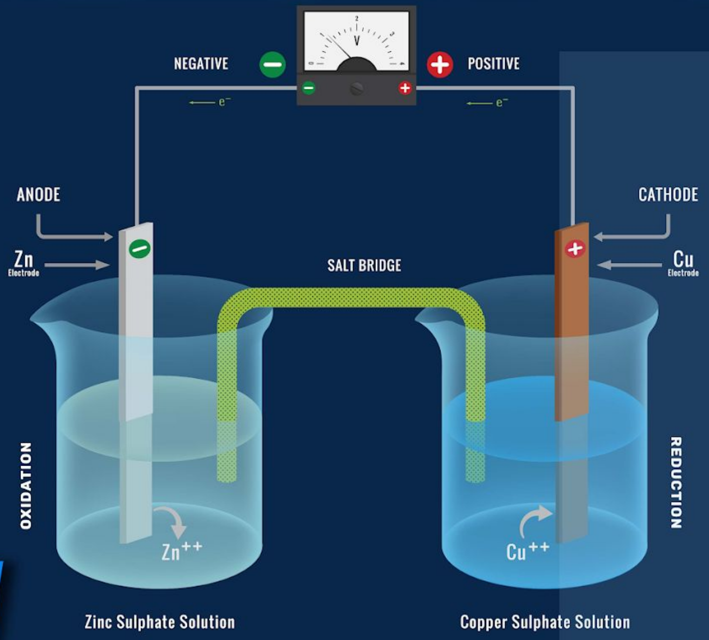
ELECTROCHEMISTRY - L6



CHEMISTRY

ANOOP SIR

NEET SCHOLAR COURSE

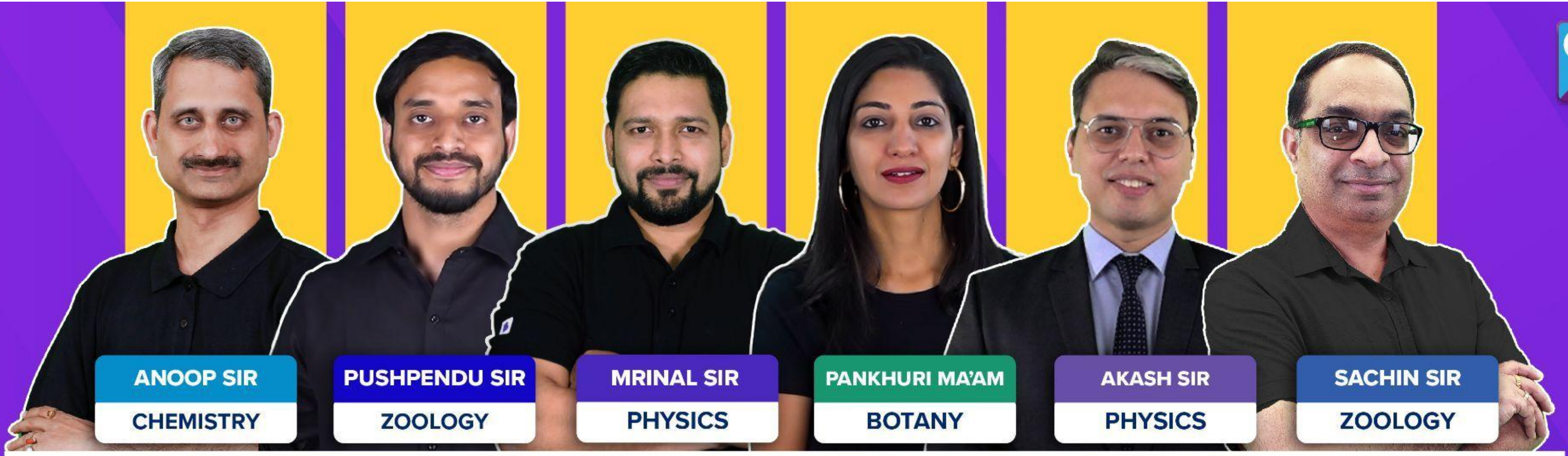
FOR CLASS 12th

VALID
till May
2023

SCHEDULE
10 LIVE sessions/Week,
2 classes/day

GET BYJU'S
SD Card, Books
& Mentor Support

FORM LINK GIVEN IN DESCRIPTION



ANOOP SIR	CHEMISTRY
PUSHPENDU SIR	ZOOLOGY
MRINAL SIR	PHYSICS
PANKHURI MA'AM	BOTANY
AKASH SIR	PHYSICS
SACHIN SIR	ZOOLOGY

MISSION MBBS

CHEMISTRY SUPER 30

BIO की रण NEETi

PHY की रण NEETi

MON - SAT | 12 PM - 8 PM

Aakash Live Webinars

6 Months NEET Strategy till May 2023

20th November, 2022

12:30 pm

Dr. Sachin Kapur

Biology Expert - NEET

ANTHE

AAKASH NATIONAL TALENT HUNT EXAM

Your Gateway To Success

For Class VII to XII

Current Students & Passouts

Aakash
+BYJU'S

NEET
STUDENTS' SURVEY

LINK IN DESCRIPTION

https://t.me/neetaakashdigital

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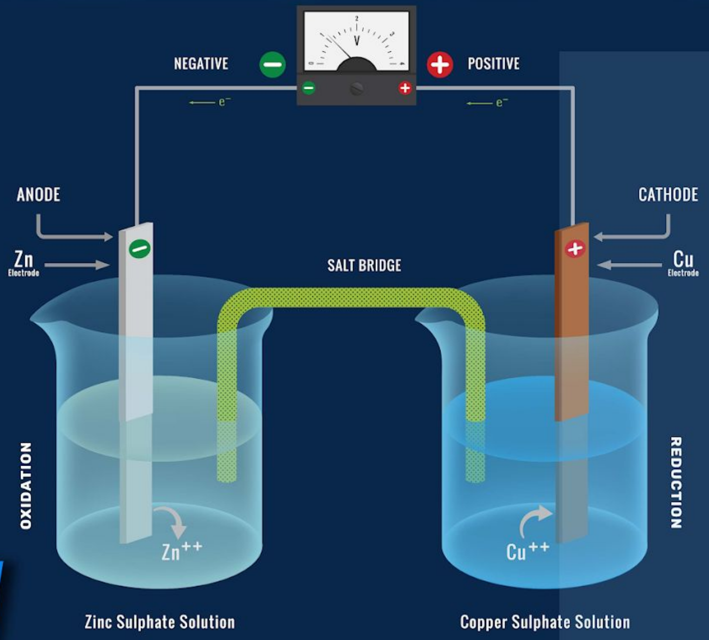

CHEMISTRY

ANOOP SIR

Nernst equation.

$E = E^\circ - \frac{0.0591}{n} \log Q$ at 298K

Hydrogen electrode

0

① $H^+_{(aq)} + \underline{\underline{e^-}} \longrightarrow \frac{1}{2} H_{2(g)}$

$n = 1$ $Q = \frac{(p_{H_2})^{1/2}}{[H^+]}$

$$E_{H^+/H_2} = -\frac{0.0591}{1} \log \frac{(p_{H_2})^{1/2}}{[H^+]}$$

② $2H^+ + 2e^- \longrightarrow H_{2(g)}$

$n = 2$ $Q = \frac{p_{H_2}}{[H^+]^2}$

$$E_{H^+/H_2} = -\frac{0.0591}{2} \log \frac{p_{H_2}}{[H^+]^2}$$

Q. Calculate the potential of hydrogen electrode under a constant pressure of 1 Bar H_2 and dipped in

