ULTIMATE KCET

CRASH COURSE 2026

Chemistry | Lecture - 01

Chemical kinetics

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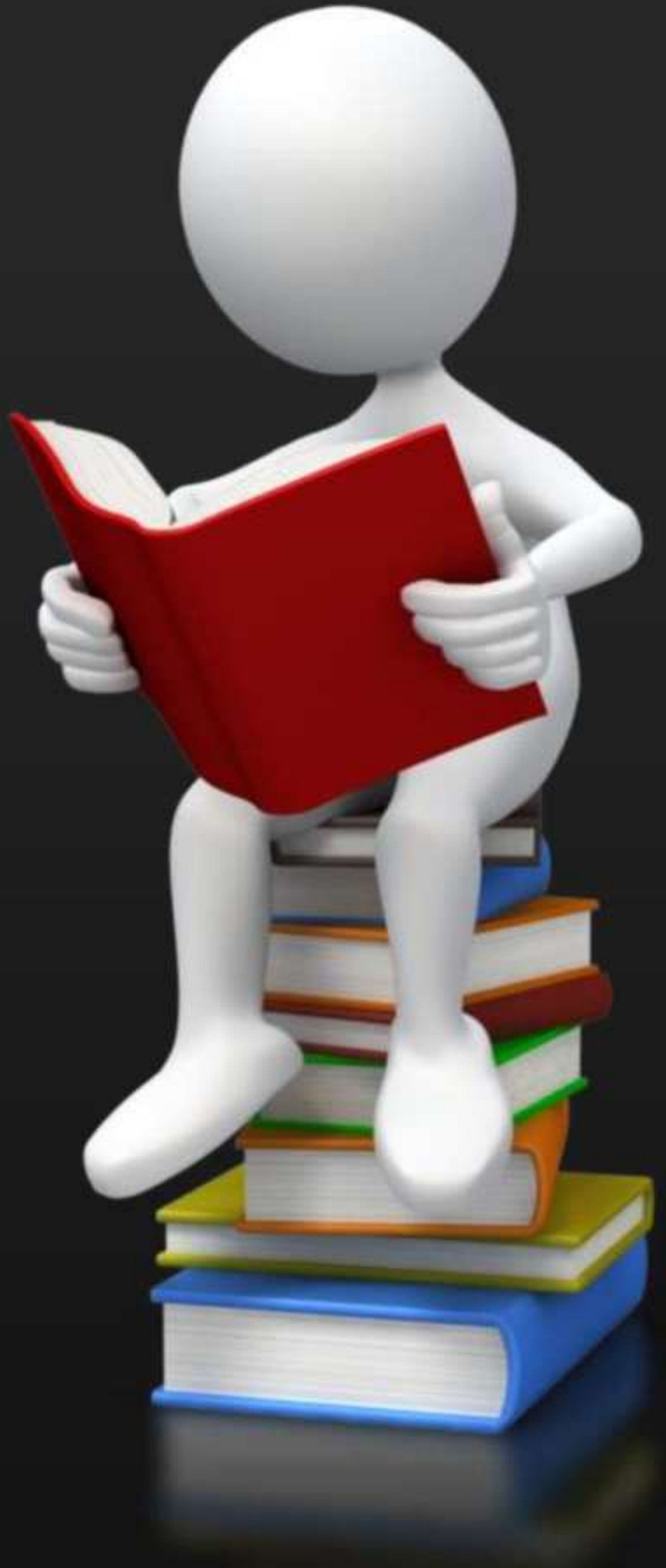
Recap *of previous lecture*

1. Electrochemistry - I

Topics *to be covered*

1. Chemical kinetics
2. Batteries – Equations and Mcq



Batteries

- ☐ Any battery (actually it may have one or more than one cell connected in series) or cell that we use as a source of electrical energy is basically a galvanic cell where the chemical energy of the redox reaction is converted into electrical energy.
- ☐ However, for a battery to be of practical use it should be reasonably light, compact and its voltage should not vary appreciably during its use.

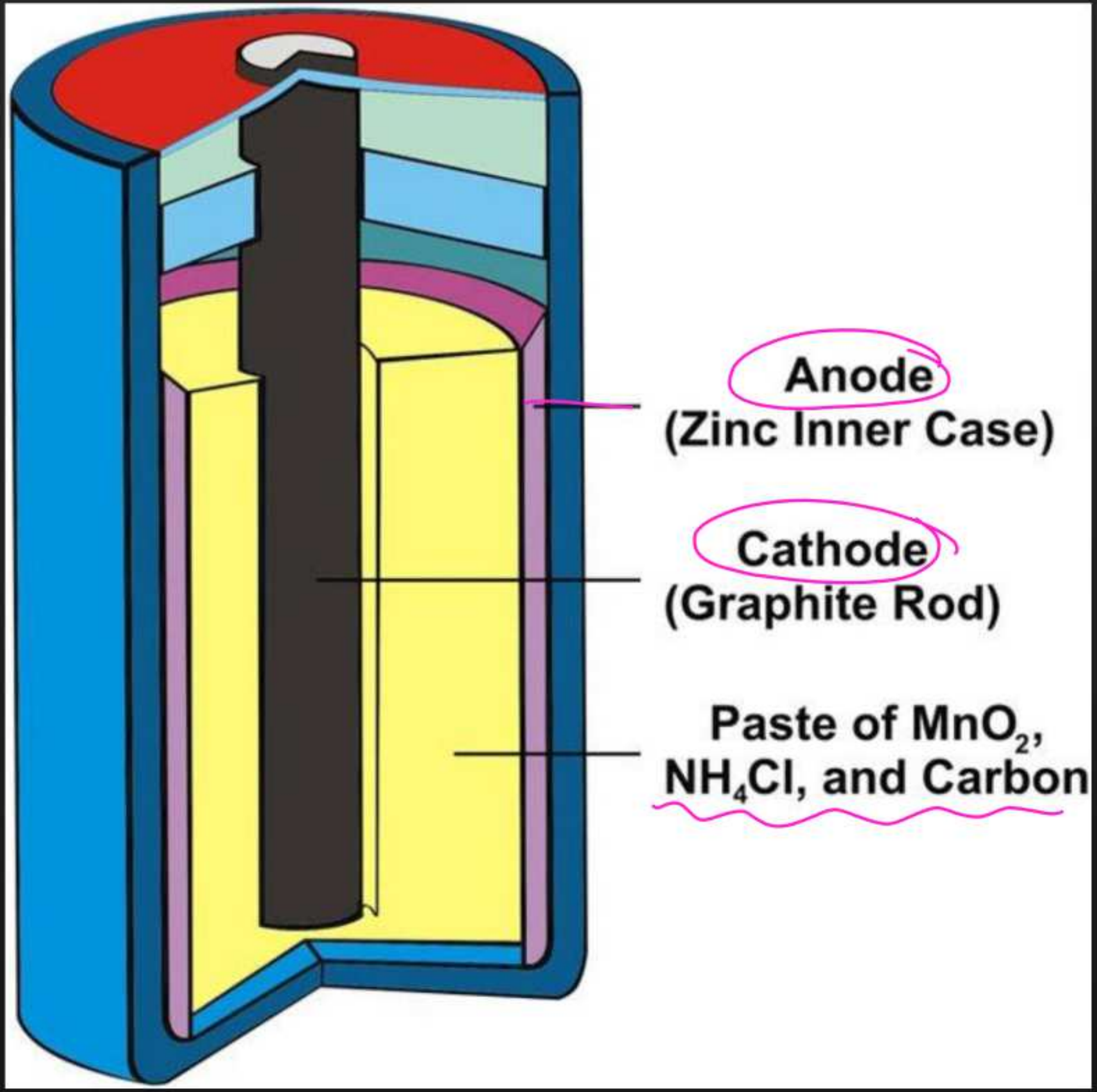
- ☐ There are mainly two types of batteries-
  - ➢ Primary Batteries ✓
  - ➢ Secondary Batteries ✓

