metal is oxidized to magnesium ions and the oxygen is reduced to oxide ions.

$$\overset{0}{2Mg_{(s)}} + \overset{0}{O_{2(g)}} \longrightarrow 2\overset{+2\ -2}{MgO_{(s)}}$$

Note: If there is no change in oxidation number for any atom in a reaction, the reaction is NOT a redox reaction.

<u>Exercise: Oxidation and Reduction</u>

1. What is the oxidation number of:
   a) Manganese in manganese dioxide, MnO_2
   b) Zinc in zinc metal
   c) Zinc in zinc ions, Zn^{+2}
   d) Nitrogen in nitrate ions, NO_3^-
   e) Sulphur in sulphur dioxide, SO_2.
2. Write the oxidation state of chlorine in each of the following:
   a) Cl^- b) Cl_2 c) HOCl   d) $HClO_4$
3. Give the oxidation numbers of each type of atom present:
   a) Mg
   b) N_2
   c) H_2O
   d) CaH_2
   e) Ag^+
   f) Fe^{+3}
   g) SO_2
   h) SO_3
   i) CH_4
   j) MgO
   k) CO_2
   l) S_8
   m) CO
   n) MnO_4^-
   o) CO_3^{-2}
   p) H_2O_2
   q) $S_2O_3^{-2}$
   r) $Cr_2O_7^{-2}$
   s) $NaHCO_3$
   t) $CuCO_3$
   u) Na_3PO_4
   v) H_3PO_4
   w) $CuSO_4$
4. Give the oxidation number of nitrogen in each of the following:
   a) NH_3
   b) NO_2
   c) NO_3^-
   d) N_2O
   e) NO
   f) N_2
   g) N^{-3}

OXIDANTS & REDUCTANTS

Learning Outcomes:

1.Oxidants :

- ◻ Define oxidants and oxidizing agent (1)
- ◻ Describe the role of an oxidant in a chemical reaction (2)
- ◻ Relate being an oxidant to oxidation state (3)
- ◻ Identify in a chemical reaction the oxidant (oxidizing agent) (1)
- ◻ Give a reason for a species being regarded as an oxidant (1)
- ◻ Name or Identify or State the formula or State the colour of a common oxidizing agents (oxygen, chlorine, metals with dilute acids, hydrogen peroxide, permanganate, dichromate) (1)
- ◻ Describe possible observations made when the common oxidizing agents undergo chemical reactions (2)
- ◻ Analyse the colours of permanganate or dichromate in their different oxidation states under various acidic conditions (3)
- ◻ Write a half equation for the change an oxidant undergoes in a chemical reaction. (Use the common oxidizing agent above)(2)

2.Reductants:

- ◻ Define reductant or reducing agent (1)
- ◻ Identify in a chemical reaction the reductant (reducing agent) (1)
- ◻ Give a reason for a species being regarded as a reductant (1)
- ◻ Describe the role of a reductant or reducing agent in a chemical reaction (2)
- ◻ Relate being a reductant to its oxidation state (3)
- ◻ Name or Identify or State the formula oe State the colour of a common reducing agents: (metals e.g zinc, magnesium and iron; carbon, sulphur dioxide, carbon monoxide) (1)
- ◻ Describe possible observations made when the common reducing agent undergo chemical reactions (2)
- ◻ Analyse given reductants in terms of their oxidation states (3)
- ◻ Write the balanced half-reaction for a reductant reaction (2)

3. Redox reaction:

- ◻ Define redox reaction (1)
- ◻ Describe the nature of a redox reaction (2)
- ◻ Explain the mechanisms involved in a redox reaction (3)
- ◻ Name or Identify or State a redox reaction (1)
- ◻ Give a reason for a reaction being regarded as a redox reaction (1)
- ◻ Describe possible observations made in a given redox reactions (1)
- ◻
- ◻ Explain why oxidation and reduction must occur simultaneously in a redox reaction (3)
- ◻ Combine half equations of reduction and oxidation reactions to write full balanced equation for a redox reaction (3)

In redox reactions:

- The reactant which is *oxidized* is called the **reductant** or (**reducing agent**), since it *reduces* the other reactant.
- The reactant which is *reduced* is called the **oxidant** (or **oxidizing agent**) since it *oxidizes* the other reactant.

  Example: Oxidants and Reductants

  For

  $$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$$

  Magnesium is the reductant because it has been oxidized. Oxygen is the oxidant because it has been reduced.

