Chemistry Class 9

Chapter 7

Important Q/A's

Electrochemistry

QUESTION .1

What is meant by oxidation and reduction?

ANSWER

Oxidation

The loss of electrons, gain of oxygen or the loss of hydrogen is termed as oxidation.

Reduction

The gain of electrons, loss of oxygen and gain of hydrogen is characterized as reduction.

Explanation

The chemical processes in which electrons are transferred from one chemical specie to another.

Redox reactions, also known as oxidation-reduction reactions, are the name given to these electron-transfer processes.

Energy transformation in the form of heat, light, electricity, etc. accompanies these reactions. Additionally, the addition of oxygen or hydrogen to various compounds is a component of the oxidation and reduction reactions.

Example

Common oxidation-reduction reactions are as follows

Combustion:

$2C_8H_{18}+25O_2 \rightarrow 16CO_{2(}g)+18H_2O$

Disproportionate Reaction:

$3ClO-(aq) \rightarrow ClO3-(aq)+2Cl-(aq)$

Single Replacement Reaction:

$Zn(s)+2HCl(aq) \rightarrow ZnCl2(aq)+H2(g$

Question. 2

What are the common examples of oxidation and reduction reactions?

Answer

Oxidation Examples

Oxidation reactions involve:

1. Gain of oxygen:

$$C + O_2 \rightarrow CO_2 \quad \text{(oxidation of carbon)}$$

2. Gain of electronegative element:

$$Fe + S \rightarrow FeS$$

3. Exit of hydrogen:

$$H_2S + Br_2 \rightarrow 2\,HBr + S \quad \text{(oxidation of sulphide)}$$

4. Exit of electropositive elements:

$$2\,KI + H_2O_2 \rightarrow I_2 + 2\,KOH$$

Reduction Examples

Reduction reactions involve:

1. Gain of hydrogen:

$$N_2 + 3\,H_2 \rightarrow 2NH_3$$

2. Gain of electropositive element:

$$SnCl_2 + 2HgCl_2 \rightarrow SnCl_4 + Hg_2Cl_2$$

3. Exit of oxygen

$$ZnO + C \rightarrow Zn + CO$$

4. Exit of electronegative element

$$2FeCl_3 + H_2 \rightarrow 2FeCl_2 + 2HCl$$

QUESTION .3

Write a note on oxidation-reduction reactions in terms of loss or gain of electrons.

ANSWER

Oxidation In Terms Of Loss Of Electrons

A molecule, atom, or ion commits to oxidation when it loses an electron/s as a result of a process.

An atom losing an electron to create a positively charged cation is known as oxidation in terms of electron transfer. One atom is reduced while another is oxidized during the creation of an ionic connection.

Examples

$Mg \text{ ----> } Mg^{2+} + 2\ e^-$

$O_2 + 4\ e^- \text{ ----> } 2\ O^{2-}$

Reduction In Terms Of Gain Of Electrons

An atom or ion is said to be reduced and the process is known as a reduction if it gains one or more electrons during the chemical reaction.

An atom when gains an electron becomes a negatively charged specie known as an anion is known as a reduction in terms of the gain of electrons.

Examples

$Cl + 1e^- \rightarrow Cl^-$

$F + 1e^- \rightarrow F^-$

QUESTION .4 (a)

What is oxidation state and example?

ANSWER

Oxidation States

The apparent charge positive or negative on an atom or a molecule is referred to as oxidation states.

An atom's oxidation state (also known as its oxidation number) gives information on how many electrons were released from it and, as a result, describes the degree of oxidation of the atom. The imagined charge that an atom would carry if all of its links to other atoms were entirely ionic in nature can be used to determine an atom's oxidation state. For instance, if both of the carbon-oxygen double bonds were entirely ionic, the carbon atom's potential charge would be equal to +4, the oxidation state of carbon in CO2 would be +4. (each oxygen atom would hold a charge of -2 since oxygen is more electronegative than carbon).

QUESTION .4 (b)

How do you assign oxidation states?