Primary Batteries

Primary Batteries

In the primary batteries, the reaction occurs only once and after use over a period of time battery becomes dead and cannot be reused again.

Example - Dry cell (known as Leclanché cell) which is used commonly in our transistors and clocks and Mercury cell

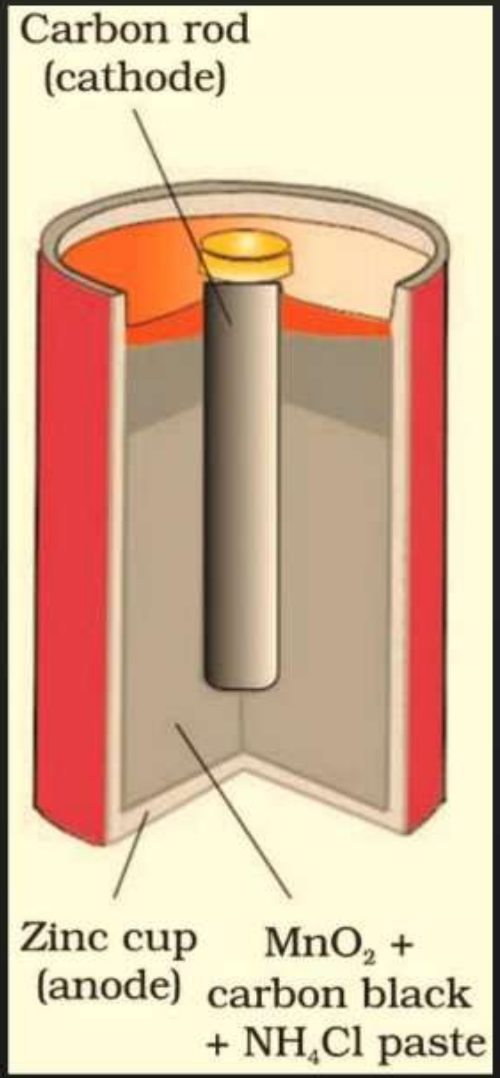


✓✓ $ZnCl_2 + NH_3 \longrightarrow [Zn(NH_3)_2]^{2+}$

Dry cell (known as Leclanché cell)

$Zn \longrightarrow Zn^{2+} + 2e^{\ominus}$

- The cell consists of a zinc container that also acts as anode
- and the **cathode is a carbon (graphite) rod** surrounded by powdered manganese dioxide and carbon.
- The space between the electrodes is filled by a moist paste of ammonium chloride (NH_4Cl) and zinc chloride ($ZnCl_2$). $+MnO_2$



A commercial dry cell consists of a graphite (carbon) cathode in a zinc container; the latter acts as the anode.

- ☐ **Anode** : $Zn(s) \rightarrow Zn^{2+} + 2e^-$
- ☐ **Cathode** : $\overset{+4}{MnO_2} + NH_4^+ + e^- \rightarrow \overset{+3}{MnO(OH)} + NH_3$ ✓
- ☐ In the reaction at cathode, manganese is reduced from the +4 oxidation state to the +3 state.
- ☐ Ammonia produced in the reaction forms a complex with Zn^{2+} to give $[Zn(NH_3)_4]^{2+}$.
- ☐ The cell has a potential of nearly 1.5 V.

drycell Mercurycell Pb storage Ni-cd.

Cathode
C electrode
anode
Zn
Electrolyte → $MnO_2 + NH_4Cl + ZnCl_2$.

H_2-O_2-fulecell

Mercury cell

Mercury cell, suitable for low current devices like hearing aids, watches, etc.

Anode - zinc – mercury amalgam

Cathode - a paste of HgO and carbon as the cathode.

The electrolyte is a paste of KOH and ZnO.