- The table below lists some of the common oxidizing agents. The oxidation number of the oxidant before and after it is reduced is shown underneath the relevant atoms.

Common Oxidizing Agents

Name	Oxidation number before and after & colour change
Permanganate	$\underset{+7}{MnO_4^-} \longrightarrow \underset{+2}{Mn^{+2}}$ Purple → colourless
Dichromate	$\underset{+6}{Cr_2O_7^{-2}} \longrightarrow \underset{+3}{Cr^{+3}}$ Orange → blue (green)
Hydrogen Peroxide	$\underset{-1}{H_2O_2} \longrightarrow \underset{-2}{H_2O}$ Colourless → colourless
Chlorine	$\underset{0}{Cl_2} \longrightarrow \underset{-1}{Cl^-}$ Pale green → white

Common Reducing Agents

Name	Oxidation number before and after & colour change
Metals high on the activity series eg. Mn, Zn and Fe	$Mg \rightarrow Mg^{+2}$ $Zn \rightarrow Zn^{+2}$ $Fe \rightarrow Fe^{+2}$
Carbon (C) and Carbon Monoxide (CO)	$C \rightarrow CO_2$ $CO \rightarrow CO_2$
Sulphur Dioxide	$SO_2 \rightarrow SO_4^{-2}$

Example 1: *Reduction of $Cr_2O_7^{-2}$ to Cr^{+3}*

Example 2: Oxidation of SO_2 to SO_4^{-2}

BALANCING REDOX EQUATIONS

- To write a balanced equation for a redox reaction, the reaction is broken into two **half equations**:

  (i) **The reduction half-equation** – the oxidant is *reduced* to a product.
  (ii) **The oxidation half-equation** – the reductant is *oxidized* to a product.

- Products are identified by carefully observing the redox reaction to see what forms.

Balancing Half-equations

- Half-equations are balanced using certain rules. The rules should be followed in the correct order for both the half equations.

  Rules for balancing half equations.

  1. Identify the reactant and the product that it forms.
  2. Balance the atoms undergoing a change in oxidation number.
  3. Balance oxygen atoms by adding water to the appropriate side.
  4. Balance hydrogen atoms by adding hydrogen ion to the appropriate side.
  5. Balance the charges by adding electrons to the most positive side.

BALANCING THE OVERALL REDOX EQUATIONS

- The oxidation and reduction half-equations (or ion-electron equations) are **added together** to give the *balanced*, *overall equation* for the redox reaction.
- Before this can be done, one or both half-equations may need to be multiplied by an appropriate number, so that the number of electrons lost in oxidation equals the number of electrons gained in reduction.
- In Example 1, the $Cr_2O_7^{-2}$ reduction half-equations needed $6e^-$ on the LHS.
- In Example 2, the SO_2 oxidation half-equation needed $2e^-$ on the RHS.
- A balanced, overall equation if acidified $Cr_2O_7^{-2}$ and SO_2 are reacted together is shown:

  Example: Sulphur dioxide reacting with dichromate.

Use ***Example 1*** and ***Example 2*** to answer the following Questions.

a) Which reactant was oxidized? : _______________

b) Which reactant was reduced? : _______________

c) Which reactant was the oxidant? : _____________

d) Which reactant was the reductant? _____________

Exercise : Redox Equations

1. Balance the following half-equations and indicate whether they involve ***oxidation*** or ***reduction:***

   a) $Na^{+} \longrightarrow Na$
   b) $H^{+} \longrightarrow H_2$
   c) $Fe^{+2} \longrightarrow Fe^{+3}$
   d) $Cu \longrightarrow Cu^{+2}$
   e) $SO_2 \longrightarrow SO_4^{-2}$
   f) $Cr_2O_7^{-2} \longrightarrow Cr^{+3}$

2. Balance the following equation and give the colours of the reactant and product:

   Fe^{+3} (aq) $\longrightarrow$ Fe^{+2} (aq)

3. Which one of the following is an oxidation react product?

   A. $MnO_4^{-} \longrightarrow Mn^{+2}$
   B. $S \longrightarrow S^{-2}$
   C. $Fe^{+3} \longrightarrow Fe^{+2}$
   D. $Cl^{-} \longrightarrow Cl_2$

4. For the ion-electron equation

$$MnO_4^{-} + \mathbf{x}H^{+} + \mathbf{y}e^{-} \longrightarrow Mn^{+2} + \mathbf{z}H_2O$$

   The numerical number for x, y and z and in order to from left to right are:

