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Give Short Answers of the Following Questions

1. What is Winkler method?

Ans. Winkler method is a technique used to measure the amount of dissolved oxygen (DO) in water samples. It is based on the idea that dissolved oxygen (DO) in the samples solution, has oxidizing properties and reacts with KI. The amount of iodine formed is measured by titrating with sodium thiosulfate using starch as an indicator which turns the solution dark blue, indicating the endpoint. The amount of iodine formed is proportional to the amount of dissolved oxygen in the solution.

2. Differentiate between dissolve and biological oxygen demand.

Ans. Dissolved oxygen (DO) is defined as the amount of oxygen dissolved in water. It is dissolved either from the atmosphere or released by plants during photosynthesis in water. Its concentration is between 5 to 8 mg/dm³. The minimum level of DO should be 5 mg/dm³. Below 5 mg/dm³, the water supposed to be polluted.

The capacity of organic matter in natural water to consume oxygen within a period of five days. The value of BOD is the amount of oxygen consumed as a result of biological oxidation of dissolved organic matter in the sample is called biological oxygen demand.

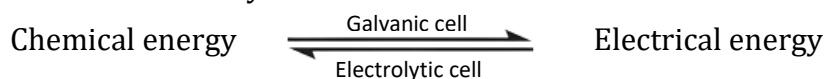
3. What is the principle of photovoltaic cell?

Ans. The central principle behind the operation of PV cells is the photovoltaic effect, which involves the generation of voltage and electric current in a material upon exposure to light. The photovoltaic effect involves the excitation of electrons from the valence band to the conduction band in semiconductors, facilitated by the energy from photons. This process is improved by the presence of a PN junction, which separates the electrons and holes, creating a voltage that drives an electric current.

4. Define electrochemistry.

Ans. Electrochemistry:

“The branch of chemistry concerned with the conversion of electrical energy into chemical energy in electrolytic cells and chemical energy into electrical energy in galvanic or voltaic cell is called electrochemistry.”



5. Differentiate between oxidizing agent and reducing agent.

Ans.	Oxidizing Agent	Reducing Agent
i.	A specie which oxidize a substance in a redox reaction is called an oxidizing agent or oxidant.	i. A specie which reduces a substance in a redox reaction is called as a reducing agent or reductant.
ii.	In the process, the oxidizing agent is reduced itself.	ii. In the process, the reducing agent is oxidized itself.
iii.	An oxidizing agent gain electrons and its oxidation number decreases in the reaction. e.g., KMnO ₄ /H ₂ SO ₄ , K ₂ Cr ₂ O ₇ /H ₂ SO ₄ , HNO ₃ , Cl ₂ , Br ₂ etc.	iii. The reducing agents lose electrons and its oxidation number is increased. e.g., FeSO ₄ , HI, SO ₂ , Na, K etc.

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6. Define oxidation and oxidation number. Give two examples.

Ans. Oxidation: There are three definitions

- (i) Loss of electron (ii) Loss of hydrogen (iii) Gain of oxygen

Oxidation Number/Oxidation State:

“The apparent charge on an atom of an element in a molecule or in an ion is called oxidation number or oxidation state.”

Example:

- (i) Oxidation number of Cr is +6 in $K_2Cr_2O_7$. (ii) Oxidation of p is +5 in Na_3PO_4 .

7. What is differences between oxidation and reduction process, give one example of each?

Ans.

Oxidation	Reduction
“Addition of oxygen or removal of hydrogen is called oxidation”	“Addition of hydrogen or removal of oxygen is called oxidation”
Loss of electron and hence increase in oxidation number is called oxidation.	Gain of electron and hence decrease in oxidation number is called reduction.
e.g., $Zn \rightarrow Zn^{+2} + 2e^-$	e.g., $Cu^{+2} + 2e^- \rightarrow Cu$
In electrochemical cells oxidation takes place at anode.	In electrochemical cells reduction takes place at cathode.

8. Write down the oxidation states of oxygen in peroxide and super oxides.

Ans. Oxidation States of Oxygen in Peroxide and Super Oxides:

In peroxides, oxygen has -1 oxidation state e.g. $H_2^{2(+1)}O^{2(-1)}$.