The electrode reactions for the cell are given below:

Anode :

$$\overset{0}{Zn}(Hg) + \cancel{2OH^-} \xrightarrow{\text{Oxidation}} \boxed{\overset{+2}{ZnO}(s)} + \cancel{H_2O} + \cancel{2e^-}$$

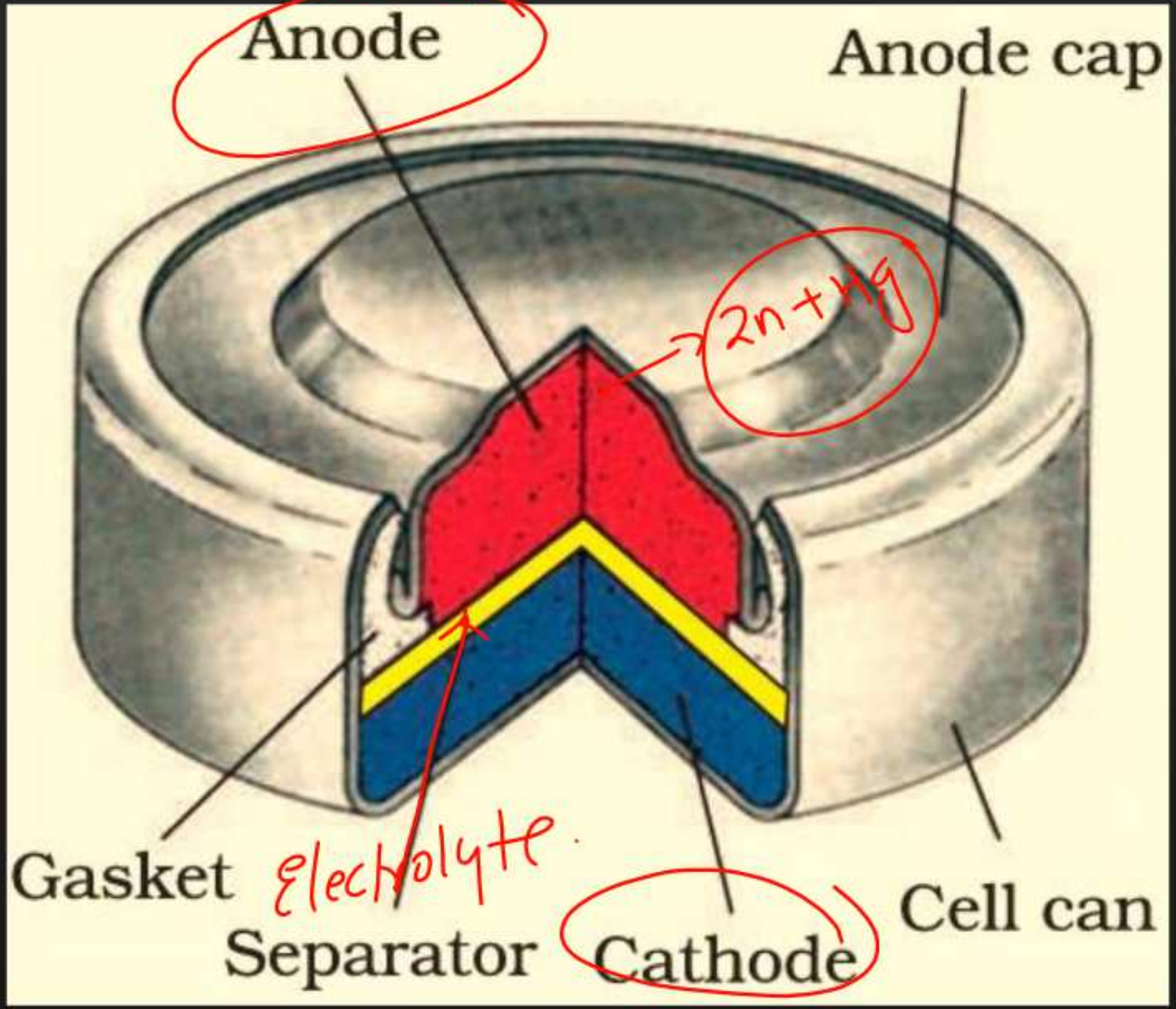
Cathode :

$$\overset{+2}{HgO} + \cancel{H_2O} + \cancel{2e^-} \rightarrow \overset{0}{Hg}(l) + \cancel{2OH^-}$$

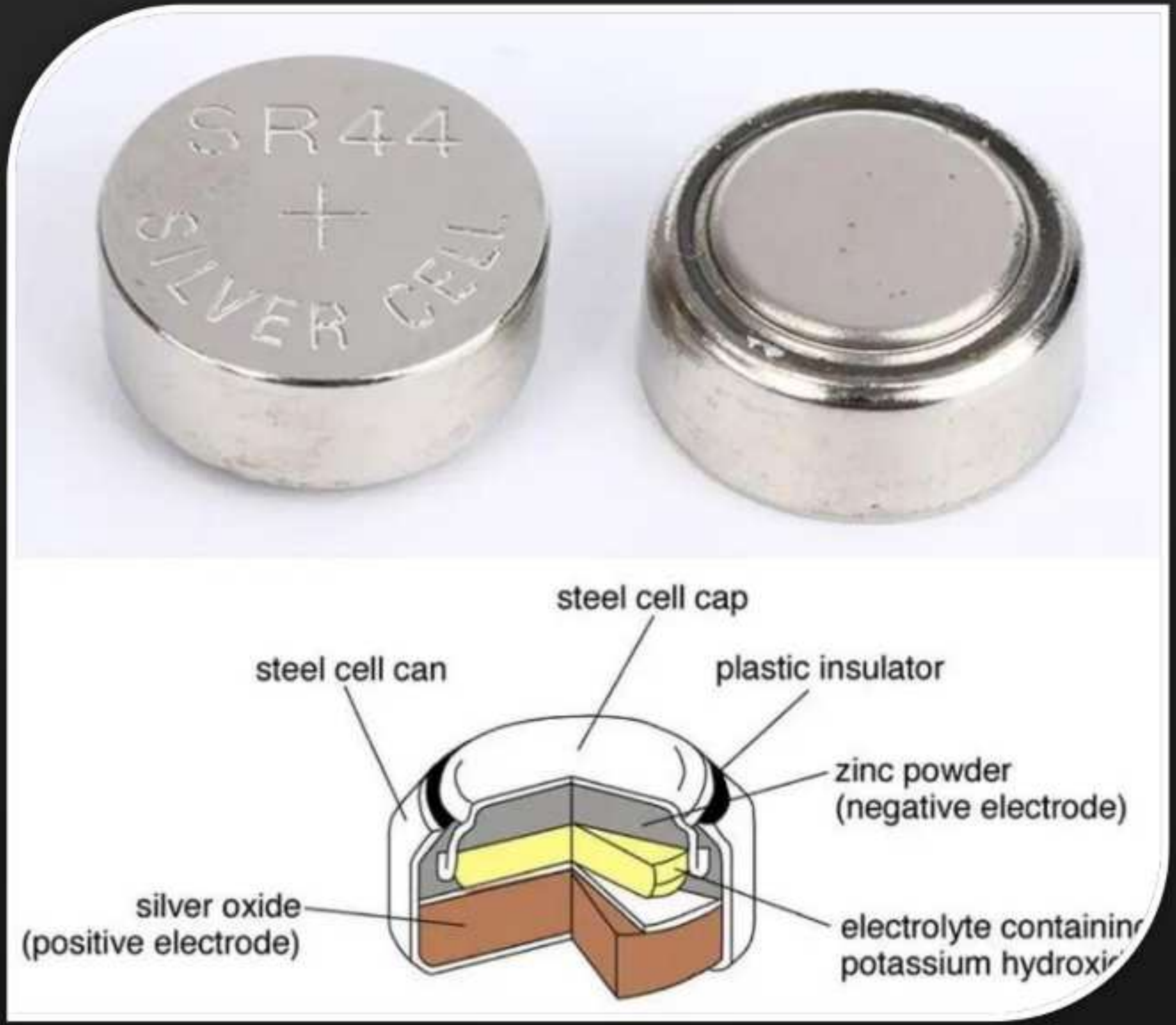
The overall Cell reaction :