   A. 8,1, 6
   B. 8,3, 4
   C. 8, 5, 4
   D. 6, 3, 2

5. Complete and balance the following equations, showing half-equations where you use them:

   a) $Fe^{+2} + MnO_4^{-} \longrightarrow Fe^{+3} + Mn^{+2}$
   b) $Mg + Ag^{+} \longrightarrow Mg^{+2} + Ag$
   c) $Fe^{+2} + Cr_2O_7 \longrightarrow Fe^{3+} + Cr^{+3}$

6. A student is asked to obtain solid copper from a solution of copper sulphate, which contains Cu^{+2}(aq). She is given three reactants:
   A – Hydrochloric acid
   B – magnesium metal
   C – solution of zinc sulphate

   a) Which one of the three reactants will solid copper sulphate solution produces solid copper?
   b) ‘Write a half-equation for the reaction of the reactant in (a).
   c) Will the student need to oxidize or reduce the Cu^{+2} to form Cu(s)?
   d) Write an ion-electron equation for the conversion of Cu^{+2} to form Cu(s)?
   e) Give the overall equation for your reaction converting Cu^{+2} to form Cu(s).
   f) State which reactant is the reducing agent.

7. Calcium oxalate (CaC_2O_4) is one of the minerals found in kidney stones. If a strong acid is added to calcium oxalate, the compound will dissolve and the oxalate ion ($C_2O_4^{-2}$) will be changed to weak acid called oxalic acid ($H_2C_2O_4$). Oxalate ion reacts with $K_2Cr_2O_7$ in anacidic solution. The reaction yields Cr3+ and CO2 among the products.

   a) Determine the oxidation numbers to the atoms indicated by **boldface** type:

   $\mathbf{C}_2O_4^{2-}$: _______________

   $\mathbf{C}O_2$: _________________

   $\mathbf{Cr}_2O_7^{2-}$: _______________

   $\mathbf{Cr}^{3+}$: __________________

   b) Balance the following ion-electron half equations.

   $C_2O_4^{-2} \longrightarrow CO_2$

   $Cr_2O_7^{-2} \longrightarrow Cr^{+3}$

   c) Give the overall balanced equation for the redox reaction

d) In this reaction with $Cr_2O_7^{-2}$, does oxalic acid is acts as an oxidant or reductant?

e) Give a reason to justify your answer for d) above.

8. The equation for the heating of copper oxide with carbon is given below,

$$CuO(s) + C(s) \xrightarrow{Heat} Cu(s) + CO_2(g)$$

a) Explain in terms of oxidation states why this is redox reaction.

b) What is the oxidation state of the chlorine atom (Cl) in $HClO_4$? (1 mark)

9. Soil organic matter is determined by its oxidation with excess potassium dichromate in acidic medium. The excess potassium dichromate is then back titrated with standardized ferrous sulfate as follows;

$$Cr_2O_7^{-2}(aq) + Fe^{2+}(aq) \rightarrow Cr^{+3}(aq) + Fe^{+3}(aq)$$

The reaction between dichromate and ferrous ions is an example of a redox reaction.

a) State the meaning of the term “**oxidizing agent**”.

b) Which is the **oxiding agent** in the reaction between $Cr_2O_7^{-2}$ and Fe^{2+}

c) Write a **half-equation** for the **reduction reaction** using the ion-electron method.

d) Write a **half-equation** for the **oxidation reaction** using the ion-electron method.

)

e) Write a **combined balanced equation** for the overall redox reaction between $Cr_2O_7^{+2}$ and Fe^{+2}.

REDOX APPLICATIONS

Learning Outcomes:

- □ Define:
  - a) Electrolysis (1)
  - b) Anode (1)
  - c) Cathode (1)
- □ Describe the electrolysis of common ionic solutions (e.g. NaCl(aq)) or molten ionic compounds. (e.g. NaCl(l)) (2)
- □ Draw a labelled diagram on the electrolysis of common ionic solutions (e.g. NaCl(aq)) or molten ionic compounds (e.g. NaCl(l)) indicating electron flow or current flow (2)
- □ Given the diagram on the electrolysis of common ionic solutions (e.g. NaCl(aq)) or molten ionic compound (e.g. NaCl(l)), name the different part(s) of the electrolysis cell (1 or 2)
- □ Write half-equation for the oxidation reaction in the electrolysis of common ionic solutions e.g. NaCl(aq) or molten ionic compounds (e.g.NaCl(l)) (2)
- □ Write half-equation for the reduction reaction in the electrolysis of common ionic solutions (e.g.NaCl(aq) or molten ionic compounds e.g. NaCl(l)) (2)
- □ Write balanced equation for the redox reaction in the electrolysis of common ionic solution (e.g. NaCl(aq) or molten ionic compounds e.g.NaCl(l)) (3)
- □ Describe observations made during the electrolysis of some common ionic solutions (e.g. NaCl(aq) or molten ionic compounds e.g. NaCl(l)) at the:
  - a) Anode (2)
  - b) Cathode (2)
- □ Apply oxidation and reduction processes to the electrolysis of some common ionic solutions e.g NaCl(aq) and molten ionic compounds e.g NaCl (l) (3)
- □ Explain why molten ionic substances or ionic solutions conduct electricity and not in the solid state (3)
- □ Discuss the occurrence of oxidation and reduction in settings commonly found in society and the environment e.g. batteries of vehicles; corrosion of metals in vehicles, buildings and bridges; oxidation of foods; galvanic protection with sacrificial electrodes; fuels; breathalyser test (4)
- □ **Define the breathalyser test (1)**
- □ **Identify or Name or State or Give the formula of the oxidant or seductant in the breathalyser test (1)**
- □ **Describe or Explain how a breathalyser test works (2 or 3)**
- □ **Write half equations for the breathalyser test reaction (2)**
- □ **Write a full balanced equation for the breathalyser test reaction (3)**
- □ **Carry-out a simplified breathalyser test (2)**
- □ **Describe the colour change involved in the breathalyser test (2)**
- □ **Interpret a breathalyser test result (3)**
- □ **Analyse the possible changes in the breathalyser test under various acidic conditions (3)**
- □ **Evaluate the mechanisms of a breathalyser test with associated validity (4)**

- Oxidation-reduction reactions, commonly known as redox reactions, are an important class of chemical reactions encountered in everyday processes.
- Redox reactions are at work all around us. Iron rusts, batteries produce electricity, and hydrocarbon fuels (such as wood, oil and coal) are burned to generate energy.
- Common examples of the application of redox reactions in our everyday life will be studied in this unit.

1. ELECTROLYSIS OF IONIC SUBSTANCES

- Electrolysis means using an electric current to seperate a mixture of ions.
- The process involves chemical changes which occur when an electric current is passed through an **electrolyte**. These chemical changes involve **redox reactions.**
- An electrolysis reactin occurs in a piece of equipment called an electrolytic cell.
- An electrolytic cell contains:
  - (i) A DC (direct current) power supply
  - (ii) An electrolyte – a liquid containing ions which are free to move.
    (An electrolyte can be either a molten ionic compound or a solution containing dissolved ions).
  - (iii) Two electrodes – carry current to and from the electrolyte.
    (The negative electrode is the cathode and the positive electrode is the anode. Electrodes are usually made of inert materials such as platinum, Pt, or graphite, C, which are not reduced or oxidised in electrolysis reactions.

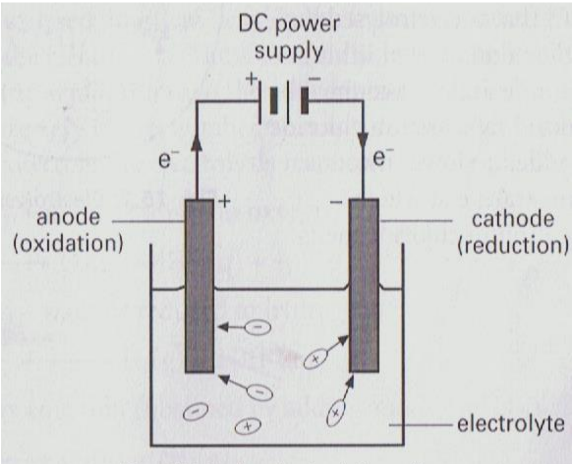


Fig. 15.1: An electrolytic cell

When a current is supplied to the cell:

- **Cations** (positive ions) move or migrate toward the cathode where reduction occurs.
- **Anions** (negative ions) migrate toward the anode where oxidation occurs.

ELECTROLYSIS OF AQUEOUS SOLUTIONS

- An aqueous solution contains water and ions which are free to move.
- When an aqueous solution is electrolysed, some of the ions are oxidised and others reduced.
- The reactions depend on the ease with which different ions can be reduced or oxidised compared with water.
- For metal ions, this relates to the position of their corresponding metal on the activity series; ions of metals 'lower down**'** theseries are more easily reduced than ions further 'up'.

