CHEMISTRY STUDY MATERIALS FOR CLASS 12 (NCERT BASED NOTES OF CHAPTER -03)

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Electrochemistry

Batteries

A battery is basically a galvanic cell in which the chemical energy of a redox reaction is converted to electrical energy. They are of mainly 2 types – primary batteries and secondary batteries.

a) Primary cells:

These are cells which cannot be recharged or reused. Here the reaction occurs only once and after use over a period of time, they become dead E.g. Dry cell, mercury button cell etc.

1. Dry Cell

It is a compact form of Leclanche cell. It consists of a zinc container as anode and a carbon (graphite) rod surrounded by powdered manganese dioxide (MnO_2) and carbon as cathode. The space between the electrodes is filled by a moist paste of ammonium chloride (NH_4Cl) and zinc chloride ($ZnCl_2$). The electrode reactions are:

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

Cathode: $MnO_2 + NH_4^+ + e^- \rightarrow MnO(OH) + NH_3$

Ammonia produced in this reaction forms a complex with Zn^{2+} and thus corrodes the cell.

The cell has a potential of nearly 1.5 V.

2. Mercury cell

Here the anode is zinc – mercury amalgam and cathode is a paste of HgO and carbon. The electrolyte is a paste of KOH and ZnO. The electrode reactions are:

Anode reaction: $Zn(Hg) + 2OH^- \rightarrow ZnO(s) + H_2O + 2e^-$

Cathode reaction: $HgO + H_2O + 2e^- \rightarrow Hg(l) + 2OH^-$

The overall reaction is : $Zn(Hg) + HgO(s) \rightarrow ZnO(s) + Hg(l)$

The cell has a constant potential of 1.35 V, since the overall reaction does not involve any ion in solution.

b) Secondary cells

A secondary cell can be recharged and reused again and again. Here the cell reaction can be reversed by passing current through it in the opposite direction. The most important secondary cell is lead storage cell, which is used in automobiles and invertors.

It consists of lead as anode and a grid of lead packed with lead dioxide (PbO_2) as the cathode.

The electrolyte is 38% H_2SO_4 solution.

The cell reactions are:Anode: $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e-$

Cathode: $PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e- \rightarrow PbSO_4(s) + 2H_2O(l)$

The overall cell reaction is: $Pb(s)+PbO_2(s)+2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$

On charging the battery, the reaction is reversed and $PbSO_4(s)$ on anode and cathode is converted into Pb and PbO_2, respectively.

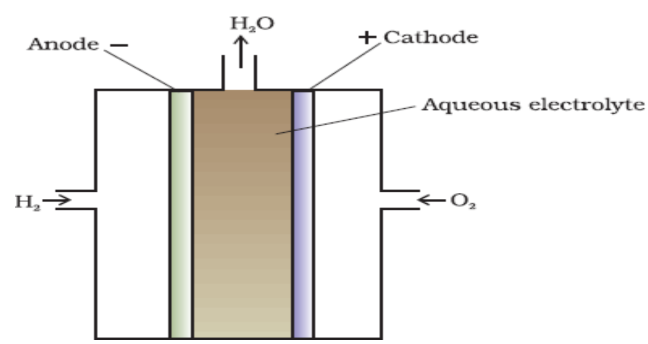
Another example for a secondary cell is nickel – cadmium cell. Here the overall cell reaction is

$$Cd(s)+2Ni(OH)_3(s) \rightarrow CdO(s) +2Ni(OH)_2(s) +H_2O(l)$$

Fuel cells

These are galvanic cells which convert the energy of combustion of fuels like hydrogen, methane, methanol, etc. directly into electrical energy.

One example for fuel cell is ***Hydrogen – Oxygen fuel cell***, which is used in the Apollo space programme. Here hydrogen and oxygen are bubbled through porous carbon electrodes into concentrated aqueous sodium hydroxide solution. To increase the rate of electrode reactions, catalysts like finely divided platinum or palladium metal are filled into the electrodes.



The electrode reactions are:

Cathode: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$

Anode: $2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$

Overall reaction is: $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$

Advantages of Fuel cells

1. The cell works continuously as long as the reactants are supplied.
2. It has higher efficiency as compared to other conventional cells.
3. It is eco-friendly (i.e. pollution free) since water is the only product formed.
4. Water obtained from $H_2 – O_2$ fuel cell can be used for drinking.

Corrosion

It is the process of formation of oxide or other compounds of a metal on its surface by the action of air, water-vapour, CO_2 etc. Some common examples are: The rusting of iron, tarnishing of silver, formation of green coating on copper and bronze (verdigris) etc.

Most familiar example for corrosion is rusting of iron. It occurs in presence of water and air. It is a redox reaction. At a particular spot of the metal, oxidation takes place and that spot behaves as anode.

Here Fe is oxidized to Fe^{2+}.

$$2\ Fe\ (s) \rightarrow 2\ Fe^{2+} + 4\ e^{-}$$

Electrons released at anodic spot move through the metal and go to another spot on the metal and reduce oxygen in presence of H^{+}. This spot behaves as cathode. The reaction taking place at this spot is:

$$O_2(g) + 4\ H^{+}(aq) + 4\ e^{-} - \rightarrow 2\ H_2O\ (l)$$

The overall reaction is: $2Fe(s) + O_2(g) + 4H^{+}(aq) \rightarrow 2Fe^{2+}(aq) + 2\ H_2O\ (l)$

The ferrous ions (Fe^{2+}) are further oxidised to ferric ions (Fe^{3+}) and finally to hydrated ferric oxide ($Fe_2O_3 .\ x\ H_2O$), which is called rust.

<u>Methods to prevent corrosion</u>

1. By coating the metal surface with paint, varnish etc.
2. By coating the metal surface with another electropositive metal like zinc, magnesium etc. The coating of metal with zinc is called galvanization and the resulting iron is called galvanized iron.
3. By coating with anti-rust solution.
4. An electrochemical method is to provide a sacrificial electrode of another metal (like Mg, Zn, etc.) which corrodes itself but saves the object (sacrificial protection).

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