$$\boxed{Zn(Hg) + HgO(s) \rightarrow ZnO(s) + Hg(l)}$$

The cell potential is approximately 1.35 V



Commonly used mercury cell. The reducing agent is zinc and the oxidising agent is mercury (II) oxide.



Secondary Batteries

- A secondary cell after use can be recharged by passing current through it in the opposite direction so that it can be used again.
- A good secondary cell can undergo a large number of discharging and charging cycles.
- lead storage battery
- Nickel-cadmium cell

lead storage battery

It consists of a lead anode and a grid of lead packed with lead dioxide (PbO_2) as cathode.

A 38% solution of Sulphuric acid is used as an electrolyte.

The cell reactions when the battery is in use are given below:

Anode : $Pb(s) + SO_4^{2-}(aq) \xrightarrow[\text{recharging/charging}]{\text{discharging}} PbSO_4(s) + 2e^-$ $(Pb \rightarrow Pb^{2+})$

Cathode : $\overset{+4}{Pb}O_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$

Pb^{2+}

i.e., overall cell reaction consisting of cathode and anode reactions is:

$$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$$

On charging the battery the reaction is reversed and $PbSO_4$ (s) on anode and cathode is converted into Pb and PbO_2, respectively.

Anode:

$$Pb + SO_4^{2-} \underset{\text{recharge}}{\overset{\text{discharge}}{\rightleftharpoons}} PbSO_4 + 2e^{\ominus}$$

Cathode

$$PbO_2 + 2e^{\ominus} + SO_4^{2-} + 4H^{\oplus} \underset{\text{recharge}}{\overset{\text{discharge}}{\rightleftharpoons}} PbSO_4 + 2H_2O$$