a) 10^{-2} M HCl

$p_{H_2} = 1$

$[H^+] = 10^{-2}$ M

$$E_{H^+/H_2} = -0.0591 \log \frac{1}{[H^+]}$$

$$E_{H^+/H_2} = -0.0591 \log \frac{1}{10^{-2}}$$

$$= -0.0591 \log 10^{2}$$

$$= -0.0591 \times 2 \text{ V}$$

$$= -0.1182 \text{ V}$$

b) a solution of pH 4

$$E_{H^+/H_2} = -0.0591 \log \frac{(P_{H_2})^{1/2}}{[H^+]}$$

if $P_{H_2} = 1$

$$E_{H^+/H_2} = -0.0591 \log \frac{1}{[H^+]}$$

$$\log \frac{1}{[H^+]} = -\log[H^+]$$

$$= pH$$

$$E_{H^+/H_2} = -0.0591\, pH \rightarrow P_{H_2} = 1\,\text{Bar}$$

$$T = 298\,K$$

$$E_{H^+/H_2} = -0.0591 \times 4 = -0.2364\,V$$

c) Pure water :

pH of pure water at 298K = 7

$E_{H^+/H_2} = -0.0591 \times 7$ V

d) 0.1M NaOH $\longrightarrow$ pH = 13

$E_{H^+/H_2} = -0.0591 \times 13$

Q What should be the partial pressure of hydrogen so that a hydrogen electrode dipped in pure water has potential 0 V

$$E_{H^+/H_2} = -0.0591 \log \frac{(P_{H_2})^{1/2}}{[H^+]} = \frac{-0.0591}{2} \log \frac{P_{H_2}}{[H^+]^2}$$

$$0 = \boxed{-0.0591} \log \frac{(P_{H_2})^{1/2}}{10^{-7}}$$

Antilog(0) = 1

$$\log \frac{(P_{H_2})^{1/2}}{10^{-7}} = 0$$

$$\frac{(P_{H_2})^{1/2}}{10^{-7}} = 1$$

$$(P_{H_2})^{1/2} = 10^{-7}$$

$$P_{H_2} = (10^{-7})^2$$

$$P_{H_2} = 10^{-14} \text{ Bar.}$$

$Pt \mid H_2(g) \mid CH_3COOH(aq) \parallel HCl(aq) \mid H_2(g) \mid Pt$ 298K

1 Bar 0.1 M product 0.1 M Reactant. 1 Bar

Ecell will be

a) zero as both acids have same concentration.

b) positive

c) negative.

d) cannot be determined.

concentration cell in which H^+ concentration is different in two half cells.