ANSWER

Rules For Assigning Oxidation Number

1. Each atom in an element holds up a zero oxidation number in either its free or uncombined state. Each atom in H_2, Cl_2, P_4, Na, Al, O_2, O_3, S_8, and Mg has an oxidation number of zero, making this crystal clear.

2. Ions with a single atom in them have an oxidation number equal to their actual charge.

The oxygen's oxidation number is typically -2 in compounds. Here, there are two exceptions.

Every oxygen atom in peroxides has an oxidation number of zero as in Na_2O_2

Every oxygen atom in superoxides has an oxidation number of -(1/2). Examples include dioxygen difluoride, where the oxygen atom is assigned an oxidation number of 1, and KO2, where oxygen is

3. Except when it is connected to metals that contain two elements, hydrogen has an oxidation number of one. For instance, CaH2 has an oxidation number of -1.

4. When they show up as halide ions in their compounds, fluorine and other halogens have an oxidation number of -1. Iodine, chlorine, and bromine all have positive oxidation numbers when mixed with oxygen.

5. The algebraic sum must equal zero when the oxidation numbers of the constituent atoms of a compound are combined together. When the oxidation numbers of an ion's constituent atoms are added up, the resulting algebraic sum must match the ion's charge in the case of

polyatomic ions. In the case of $(CO3)^{-2}$, the oxidation number of one carbon atom and three oxygen atoms adds up to -2 algebraically.

QUESTION .5

What is oxidizing and reducing agent give examples.

ANSWER

Oxidizing Agents

The specie which oxidizes other substances and itself gets reduced in the process is known as an oxidizing agent. Its oxidation number decreases as it gains electrons due to reduction.

They are chemical compounds whose atoms, as part of a chemical process, take at least one electron away from another atom. According to this definition, reactants in redox processes that go through reduction are oxidizing agents.

A specie that transfers at least one electronegative atom to a species during a chemical reaction is known as an oxidizing agent. Usually, an oxygen atom is the

transferred atom. An electronegative atom is transferred between two reactants in several combustion processes and organic redox reactions.

Reducing Agents

A reducing agent is a specie that, in a redox process, loses electrons to other chemical species and commits to oxidation to a higher valency state

One of the reactants of an oxidation-reduction process is a reducing agent, which reduces the other reactant by discharging electrons to it. The reduction process cannot take place if the reducing agent does not transfer electrons to other components during a reaction. For instance, in the reaction given, $H_2(g) + F_2(g)$ 2HF (g)

Because it provides its electrons to fluorine, hydrogen functions as a reducing agent, decreasing fluorine.

Atoms with low electronegativity, or the capacity to draw bonding electrons, are a common component of good reducing agents, albeit they can also include elements with low ionization energies.

QUESTION .6

What methods are used to extract metals from their ores?

ANSWER

Mining is the process of obtaining (extracting) metal ores from subterranean (**underground)** deposits. The earth's crust contains variable amounts of metal ores. We can utilize the minerals in the earth because metals can be extracted from ores. The finished metals that we see in buildings and bridges are substantially different from the ores.

Ores are made up of the desired metal combination as well as gangue, which are earthy impurities.

Metals are extracted and isolated over the course of several key stages, including:

Ore concentration: At this stage, earthy impurities are removed from the ore.

The process of isolating metal from concentrated ore involves first oxidizing the ore and then reducing it.

Either roasting or calcination followed by heating with a reducing agent are the procedures required.

Metal purification is done in this case for functional reasons.

Aluminium:

Bauxite $AlOx(OH)_3$-2x

Kaolinite $[Al_2(OH)_4\ Si_2O_5]$

Iron:

Haematite Fe_2O_3

Magnetite Fe_3O_4

Siderite $FeCO_3$

Iron pyrites FeS_2

Copper:

Copper pyrites $CuFeS_2$

Malachite $CuCO_3.Cu(OH)_2$

Cuprite Cu_2O

Copper glance Cu_2S

Zinc:

Zinc blend/Sphalerite ZnS

Calamine $ZnCO_3$

Zincite ZnO

QUESTION .7

Which is an electrochemical cell in which electrolysis occurs?