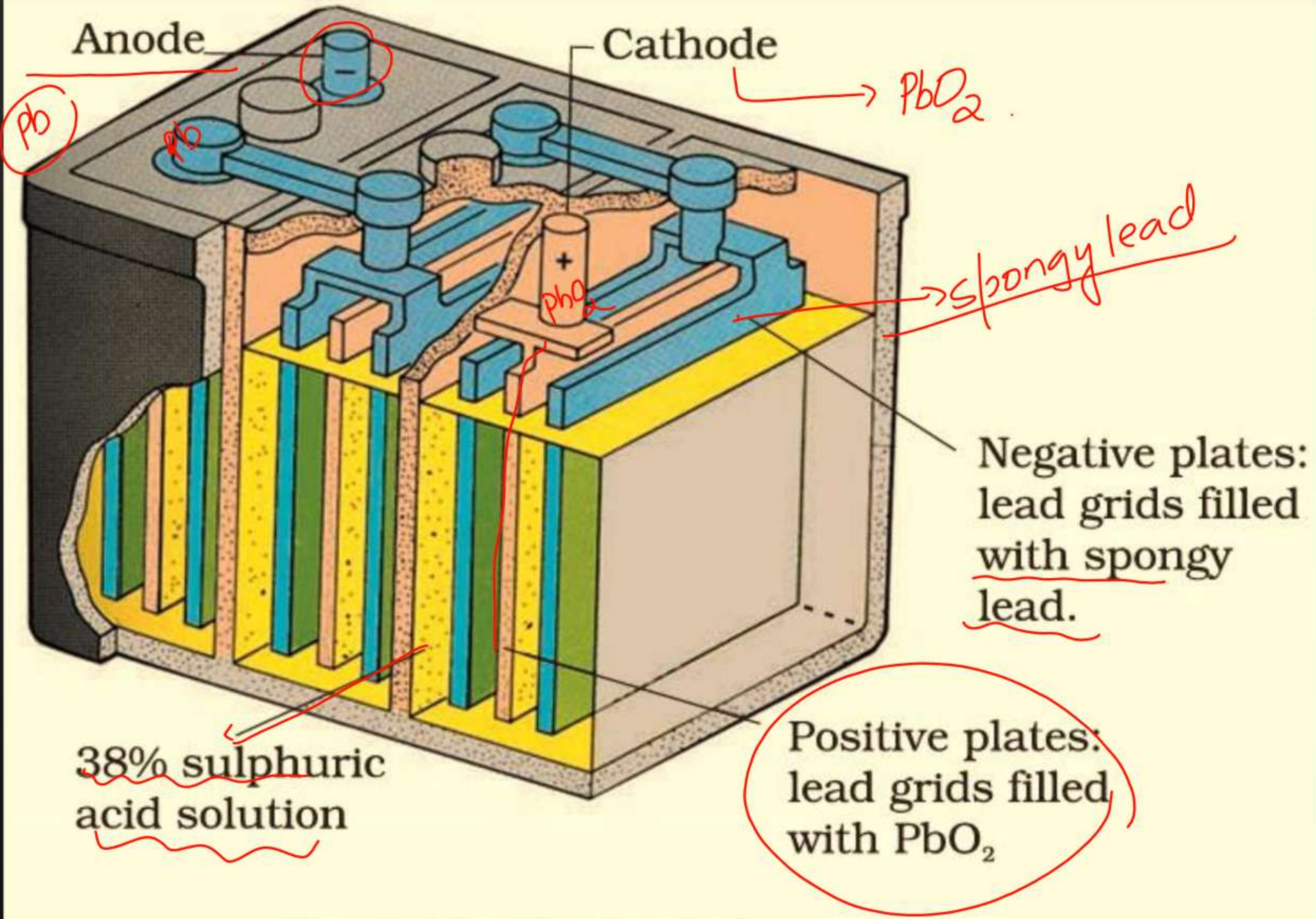


Fig. 3.10: The Lead storage battery.

Nickel-cadmium cell

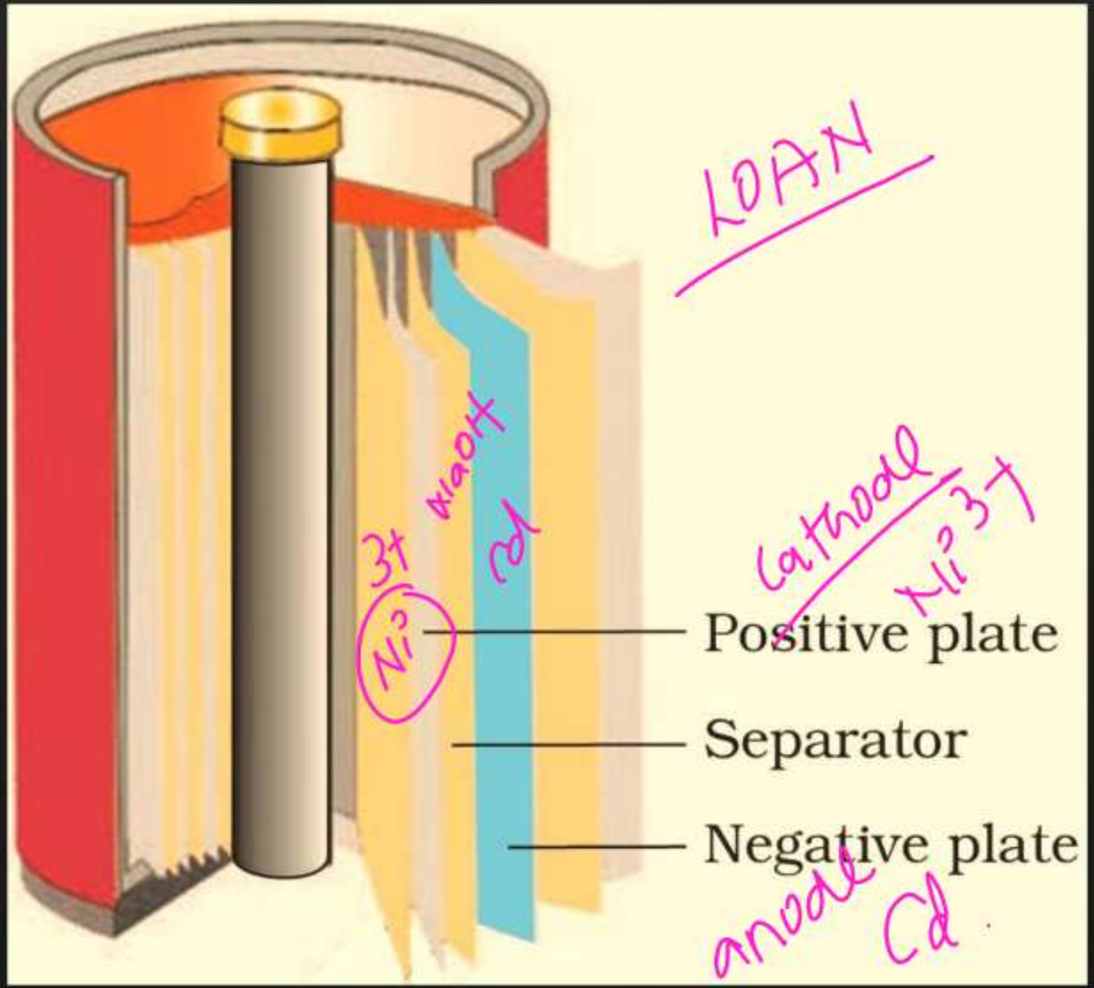
- ❑ This has longer life than the lead storage cell but more expensive to manufacture.
- ❑ The overall reaction during discharge is:

$$\overset{0}{Cd}_{(s)} + 2Ni(OH)_{3\,(s)} \rightarrow \overset{+2}{Cd}O_{(s)} + 2Ni(OH)_{2\,(s)} + H_2O_{(l)}$$

anode → Cadmium

Cathode → $Ni^{+3} + e^{\ominus} \longrightarrow Ni^{+2}$

$$\boxed{Ni(OH)_3 \longrightarrow Ni(OH)_2}$$



A rechargeable nickel-cadmium cell in a jelly roll arrangement and separated by a layer soaked in moist sodium or potassium hydroxide.

- Galvanic cells that are designed to convert the energy of combustion of fuels like hydrogen, methane, methanol, etc. directly into electrical energy are called fuel cell.

- **H_2-O_2 fuel cell**

- The cell was used for providing electrical power in the Apollo space programme
- Fuel cells produce electricity with an efficiency of about 70 % compared to thermal plants whose efficiency is about 40%.

