

Hydrogen Half Cell

To establish standard conditions, we place an electrode of the metal to be tested in a 1-molar solution of its ions (Fig.41).

A semi permeable membrane divides the cell. For the other half of the cell a platinum electrode is placed in a 1-molar solution of hydrogen ions.

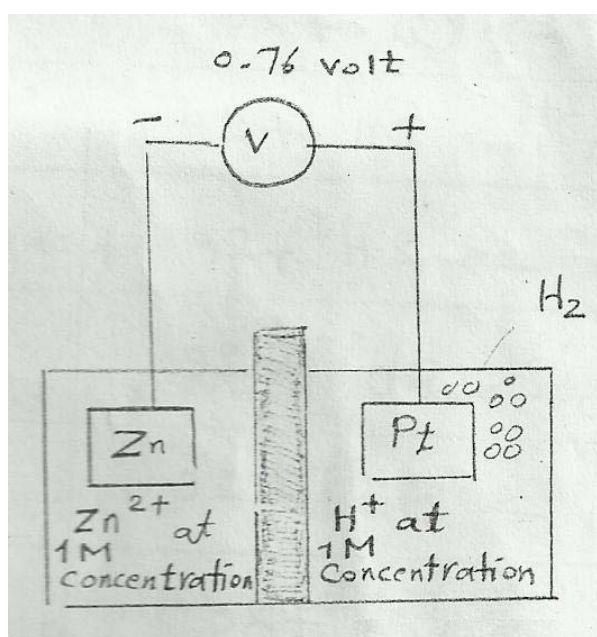
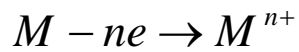


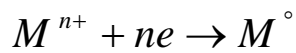
Fig. (41) Diagram of a half cell for measurement of potential

A stream of hydrogen gas is bubbled around the platinum electrode. Therefore, the platinum serves to adsorb the gas and does not take part in the reaction.

If the metal is more reactive than hydrogen, we have the following half cell reaction.



On the other hand, if the metal is nobler (less reactive) than hydrogen, we have :



If the metal is not surrounded by a 1-molar concentration of its ions, the driving voltage is correspondingly less. This effect is derived from Nernst equation, which can be simplified:

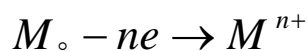
$$E = E_0 + \frac{0.0592}{n} \text{LogCoin}$$

Where:

E : new electromotive force (emf)

E_0 : Standard emf of Table.1

n : number of electrons transferred as



Coin = molar concentration of ions.

Ex: What would be the new half-cell voltages for zinc and copper if the ion concentrations were changed to 0.1 molar?

Answer:

$$E = E_{\circ} + \frac{0.0592}{n} \log \text{Coin}$$

$$\begin{aligned} \text{Zinc: } E &= -0.763 + \frac{0.0592}{2} \log 0.1 \\ &= -0.763 - 0.0296 = -0.793 \text{ volt} \end{aligned}$$

$$\text{Copper: } E = 0.337 + \frac{0.0592}{2} \log 0.1$$

$$= +0.337 - 0.0296 = +0.307 \text{ volt}$$

The result for zinc shows a greater tendency to corrode.

Corrosion Rates

We have introduced the concept of solution potential, and we find the emf (electro motive force) or (potential) indicates the tendency to corrode as well as the anode and cathode, but it does not directly govern the corrosion rate. The corrosion rate depends on Faraday's law:

Weight of metal dissolving (g) = KIT

$$K = \frac{\text{at.wt.of metal}(M)}{\text{no of electron transferred} \times 96,500 \text{ amp. sec}} \quad \text{Where}$$

I = current (amp) and t = time (sec)

In corrosion we are usually interested in the loss in weight from a given area per unit time.

Therefore, dividing both sides of the above equation by the area gives us the corrosion rate. In other words, the rate of weight loss from a given area is proportional to the current density.

$$W = \frac{M}{nf} It$$

W : weight loss (g)

$\left(\frac{\text{g}}{\text{mole}}\right)M$: atomic weight

I : current (amp)

t : time (sec.)

n : no of valence electrons

F : Faraday constant 96500 amp.sec

Example: It is required to electro-plate one side ($2 \times 2 \text{ cm}^2$) carbon steel sample with silver using a current of 10 amp. Determine the time require to make silver plate of 0.2 cm.

($\frac{\text{g}}{\text{mole}}$) Using (M) atomic weight of silver = 107.89

(f) density of silver = 10.5 g/cm^3

Answer

$$W = \frac{M}{nF} It$$

$$\frac{W}{t} = \frac{M}{nF} I$$

$$\frac{W}{t} = \frac{107.87}{(1)(96500)} (10) \quad (10)$$

$$\frac{W}{t} = 0.011 \frac{\text{g}}{\text{sec}}$$

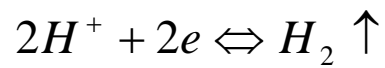
$$\frac{V \cdot f}{t} = 0.011$$

$$\therefore t = \frac{V \cdot f}{0.011} = \frac{(2 \times 2 \times 0.2)(10.5)}{0.011}$$

$$\therefore t = 763.63 \text{ sec}$$

Ex: Calculate the corroded potential for H₂ electrode at PH = 1,2, and 3

Answer



$$E = E_0 + \frac{0.059}{n} \log[H^+]$$

$$PH = -\log[H^+]$$

When PH = 1

$$E = 0 + \frac{0.059}{1} (-1) = -0.0592 \text{ volt}$$

When PH = 2

$$E = 0 + \frac{0.059}{1} (-2) = -0.1184 \text{ volt}$$

When PH = 3

$$E = 0 + \frac{0.059}{1} (-3) = -0.1776 \text{ volt}$$