In superoxide it has -1/2 oxidation state e.g. $K^{+1}O^{(-1/2)}$.

9. Calculate the oxidation number of underlined elements: HNO_3 : CrO_3

Ans. HNO_3

Given Data: Oxidation Number of O = -2

Oxidation Number of H = +1

Required: Oxidation Number of N = x = ?

Solution:

(Oxidation No. of H) + (Oxidation No. of N) + 3 (Oxidation No. of O) = 0

$$(+1) + x + 3(-2) = 0$$

$$+1 + x - 6 = 0$$

$$x = +6 - 1 = +5$$

Oxidation Number of N is +5 in HNO_3 .

CrO_3

Given Data: Oxidation Number of O = -2

Required: Oxidation Number of Cr = x ?

Solution:

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(Oxidation No. of Cr) + 3(Oxidation No. of O) = 0

$$x + 3(-2) = 0$$

$$x - 6 = 0$$

$$x = +6$$

Oxidation Number of Cr is +6 in CrO_3 .

10. Define anode and cathode.

Ans. Anode:

An anode is an electrode of a polarized electrical device through which conventional current enters the device.

Cathode:

An electrode of the device through which conventional current leaves the device.

11. Define electrolytic cell. Give example.

Ans. Electrolytic cell:

In this cell, electrical energy is converted energy. Non-spontaneous redox reaction takes place.

Electrical energy is used to carry out reaction. Electrolysis takes place. Anode is positive and cathode is negative. e.g., Down's cell (to prepare pure Na-metal), Nelson's cell (to prepare caustic soda commercially) etc.

12. Define electrode potential and standard electrode potential.

Ans. Electrode Potential:

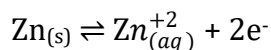
The tendency of a metal to form its ions or get deposited on the metal when it is dipped into solution of its own ions is called electrode potential.

Standard Electrode Potential

"The potential set up when an electrode is in contact with one molar solution of its own ions at 298 K is known as standard electrode potential or standard reduction potential of the element".

It is represented as E° .

For example, a rod of Zn will have an accumulation of negative charge owing to the net ionization of some of its atoms. This layer of negative charge will attract Zn^{+2} ions around the rod. In this way, an electrical double layer is formed around the Zn rod as shown in figure. The equilibrium state is represented as



13. What is Standard Hydrogen Electrode (SHE)? Explain with the help of diagram.

Ans. Standard Hydrogen Electrode (SHE):

A standard hydrogen electrode consists of a piece of platinum foil, coated with a layer of finely divided platinum black (to give it a large surface area) led with a platinum wire. It is immersed in 1 M HCl solution. Pure H_2 gas at one atmospheric pressure is continuously bubbled into 1 M HCl solution.

The platinum acts as an electrical conductor and also facilitates the attainment of equilibrium between the gas and its ions in solution. The electrode potential (oxidation or reduction) of this standard hydrogen electrode is arbitrarily taken as zero at all temperatures.

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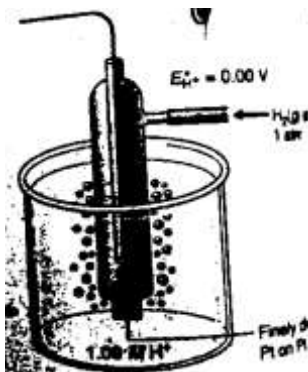
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$$E^{\circ} = 0.00 \text{ V}$$



14. Write down the importance of standard Hydrogen Electrode (SHE).

Ans. Importance of Standard Hydrogen Electrode (SHE):

- SHE is used as a standard electrode.
- Standard Hydrogen Electrode is used as a reference electrode and has a potential of 0.0 volts.
- SHE is used to find the reduction potential of different electrodes.
- SHE serves as a medium for cell potential calculations.

15. Differentiate between Galvanic cell and Electrolytic cell.

Ans.