H_2-O_2 fuel cell

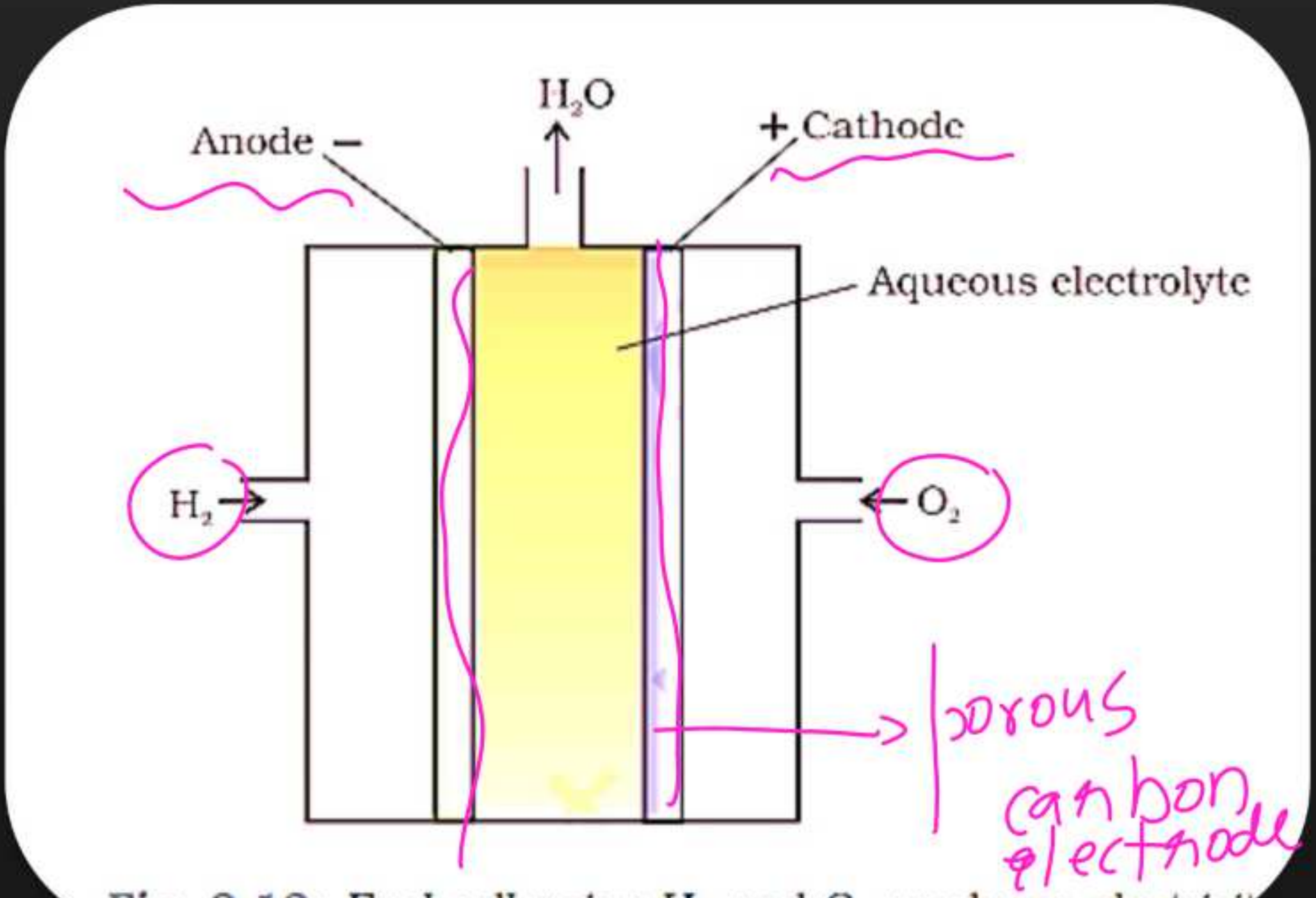
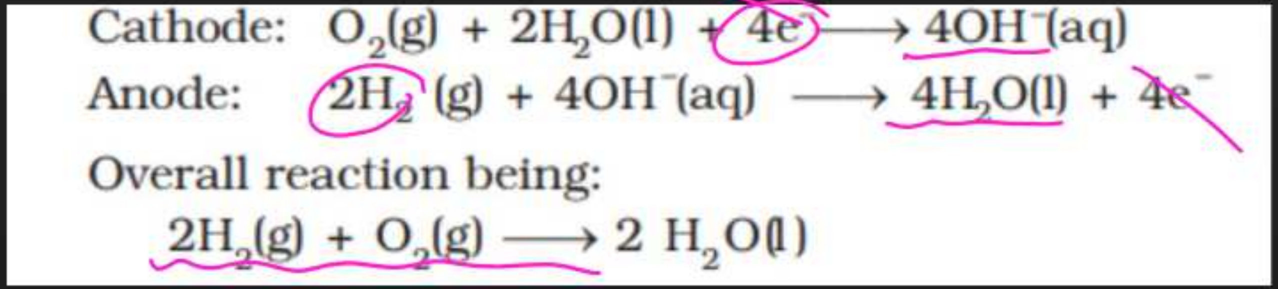


Fig. 3.12: Fuel cell using H_2 and O_2 produces electricit



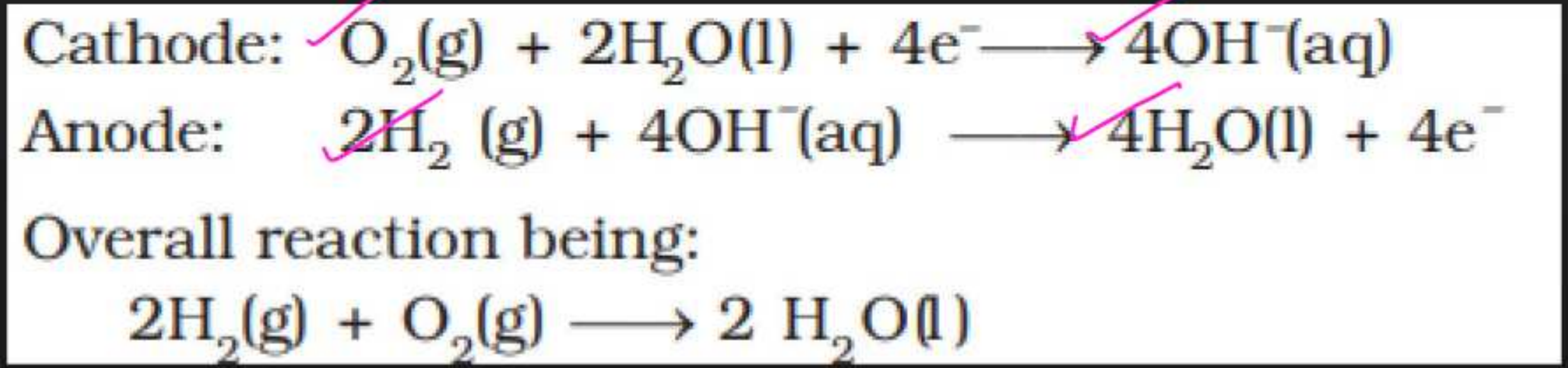
Cathode: $O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$

