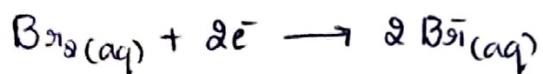


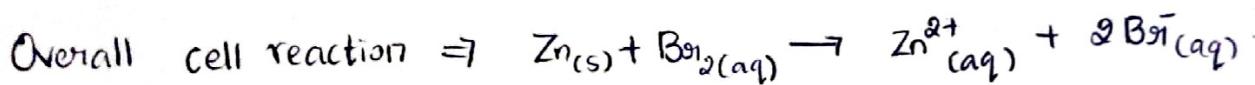
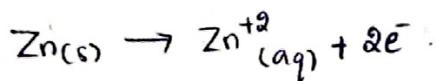
1. This because in every electrochemical cell involves a redox reaction takes place. We know that in redox reaction both oxidation and reduction takes place simultaneously so we can conclude that in an electrochemical cell reduction reaction is accompanied by oxidation reaction.

Example:

Reduction half-reaction :



Oxidation half-reaction:



3. Generally we use standard potential of 2 cells for comparing their output. This is because the emf of cell is independent of parameters like concentration, temperature etc. So, the emf or cell potential of any 2 cells makes no sense unless we define the conditions of temperature and concentration.

4. Since we usually measure the standard reduction potential of a halfcell as its ability to gain electron, a negative standard reduction potential implies that a given compound has a greater tendency to lose electrons than gaining of electron and since we know that oxidation is the process of loss of

electrons from a system, we could conclude that a large negative value at standard reduction potential for a cell indicates that the half cell is likely to undergo oxidation.

5. Since Pb is given,

from the data given Co ion is on L.H.S and Cu is on R.H.S.

Here,

Co is anode, Cu is cathode

Thus

$$E^\circ_{\text{Cu}^{+2}/\text{Cu}} - E^\circ_{\text{Co}^{+2}/\text{Co}} = 0.614 \rightarrow ①$$

and similarly

$$E^\circ_{\text{Cu}^{+2}/\text{Cu}} - E^\circ_{\text{Fe}^{+2}/\text{Fe}} = 0.777 \rightarrow ②$$

Now,

$$\underline{\underline{② - ① \text{ gives}}}$$

$$E^\circ_{\text{Cu}^{+2}/\text{Cu}} - E^\circ_{\text{Fe}^{+2}/\text{Fe}} = 0.777$$

$$(+) E^\circ_{\text{Cu}^{+2}/\text{Cu}} - E^\circ_{\text{Co}^{+2}/\text{Co}} = 0.614$$

$$\underline{\underline{E^\circ_{\text{Co}^{+2}/\text{Co}} - E^\circ_{\text{Fe}^{+2}/\text{Fe}} = 0.163}}$$

The total potential of the given cell is 0.163 V.

Given, $[\text{Cu}^{+2}] = 4.8 \times 10^{-3} \text{ M}$

$$[\text{Zn}^{+2}] = 0.4 \text{ M}$$



(It is a voltaic cell given)

\therefore From nearest equation,

$$\begin{aligned} E_{\text{cell}} &= E_{\text{cell}}^{\circ} - \frac{0.0592}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} \\ &= 1.1 - 0.0296 \log \left(\frac{4 \times 10^3}{48} \right) \\ &= 1.1 - 0.0296 (1.92) \\ &= 1.1 - 0.0568 \\ &= 1.0432 \text{ V} \end{aligned}$$

2.

Galvanic cell

* Spontaneous redox reactions

Convert the chemical energy to
an electric energy.

* Electrical energy is generated
by redox reactions.

* The cathode is the positive
electrode and anode is the
negative electrode.

Voltaic cell

* Non spontaneous redox reactions

Convert the electrical energy to
chemical energy.

* Electrical energy brings about
the chemical reaction with the
help of external source.

* The anode is the positive
electrode and cathode is the
negative electrode.

- * The process of oxidation takes place at the anode and the reduction process occurs at the cathode.
- * Here the oxidation process occurs at the cathode while the reduction process takes place at anode.
- * Half cells are set up in different containers and are connected through salt bridges.
- * Electrode are kept in the same container in a molten or solution electrolyte.
- * Application lies in batteries.
- * Application lies in purifying Copper and electroplating.