Electrolytic Cell	Voltaic/Galvanic Cell
• In this cell, electrical energy is converted into chemical energy.	• In this cell, the chemical energy is converted into electrical energy.
• Non-spontaneous redox reaction takes place.	• Spontaneous redox reaction takes place.
• Electrical energy is used to carry out reaction.	• Electrical energy is generated by a chemical reaction.
• Electrolysis takes place.	• No electrolysis takes place.
• Anode is positive and cathode is negative.	• Anode is negative and cathode is positive.
• Example: Down's cell (to prepare pure Na-metal), Nelson's cell (to prepare caustic soda commercially) etc.	• Example: Zn-Cu cell (Daniel cell), lead storage battery, silver oxide cell etc.

16. How and why electrical double layer is formed?

Ans. An electrical double layer forms at the interface between a solid electrode and an electrolyte solution. It consists of two layers: a layer of charge on the electrode surface and a layer of opposite charge in the solution. This structure arises due to electrostatic interactions between the electrode and the ions in the solution.

17. Why electrode potential of Cu is called reduction potential?

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Ans. The electrode potential of Cu is called a reduction potential because it represents the tendency of Cu^{2+} ions to gain electrons (reduction) and form solid copper. A positive value indicates a strong tendency to gain electrons.

18. What are the advantages of salt bridge in a galvanic cell?

Ans. A salt bridge:

- Maintains electrical neutrality by allowing ion flow between half-cells.
- Prevents the solutions from mixing while completing the circuit.
- Enables continuous electron flow, allowing the cell to produce electricity.

19. How can we predict the feasibility of a chemical reaction using the cell voltage?

Ans. The feasibility of a reaction can be predicted by calculating the standard cell potential (E_{cell}°). A positive E_{cell}° indicate a spontaneous reaction, while a negative E_{cell}° suggests a non-spontaneous reaction.

20. During electrolysis of aqueous NaCl, why Na is not liberated at the cathode?

Ans. Sodium (Na) is not liberated at the cathode during the electrolysis of aqueous NaCl because water has a lower reduction potential than Na^{+} ions. Therefore, water is reduced to hydrogen gas instead of Na^{+} being reduced to Na metal.

21. Calculate the Ox. No. of chromium (Cr) in the following compounds:

i) CrCl_3

ii) $\text{Cr}_2(\text{SO}_4)_3$

iii) $\text{Cr}_2\text{O}_7^{2-}$

Ans. Oxidation Number of Chromium (Cr)

(i) CrCl_3 :

Given Data: Oxidation number of Cl = -1

Required: Oxidation number of Cr = x ?

Solution:

Oxidation No. of Cr + 3(Oxidation No. of Cl) = 0

$$x + 3(-1) = 0$$

$$x - 3 = 0$$

$$x = +3$$

Oxidation number of Cr is +3 in CrCl_3 .

(ii) $\text{Cr}_2(\text{SO}_4)_3$:

Given Data: Oxidation number of O = -2

Oxidation number of S = +6

Required: Oxidation number of Cr = x = ?

Solution:

2(Oxidation No. of Cr) + 3(Oxidation No. of S)

$$+ 12(\text{Oxidation No of O}) = 0$$

$$2x + 3(+6) + 12(-2) = 0$$

$$2x = 24 - 18 = +6$$

$$x = +3$$

Oxidation number of Cr is +3 in $\text{Cr}_2(\text{SO}_4)_3$.

(iii) $\text{Cr}_2\text{O}_7^{2-}$:

Given Data: Oxidation number of O = -2

Required: Oxidation number of Cr = x ?

Solution:

2(Oxidation No. of Cr) + 7(Oxidation No. of O) = -2

$$2x + 7(-2) = -2$$

$$2x - 14 = -2$$

$$2x = +14 - 2 = +12$$

$$x = +6$$

Oxidation number of Cr is +6 in $\text{Cr}_2\text{O}_7^{2-}$.