Example: Sodium and Copper chlorides

- When a dilute solution of sodium chloride, NaCl, and a dilute solution of copper (II) chloride, $CuCl_2$, are each electrolysed, water is oxidised at the anode in both cases in preference to chloride ions.

  Anode reaction:

  $2H_2O \longrightarrow H_{2(g)} + 2OH^-_{(aq)}$

- The reaction at the cathode depends upon the metal ion (Na^+ or Cu^{2+}), which in turn depends upon the position of the corresponding metal (Na or Cu) on the activity series.
- In dilute sodium chloride solution, water is reduced at the cathode in preference to sodium ions.

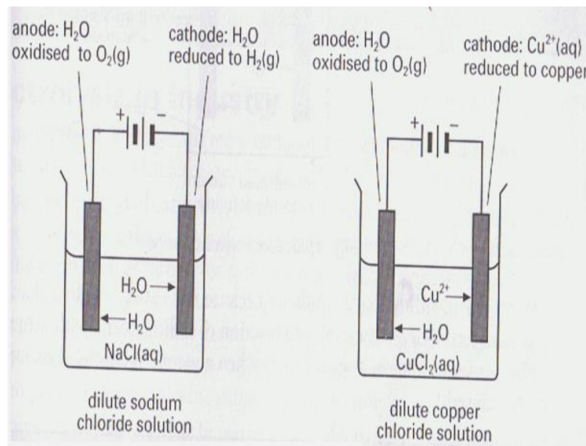
  Cathode reaction for NaCl (aq):

  $2H_2O + 2e^- \longrightarrow H_{2(g)} + 2OH^-_{(aq)}$

- In copper chloride solution, the copper ions (Cu^{+2}) are reduced rather than water, ie Cu^{+2} are reduced in preference to water molecules.

  Cathode reaction for $CuCl_2$ (aq):

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



NOTE:

- In the electrolysis of dilute sodium chloride, water is 'split' into its elements rather than the NaCl being split into its elements.
- The different reduction reactions occurring in the two solutions are consistent with the activity series.
- Copper is lower down the activity series than sodium , so copper ions are more easily reduced than are sodium ions.
- Copper ions are a better oxidising agent than water; when both are present, $Cu^{+2}(aq)$ will be reduced (to Cu(s)) in preference to water being reduced (to $H_2(g)$).
- Copper is below hydrogen on the activity series, sodium is above hydrogen.
- Na^+ ions are a poorer oxidising agent than water – when both are present, water is reduced in preference to Na^+.
- The concentration of ions also affects the reduction and oxidation reactions that occur.

Example: Electrolysis of concentrated sodium chloride

- If a concentrated solution of sodium chloride (called **brine**) is electrolysed, the chloride ions are in high enough concentration to be oxidised at the anode in preference to water.

Anode Reaction:

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

Cathode Reaction:

$2H_2O \longrightarrow H_{2(g)} + 2OH^-_{(aq)}$

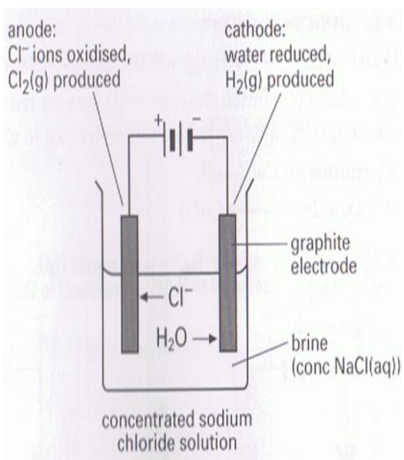


Fig. 15.5: Electrolysis of brine

- Water is still reduced at the cathode because it is easier to reduce than sodium ions. (Given the explosive reaction of sodium metal with water, it is obvious sodium metal does not form when aqueous solutions of its ions are electrolysed.)
- In an electrochemical cell, chemical redox reactions drive the movement of electrons through a wire—generating electricity.

2. COMBUSTION AND EXPLOSIONS

- As with any type of chemical reaction, combustion takes place when chemical bonds are broken and new bonds are formed.
- It so happens that combustion is a particularly dramatic type of oxidation-reduction reaction: whereas we cannot watch iron rust, combustion is a noticeable event.
- Even more dramatic is combustion that takes place at a rate so rapid that it results in an explosion.
- As one might expect from what has already been said about oxidation-reduction, the oxygen is reduced while the carbon is oxidized.