Anode: $2H_2(g) + 4OH^-(aq) \longrightarrow 4H_2O(l) + \cancel{4e^-}$

Overall reaction being:

$2H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l)$

- In the cell, hydrogen and oxygen are bubbled through porous carbon electrodes into concentrated aqueous sodium hydroxide solution.
- Catalysts like finely divided platinum or palladium metal are incorporated into the electrodes for increasing the rate of electrode reactions.



Cathode: $O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$

Anode: $2H_2(g) + 4OH^-(aq) \longrightarrow 4H_2O(l) + 4e^-$

Overall reaction being:

$$2H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l)$$

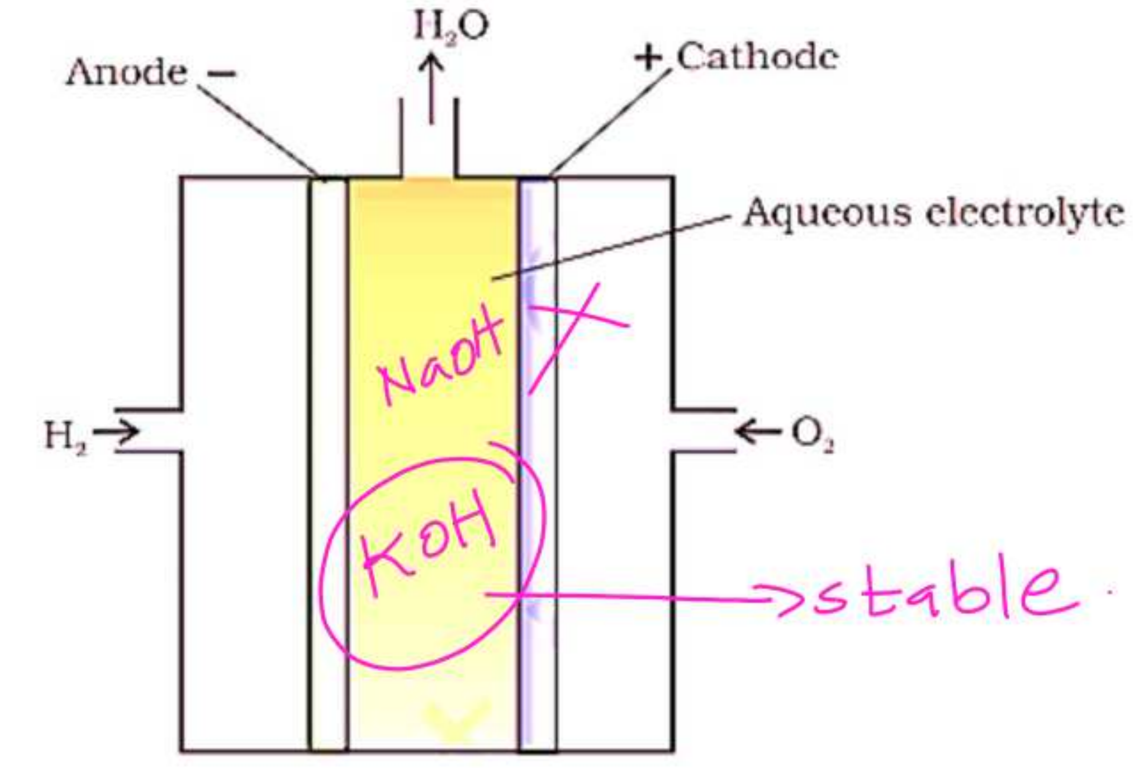


Fig. 3.12: Fuel cell using H_2 and O_2 produces electrici

Corrosion

→ $Fe_2O_3 \cdot xH_2O$

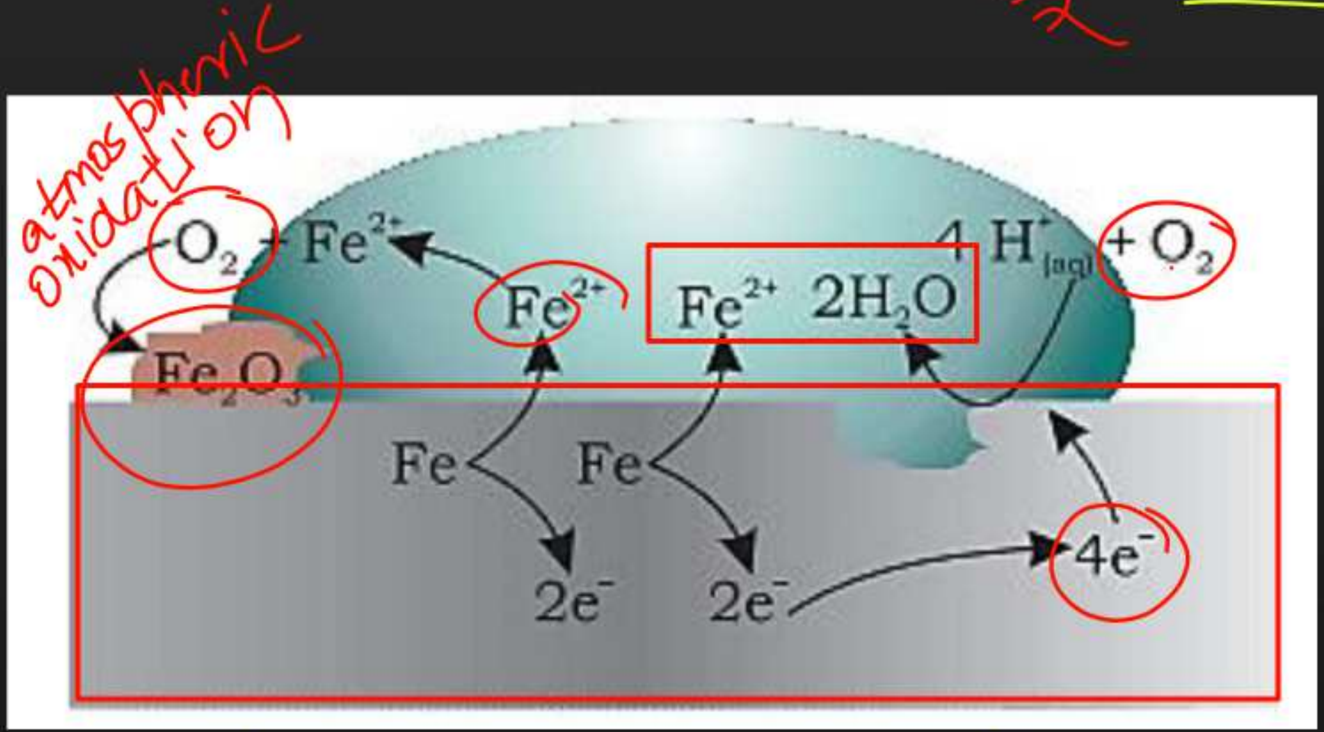
→ Ag_2S.