ANSWER

Electrochemical Cell

An electrochemical cell is a device that can use electrical energy to facilitate chemical reactions inside of it or create electrical energy from those reactions. These systems can change chemical energy into electrical energy or the other way around.

Electrochemical Cell Examples

The typical 1.5-volt cell, which powers numerous electrical devices like TV remote controls and clocks, is an example of an electrochemical cell.

Galvanic cells or voltaic cells are those that may produce an electric current as a result of chemical reactions taking place inside of them.

Alternately, **electrolytic cells** are those that undergo chemical reactions when an electric current is passed through them.

Electrolysis

Electrolysis refers to the process of changing a substance's chemical composition by passing an electric current through it. An electron is lost or gained during the chemical reaction, depending on the material (oxidation or reduction). The procedure is carried out in an electrolytic cell, a device made up of positively and negatively charged electrodes that are held apart and submerged in a solution with oppositely charged ions

Electrolytes

Electrolytes are substances that form ions when put in water. Because all of them produce ions when mixed with water, they can be noted as acids, bases, and salts. Due to the mobile action of the positive and negative ions, known as cations and anions, respectively, these solutions conduct electricity. Strong electrolytes do not

generate any neutral molecules in solution; instead, they entirely ionize when dissolved.

QUESTION .8

What is the structure of an electrolytic cell?

ANSWER

Electrolytic Cell

An electrolytic device that employs electrical energy to induce a non-spontaneous redox reaction is known as an electrolytic cell. Certain chemicals can be electrolyzed using electrolytic cells, which are electrochemical cells. For instance, water can be electrolyzed to create gaseous oxygen and gaseous hydrogen with the use of an electrolytic cell. This is accomplished by leveraging the flow of electrons (into the reaction environment) to break through the non-spontaneous redox reaction's activation energy barrier.

Components of Electrolytic Cell:

The following are the three main components of electrolytic cells:

Cathode (which is negatively charged for electrolytic cells) (which is negatively charged for electrolytic cells)

Anode (which is positively charged for electrolytic cells) (which is positively charged for electrolytic cells)

Electrolyte

The cathode and anode transfer electrons through a solution that is provided by the electrolyte. standard electrolytes NaCl,KCl

---

QUESTION .9

Write a note on the uses of electrolytic cells.

ANSWER

Uses

- The main use of electrolytic cells is to create oxygen and hydrogen gas from water.
- They are also employed in the process of removing aluminum from bauxite.
- The technique of creating a thin protective layer of one metal on the surface of another metal, known as electroplating, is another noteworthy use of electrolytic cells.

- Electrolytic cells are used to electrorefine a variety of non-ferrous metals.
- Such electrochemical cells are also employed in electrowinning procedures.
- It can be noted that the industrial production of high-purity copper, high-purity zinc, and high-purity aluminum is virtually usually done by electrolytic cells.

QUESTION .10

What is the function of Daniel cell?

ANSWER

Daniel Cell

The greatest illustration of a galvanic cell that transforms chemical energy into electrical energy is a Daniell cell.

The Daniell cell has two electrodes made of different metals, Zn and Cu, and each electrode is in contact with a solution of its ion, which is, respectively, zinc sulfate and copper sulfate.

Construction

It is a typical galvanic cell that is made to leverage the naturally occurring redox reaction between zinc and cupric ions to generate an electric current.

This cell is a copper container,a saturated $CuSO_4$ solution that serves as a depolarizer. It contains H_2SO_4, which serves as an electrolyte. Zn_2SO_4 is used to submerge an amalgamated zinc rod. $CuSO_4$ crystals are kept in contact with the $CuSO_4$ solution in the transparent layer that surrounds copper vessels just below the upper surface, ensuring that the solution is always saturated.

Salt Bridge

By enabling anions to travel toward the anodic compartment and cations to go toward the cathodic compartment, it maintains electrical neutrality in two compartments.

It is a **glass tube** containing gelatinized potassium chloride or ammonium nitrate.

Gelatin permits ionic flow but forbids mixing in any way.