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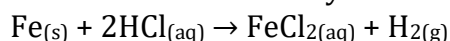
22. The order of decreasing reactivity of metals based on their position is $K > Mg > Zn > Fe > Cu$. Write balanced chemical equations for the reactions that would occur (if any) when:

- Copper is added to a solution of magnesium sulphate.**
- Iron is added to a dilute solution of hydrochloric acid.**

Ans. Copper is added to a solution of magnesium sulphate:

No reaction occurs because copper is less reactive than magnesium.

Iron is added to a dilute solution of hydrochloric acid:



23. Explain why some metals higher in the activity series can displace hydrogen from acids, while other lower in the series cannot.

Ans. Metals higher in the activity series (e.g., Zn, Mg) are more reactive and have a greater tendency to lose electrons (oxidize) compared to hydrogen. When reacted with acids, these metals displace H^+ ions (reducing them to H_2 gas) because their reduction potential is lower than hydrogen's.

Metals lower in the series (e.g., Cu, Ag) are less reactive than hydrogen. Their reduction potential is higher meaning they cannot supply electrons to reduce H^+ ions, so they do not displace hydrogen from acid.

In short, electrochemical reactivity (position in the series) determines whether a metal can reduce H^+ to H_2 .

24. Calculate the number of Faradays required to deposit 108 g of Ag^{1+} , 63.5 g of Cu^{2+} and 27 g of Al^{3+} .

Ans. Number of Faradays = Mass of substance (g)/equivalent mass of substance

The equivalent mass (E) is calculated as:

$$E = \frac{\text{Molar mass (M)}}{\text{Valency (n)}}$$

1. For 108 g of Ag^+ (Silver)

Molar mass (Ag) = 108 g/mol

Valency = 1

Equivalent mass (E) = $\frac{108}{1} = 108$ g/equivalent

Number of Faradays = $\frac{108g}{108g/equivalent} = 1F$

2. For 63.5 g of Cu^{2+} (Copper)

Molar mass (Cu) = 63.5 g/mol

Valency = 2

Equivalent mass (E) = $\frac{63.5g}{2} = 31.75$ g/equivalent

Number of Faradays = $\frac{63.5g}{31.75g/equivalent} = 2F$

3. For 27 g of Al^{3+} (Aluminum)

Molar mass (Al) = 27 g/mol

Valency = 3

Equivalent mass (E) = $\frac{27}{3} = 9$ g/equivalent

Number of Faradays = $\frac{27g}{9g/equivalent} = 3F$

Final Answer:

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- Ag^+ (108 g): 1 Faradays
- Cu^{2+} (63.5 g): 2 Faradays
- Al^{3+} (27 g): 3 Faradays

(Each Faradays = 96,485 C/mol, but here we only calculate the number of Faradays.)

25. In an electrolysis experiment, a current of 0.500 A was passed through a solution of AgNO_3 for 30.0 minutes. The mass of silver deposited on the cathode was found to be 0.503 g. Given that the molar mass of silver is $107.87 \text{ g mol}^{-1}$ and the charge on a silver ion is +1. Calculate the value of Avogadro's number (N_A) from this data.

Ans. Given Data:

Current (I)	= 0.500 A
Time (t)	= 30.0 minutes = 1800 seconds
Mass of Ag deposited (m)	= 0.503 g
Molar mass of Ag (M)	= $107.87 \text{ g mol}^{-1}$
Change on Ag^+ ion (n)	= 1

Step 1: Calculate the total charge passed (Q)

$$Q = I \times t = 0.500 \text{ A} \times 1800 \text{ s} = 900 \text{ C}$$

Step 2: Determine moles of Ag deposited

$$\text{Moles of Ag} = \frac{m}{M} = \frac{0.503 \text{ g}}{107.87 \text{ g mol}^{-1}} = 4.663 \times 10^{-3} \text{ mol}$$

Step 3: Related charge to moles of electrons (Faradays law)

Since $\text{Ag}^+ + e^- \rightarrow \text{Ag}$, 1 mole of Ag requires 1 mole of electrons (1 Faraday = 96,485 C).