CH_3COOH is weak acid so $[H^+]$ is less.

To find ΔG from Ecell.

$$\Delta G^\circ = -nFE^\circ$$

ΔG° → standard Gibb's energy change
n = number of electron transferred
E° → standard cell potential.

ΔG° is an extrinsic quantity, so the change in balancing of the equation changes the value of n and so changes ΔG°

$\frac{4}{3}Al + O_2 \longrightarrow \frac{2}{3}Al_2O_3$ $n = 4$

$2Al + \frac{3}{2}O_2 \longrightarrow Al_2O_3$ $n = 6$

$4Al + 3O_2 \longrightarrow 2Al_2O_3$ $n = 12$

Never change the balancing of equations.

$Fe^{2+} + 2e^- \longrightarrow Fe$ $\quad E^\circ_{Fe^{2+}/Fe} = -0.44V$

$Fe^{3+} + e^- \longrightarrow Fe^{2+}$ $\quad E^\circ_{Fe^{3+}/Fe^{2+}} = +0.77V$

$Fe^{3+} + 3e^- \longrightarrow Fe$ $\quad E^\circ_{Fe^{3+}/Fe} = ?$

Electrode potentials are intrinsic quantities so cannot be directly added or subtracted.

It becomes possible if the final equation does not have any remaining electron.

If both equations are reduction or oxidation, we must convert electrode potentials to ΔG°.

$Fe^{2+} + 2e^- \rightarrow Fe \quad E^\circ = -0.44\,V \qquad \Delta G_1^\circ = -2F \times (-0.44\,V)$

$Fe^{3+} + e^- \rightarrow Fe^{2+} \quad E^\circ = +0.77\,V \qquad \Delta G_2^\circ = -1 \times F \times (+0.77\,V)$

$Fe^{3+} + 3e^- \rightarrow Fe \quad E^\circ = ? \qquad \Delta G_3^\circ = -3 \times F \times E_3^\circ$

$$\Delta G_3^\circ = \Delta G_1^\circ + \Delta G_2^\circ$$

$$-3FE_3^\circ = -2F(-0.44) + (-1F(0.77))$$

$$E_3^\circ = \frac{2E_1^\circ + E_2^\circ}{3} = \frac{-0.88 + 0.77}{3}$$

$$= -\frac{0.11}{3} \simeq -0.036\,V$$

Relation between $\Delta_r G$ and E_{cell}

Relation between $\Delta_r G$ and E_{cell}

The **maximum non-expansion (electrical) work** that a system (the cell) can do is equal to the decrease in Gibbs free energy.

$$W_{e,\,max} = |\Delta G|$$

ΔG identified with the Gibbs energy of the cell reaction, **$\Delta_r G$**

Relation between $\Delta_r G$ and E_{cell}

Maximum electrical work obtained from a cell $= |\Delta_r G|$

$$\Delta_r G = -q \times V = -nFE$$

where F is Faraday’s constant

- n = Number of moles of electrons
- E = Electrode potential or cell EMF
- $\Delta_r G$ = Change in Gibbs energy for half-cell or cell

Relation between $\Delta_r G$ and E_{cell}

1 F = Charge present on 1 mole of electrons

= 96,485 C/mol

≈ 96,500 C/mol

Note

By knowing the $\Delta_r G$ at a specified composition, the **cell emf** at that composition can be stated.

Spontaneous cell reaction

$\Delta_r G < 0$

$E_{cell} > 0$

Non-spontaneous cell reaction

$\Delta_r G > 0$

$E_{cell} < 0$

Any redox reaction would occur spontaneously, if:

[✓ a $\Delta_r G^0$ is negative]

b $\Delta_r G^0$ is positive

c E^0 is negative

d All of these

The Gibbs energy for the decomposition of Al_2O_3 at 500°C is as follows: $\frac{2}{3}Al_2O_3 \rightarrow \frac{4}{3}Al + O_2$; $\Delta_r G = +966\ kJ\ mol^{-1}$. The potential difference needed for electrolytic reduction of Al_2O_3 at 500°C is at least:

$\Delta G = -4FE^{\circ}$

a 4.5 V

b 3.0 V

c 2.5 V

d 5.0 V

Given: $\frac{2}{3}Al_2O_3 \longrightarrow Al + O_2; \Delta_rG = +966\ \text{kJ mol}^{-1}$ (at 500^0C)

To find: Minimum potential difference needed for electrolytic reduction of Al_2O_3 at 500^0C

For the reduction of silver ions with copper metal, the standard cell potential was found to be $\underline{+0.46\ V}$ at $\underline{25°C}$. The value of standard Gibbs energy, $\underline{\Delta G^0}$, will be:

($F = 96500\ C\ mol^{-1}$)

AIPMT 2010

a	-89.0 kJ
b	-89.0 J
c	-44.5 kJ
d	-98.0 kJ

“Stay Positive, Work Hard, Make It Happen”

THANK YOU