The ionic mobilities of cations and anions are the same for potassium chloride and ammonium nitrate

Reactions In The Cell

The cell can be represented as

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

In a Daniell cell, metal ions go from one half of the cell to the other over the salt bridge while electrons move from the zinc electrode to the copper electrode through an external circuit.

Here, a circuit external to the cell permits current to move from a copper electrode to a zinc electrode, or a cathode to an anode.

A voltaic cell can be both reversible or irreversible, whereas the Daniell cell is only reversible.

Ion Zn/ZnSO4 half cell, oxidation occurs at the anode

$Zn \rightarrow Zn^{2+} + 2e^{-}$

Ion Cu/$CuSO_4$ half-cell, reduction occurs at the cathode.

$Cu^{2+} + 2e^{-} \rightarrow Cu$

The net cell reaction is

$Zn + Cu^{2+} \rightleftharpoons Zn^{2+} + Cu$

Cell is represented as

$Zn/Zn^{2+} \,||\, Cu^{2+}/Cu$

QUESTION .11

What is the function of dry cell?

ANSWER

Dry Cell

One sort of electric battery, known as a dry cell, is frequently used in portable and home electronics. One or more electrochemical cells make up a battery, a device that transforms chemical energy into electrical energy. After Georges Leclanche created wet zinc-carbon batteries in 1866, the "German scientists Carl Gassner" created dry cells, one of the electrochemical cells, in 1886. Japanese inventor Yai Sakizo created the first modern dry cell in 1887. Dry cell batteries, which range in size from large flashlight batteries to small flashlight batteries and are usually used in wristwatches or calculators, are currently the most widely used batteries

A graphite rod or a metal electrode is covered in a low humidity electrolyte paste in a dry cell, which is made of a metal container. In most cases, the metal container will be made of zinc, whose base serves as a negative electrode (anode) and a positive electrode (carbon road)

(cathode). Manganese dioxide and a non-reversible low moisture electrolyte, such as ammonium chloride paste, that can only produce a maximum voltage of 1.5V, surround it

Cathodic Reaction

$2NH_4^+ + 2MnO_2 \rightarrow Mn_2O_3 + 2NH_3 + H_2O$

Anodic Reaction

$Zn \rightarrow Zn^{2+} + 2e^-$

QUESTION .12

What is down cell used for?

ANSWER

Down Cell

Industry uses the Downs cell to make sodium metal. A cylindrical iron cathode surrounds a central carbon anode in the cell. Chlorine gas produced at the anode is kept from coming into contact with sodium metal produced at the cathode using an iron mesh screen.

In the molten mixture, sodium metal rises to the top and is collected in a watertight storage container. The cell's

strong current flow generates enough heat to keep the mixture molten. Since sodium chloride melts at a temperature of about 800°C, a substantial quantity of energy must be supplied. In addition to the viability issue, the reaction vessel's interior is filled with a thick fog of molten sodium at this temperature.

Calcium chloride is mixed with sodium chloride in a 1:2 ($NaCl : CaCl_2$) part ratio to lower the melting temperature. The mixture doesn't generate sodium fog because of its significantly lower melting point, which is 600°C. Molten sodium is collected from the sodium collecting ring while chlorine is obtained in a funnel from the top of the cell.

Anodic Reaction

$2Cl^-_{(l)} \Rightarrow Cl2_{(g)} + 2e^-$

Cathodic Reaction

$Na^+_{(l)} + e^- \Rightarrow Na_{(l)}$

QUESTION .13

What is Nelson cell used for?

ANSWER

On an industrial scale, brine, an aqueous solution of NaOH, is electrolyzed in Nelson's cell to create caustic soda (sodium hydroxide, or NaOH). It comprises of an iron cathode with a U-shaped perforation suspended in the middle of a steel tank. Asbestos diaphragm is used to line the interior of this iron cathode. The iron cathode contains electrolyte brine.

Working

Na, Cl, H, and OH ions make up sodium chloride's aqueous solution. These ions migrate in the direction of their corresponding electrodes, where redox reactions occur. When electrolysis occurs, the anode discharges Cl ions, and Cl 2 + gas rises into the cell's dome. At the cathode, the H ions are released, and H 2 gas exits through a conduit. The sodium hydroxide solution trickles into the catch basin over time.