3. OXIDATION: FOOD SPOILING

- At the same time, oxidation-reduction reactions are responsible for the spoiling of food, the culprit here being the oxidation portion of the reaction.
- To prevent spoilage, manufacturers of food items often add preservatives, which act as reducing agents.

4. FORMING A NEW SURFACE ON METAL - CORROSION

- Clearly, oxidization can have a corrosive effect, and nowhere is this more obvious than in the corrosion of metals by exposure to oxidizing agents—primarily oxygen itself.
- Most metals react with O_2, and might corrode so quickly that they become useless, Iron forms an oxide, commonly known as **rust,** but this in fact does little to protect it from corrosion, because the oxide tends to flake off, exposing fresh surfaces to further oxidation.
- Every year, businesses and governments devote millions of dollars to protecting iron and steel from oxidation by means of painting and other measures, such as galvanizing with zinc.
- In fact, oxidation-reduction reactions virtually define the world of iron.

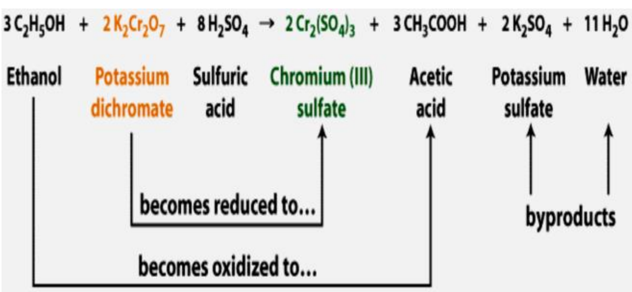
5. COINAGE METALS. –COPPER,SILVER, and GOLD

- Copper, as we have seen, responds to oxidation by corroding in a different way: not by rusting, but by changing color.
- A similar effect occurs in silver, which tarnishes, forming a surface of silver sulfide, or Ag_2S.
- Copper and silver are two of the "coinage metals," so named because they have often been used to mint coins.
- They have been used for this purpose not only because of their beauty, but also due to their relative resistance to corrosion.
- The third member of this mini-family is gold, which is virtually noncorrosive. Wonderful as gold is in this respect, however, no one is likely to use it as a roofing material, or for any such large-scale application involving its resistance to oxidation. Aside from the obvious expense, gold is soft, and not very good for structural uses, even if it were much cheaper.
- Yet there is such a "wonder metal": one that experiences virtually no corrosion, is cheap, and strong enough in

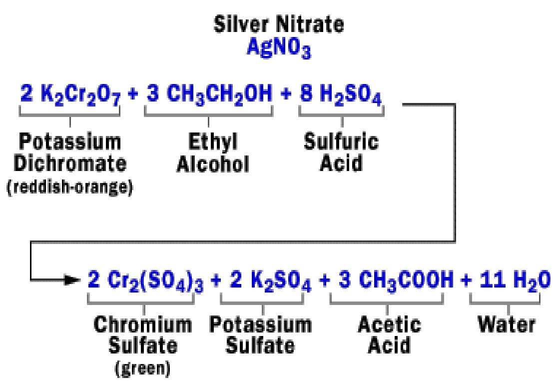
alloys to be used for structural purposes. Its name is aluminium.

6. BREATHALYSER TEST

- A breathalyser is a device for estimating blood alcohol content (BAC) from a breath sample.
- The device is of popular use with Traffic Department to check for drunk drivers and safety of people on the road.
- The device has a **mouthpiece**, a tube through which the suspect blows air, and a **sample chamber** where the air goes.
- To measure alcohol, a suspect breathes into the device. The breath sample is bubbled in one vial through a mixture of sulfuric acid, potassium dichromate, silver nitrate and water. The principle of the measurement is based on the following chemical reaction:



Again:



In this reaction:

1. The **sulfuric acid removes the alcohol from the air** into a liquid solution.
2. The **alcohol reacts with potassium dichromate** to produce:
   - chromium sulfate
   - potassium sulfate
   - acetic acid
   - water

- The silver nitrate is a **catalyst**, a substance that makes a reaction go faster without participating in it.
- The sulfuric acid, in addition to removing the alcohol from the air, also might provide the acidic condition needed for this reaction.

- During this reaction, the *reddish-orange* dichromate ion **changes color** to the *green* chromium ion when it reacts with the alcohol; the degree of the color change is directly related to the level of alcohol in the expelled air.