Total charge required for $4.663 \times 10^{-3} \text{ mol Ag}$.

$$Q_{\text{theory}} = \text{Moles of Ag} \times F = 4.663 \times 10^{-3} \times 96,485 \text{ C}$$

But the actual charge used (Q) = 900 C, so:

$$F = \frac{Q}{\text{Moles of Ag}} = \frac{900}{4.633 \times 10^{-3}} = 1.930 \times 10^5 \text{ C mol}^{-1}$$

Step 4: Calculate Avogadro's number (N_A)

Since $F = N_A \times e$ (where $e = 1.602 \times 10^{-19} \text{ C}$):

$$N_A = \frac{F}{e} = \frac{1.930 \times 10^5}{1.602 \times 10^{-19}} = 1.2 \times 10^{24}$$

Final Answer: $N_A = 1.20 \times 10^{24} \text{ mol}^{-1}$

(Note: Due to error in statement of question the answer comes 1.20×10^{24} molecules, however it should be 6.02×10^{23}).

26. A cell is set up with a standard nickel electrode ($\text{Ni}_{(\text{aq})}^{2+} + 2e^- \rightleftharpoons \text{Ni}_{(\text{s})}$ $E^\circ = -0.25 \text{ V}$) and a standard cobalt electrode ($\text{Co}_{(\text{aq})}^{2+} + 2e^- \rightleftharpoons \text{Co}_{(\text{s})}$ $E^\circ = -0.28 \text{ V}$).

- Identify which metal will be the anode and which will be the cathode. Justify your answer.
- Write the balanced overall cell reaction.
- Calculate the standard cell potential (E°_{cell}).

Ans. Solution:

i) **Identifying Anode and Cathode**

• **Standard Reduction Potentials (E°):**

- $\text{Ni}^{2+} + 2e^- \rightleftharpoons \text{Ni}_{(\text{s})}$ $E^\circ = -0.25 \text{ V}$
- $\text{Co}^{2+} + 2e^- \rightleftharpoons \text{Co}_{(\text{s})}$ $E^\circ = -0.28 \text{ V}$

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- **Comparison:**
 - Nickel has a less negative (higher) reduction potential than cobalt.
 - **Cathode (Reduction):** The electrode with the higher reduction potential (Ni^{2+}/Ni).
 - **Anode (Oxidation):** The electrode with the lower reduction potential (Co^{2+}/Co), where oxidation occurs.

Reason:

Since, $E_{(\text{Ni}^{2+}/\text{Ni})}^{\circ} > E_{(\text{Co}^{2+}/\text{Co})}^{\circ}$, Ni^{2+} gains electrons (reduction) at the cathode, while Co less electrons (oxidation) at the anode.

ii) Balanced Overall Cell Reaction

- **Oxidation at Anode (Co):** $\text{Co}_{(\text{s})} \rightarrow \text{Co}^{2+} + 2\text{e}^{-}$
- **Reduction at Cathode (Ni):** $\text{Ni}_{(\text{aq})}^{2+} + 2\text{e}^{-} \rightleftharpoons \text{Ni}_{(\text{s})}$
- **Overall Reaction:** $\text{Ni}_{(\text{aq})}^{2+} + 2\text{e}^{-} \rightarrow \text{Ni}_{(\text{s})}$

iii) Standard Cell Potential (E° cell)

- $E_{\text{cell}}^{\circ} = E_{\text{cathod}}^{\circ} = -E_{\text{anode}}^{\circ}$
- $E_{\text{cell}}^{\circ} = (-0.25 \text{ V}) - (-0.28 \text{ V}) = +0.03 \text{ V}$
- The positive E_{cell}° confirms the reaction is spontaneous.
- Cobalt (anode) dissolves, while nickel (cathode) deposits.

Final Answer:

- i) **Anode:** $\text{Co}_{(\text{s})}$ (oxidation) ii) $\text{Ni}_{(\text{aq})}^{2+} + \text{Co}_{(\text{s})} \rightarrow \text{Ni}_{(\text{s})} + \text{Co}^{2+}$ iii) $E_{\text{cell}}^{\circ} = +0.03 \text{ V}$
Cathode: $\text{Ni}_{(\text{aq})}^{2+}$ (reduction)