Ionization Of Brine

$2NaCl \rightarrow 2Na^{+} + 2Cl^{-}$

Reaction At Anode

$2Cl^{-} \rightarrow Cl_2 + 2e^{-}$

Reaction At Cathode

$4OH^- \rightarrow 2H_2O + O_2 + 2e^-$

$2Na^+ + 2OH^- \rightarrow 2NaOH$

QUESTION .15

What is corrosion and its prevention?

ANSWER

Corrosion is the gradual and ongoing eating away of a metal by its environment. Through the interaction of air and moisture with the metals, a redox chemical reaction occurs. Rusting of iron is the most typical manifestation of corrosion.

Rusting Of Iron

Corrosion is a generic phrase, however, rusting is used to describe the corrosion of iron. The crucial element for rusting is wet air (air having water vapours in it). In the absence of either air or water, there won't be any rusting.

Our current topic of study is the chemistry of rusting. The locations for this process to take place on the iron's surface are stains and dents. This area is known as the anodic region, and an oxidation process occurs here.

$2Fe \rightarrow 2Fe^{+2} + 2e^-$

The object suffers harm from this electron loss. The free electrons travel through the iron sheet until they arrive at a water-surrounded location with a rather high O2 concentration close to the surface. In the presence of H+ ions, electrons decrease the oxygen molecule in this area, which serves as the cathode:

$O_2 + 4H^- + 4e \rightarrow 2H_2O$

Carbonic acid, which is created when CO is present in water, supplies the H+ ions. Because of this, rusting is accelerated by the acidic medium.

Rust does not occur as a result of the whole redox process.

$2Fe + O_2 + 4H^+ \rightarrow 2Fe^{+2} + 2H_2O$

Rust is the salt $Fe_2O_3.nH_2O$ that is created when the Fe^{+2} produced spreads across the surrounding water and reacts with oxygen.

QUESTION .16

What are the different methods to prevent corrosion?

ANSWER

Prevention Of Corrosion

Removal Of Stains

An iron rod's stained areas serve as the location for corrosion. Rusting could be avoided by thoroughly cleaning and removing stains from iron surfaces.

Paints And Greasing

Iron can be protected against rusting by painting or greasing the surface. Modern paints contain a combination of chemicals called stabilizers that, thanks to technological advancement, offer protection against corrosion in addition to prevention against the influence of other atmospheric factors and weathering.

Alloying

A homogenous combination of one metal and one or more other metals or non-metals is called an alloy. Iron has been alloyed with other metals, which has proven to

be a very effective anti-rusting treatment. The term "stainless steel" is the best illustration of alloying.

excellent combination of nickel, chromium, and iron.

Metallic Coating

Coating the metal with another metal is the greatest approach for preventing the corrosion of metals exposed to acidic conditions. Iron is coated with corrosion-resistant metals like Zn, Sn, and Cr to prevent corrosion. It has the greatest reach.

a method used in the food business to "tin-pack" goods. Tin is applied to the iron vessels to extend their lifespan. Physical and electrolytic methods can both be used to cover metal.

QUESTION .17

Briefly explain zinc coating and tin coating.

ANSWER

Zinc Coating

Galvanizing is the process of coating iron with a thin layer of zinc. A clean iron sheet is heated after being dipped in a zinc chloride bath to perform this operation. This iron sheet is then taken out, rolled in a bath of molten zinc metal, and ultimately air-cooled. Galvanizing has the benefit of keeping iron from corroding even after the coated surface is damaged by zinc

Tin Coating

The process entails dipping a fresh iron sheet in a molten tin and then running it through heated rollers. Cans of food and beverages employ these sheets. Only as long as the iron's barrier is unbroken does the tin serve as protection. when it is damaged and When iron is exposed to water and the elements, a galvanic cell forms, causing iron to corrode quickly

---

QUESTION .18

What is the process of electroplating silver?

ANSWER

Electroplating of Silver

By using electrolysis, one metal is deposited on top of another during electroplating.

Metals are improved in appearance and are protected against corrosion with this method.

The foundation of electroplating is the creation of an electrolytic cell with an anode formed of the cathode of the item on which metal is to be placed and the metal that is to be deposited. The electrolyte is dissolved in an aqueous solution of the corresponding metal salt.

Procedure

The item to be electroplated is cleansed with sand in this process, then washed with a caustic soda solution before being completely washed with water. The metal that is to be deposited, such as Cr or Ni, serves as the anode. The item serves as the cathode.

It will be electroplated to look like an iron sheet. A salt of the metal that is being deposited serves as the electrolyte in this system. Anode and cathode are suspended in the electrolytic tank, which is constructed of cement, glass, or wood. A battery holds these electrodes together. Metallic ions migrate to the cathode and discharge or deposit on the cathode when the

current is transferred, dissolving the metal from the anode in the solution (object). This discharge leaves a thin layer of metal on the object, which is then removed and cleaned.

Electroplating Of Silver

Setting up an electrolytic cell allows for the electroplating of silver.

The pure silver strip serves as the anode and is submerged in the silver nitrate solution.

The metallic item that needs to be coated, like a spoon, is the cathode. once the current as it moves through the cell, the anode disintegrates to create Ag+ ions, which move in the direction of the cathode. They are released and dropped on the object, such as a spoon, at the cathode. Chemical reactions can be visualized as

At Anode

$Ag \rightarrow Ag^{+} + 1e^{-}$

At Cathode

$Ag^{+} + 1e^{-} \rightarrow Ag$

QUESTION .19

Describe electroplating of chromium.

Or

How is chromium electroplated?

ANSWER

Electroplating Of Chromium

As with silver, chromium is electroplated using the same process.

The item to be electroplated is submerged in an electrolyte solution of aqueous chromium sulfate and a small amount of sulfuric acid. The thing that will be electroplated whereas the anode is constructed of antimonial lead, acts as the cathode. Cr3+ ions are produced by the electrolyte and ionization, which decrease and deposit at the cathode

At Anode

$4OH^- \rightarrow 2H_2O + O2 + 4e^-$

At Cathode

$Cr^{+3} + 3e \rightarrow Cr$

Because chromium does not attach well to the surface of the steel, it is typically first plated with nickel or copper for practical convenience before being chrome plated. Additionally, it enables moisture to seep through it, which strips metal. Copper or nickel offers stickiness, and after copper is put over it, chromium, the copper's sticky layer, lasts longer. This kind of electroplating provides the thing a dazzling silvery sheen and prevents corrosion.

QUESTION .20

Write a note on electrolytic refining of Copper

ANSWER

Electrolytic Refining Of Copper

The electrolytic process refines impure copper in the electrolytic cell.

The cathode is a plate of pure copper, whereas the anode is impure copper. An electrolyte is a solution of copper sulfate in water.

A reaction called oxidation occurs at the anode. Impure copper atoms lose electrons to the anode and dissolve as copper ions in solution.

$Cu \rightarrow Cu^{+2} + 2e$

The cathode is the site of the reduction reaction. The cathode is drawn to the copper ions in the solution. they become neutral, pick up electrons from the cathode, and deposit on the cathode there.

$Cu^{+2} + 2e \rightarrow Cu$

Purified copper atoms settle on the cathode as impure copper is consumed in the process.

Examples 1:

Oxidation state of chlorine in KCl

Solution

KCl, net charge = 0

Oxidation state of KCl = Oxidation state of potassium + oxidation state of chlorine = 0.

Oxidation number of potassium = +1

Oxidation states → +1 + x = 0: x = -1

Atoms in the species → K Cl

Oxidation number of chlorine in KCl = -1

Example 2:

Oxidation number of Manganese in permanganate ion MnO4–

Solution

The charge on permanganate ion is -1

Oxidation number of permanganate ion =Oxidation state of manganese + 4 oxidation state of oxygen = -1.

Oxidation number of oxygen = -2

Oxidation states → x + (4*-2) = -1: x = +7

Atoms in the species → Mn 4O

Oxidation state of manganese = +7