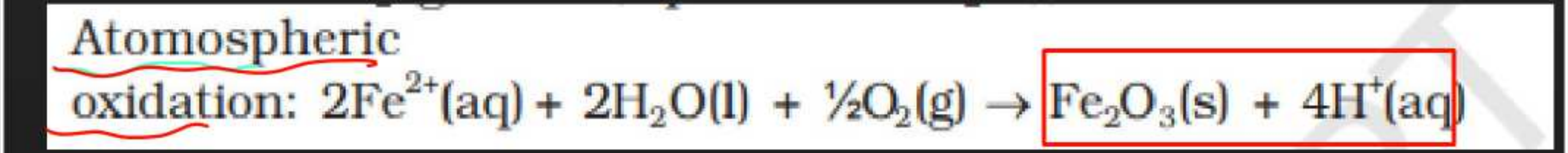
$Cu(OH)_2 \cdot CuCO_3$

- The rusting of iron, tarnishing of silver, development of green coating on copper and bronze are some of the examples of corrosion.
- It causes enormous damage to buildings, bridges, ships and to all objects made of metals especially that of iron
- In corrosion, a metal is oxidised by loss of electrons to oxygen and formation of oxides
- Corrosion of iron (commonly known as rusting) occurs in presence of water and air.

Corrosion- Corrosion of iron in atmosphere

$CO_2 \longrightarrow H_2CO_3 \longrightarrow H^{\oplus}$




Atomospheric oxidation: $2Fe^{2+}(aq) + 2H_2O(l) + \frac{1}{2}O_2(g) \rightarrow \boxed{Fe_2O_3(s) + 4H^+(aq)}$

O

$O \rightarrow 2e^{\ominus}$ H_2O

$H^{\oplus}$ O^{2-}

Anode → $2Fe \longrightarrow 2Fe^{2+} + \cancel{4e^{\ominus}}$

Cathode → $O_2 + \cancel{4e^{\ominus}} + 4H^{\oplus} \longrightarrow 2H_2O$

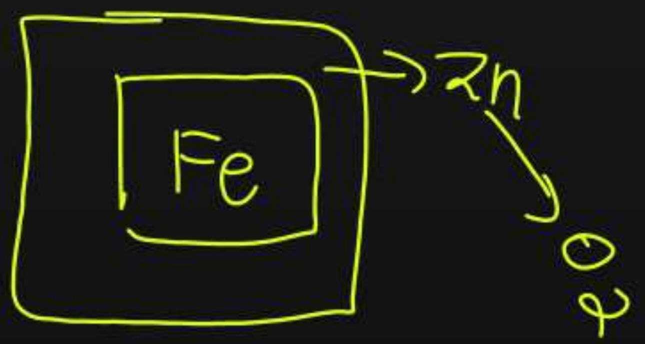
Cell reaction → $2Fe + O_2 + 4H^{\oplus} \longrightarrow \boxed{2Fe^{2+} + 2H_2O}$

- At particular spot of an object made of iron, oxidation takes place and that spot behaves as anode.
- Anodic reaction:
- Electrons released at anodic spot move through the metal and go to another spot on the metal and reduce oxygen in presence of H^+. This spot behaves as cathode with the reaction.
- Cathodic reaction:
- Over all cell reaction:

Prevention of Corrosion

K
Na
Ca
Mg
Al
Zn
Fe
Sn
Pb

- One of the simplest methods of preventing corrosion is to prevent the surface of the metallic object to come in contact with atmosphere.
- This can be done by covering the surface with paint or by some chemicals (e.g. bisphenol).
- Another simple method is to cover the surface by other metals (Sn, Zn, etc.) that are inert or react to save the object.
- An electrochemical method is to provide a sacrificial electrode of another metal (like Mg, Zn, etc.) which corrodes itself but saves the object.



product of electrolysis

NaCl
molten

$Na^{\oplus} \longrightarrow$ cation $\longrightarrow$ cathode deposition
$Cl^{\ominus} \longrightarrow$ anion $\longrightarrow$ anode, gas liberation

aq. NaCl
cathode $\rightarrow H_2$
anode $\rightarrow Cl_2$
NaOH $p^H < 7$

aq. $AgNO_3$
product of electrolysis

$Ag^{\oplus}$, $H^{\oplus}$ | $NO_3^{\ominus}$, $OH^{\ominus}$ → anode → $O_2 \uparrow$

cathode Ag

$CuSO_4$ aq.

$H^{\oplus}$ Cu^{2+}	SO_4^{2-} $OH^{\ominus}$

cathode → Cu deposite

anode
$OH^{\ominus} \longrightarrow O_2 \uparrow$

conc H_2SO_4	dil H_2SO_4
H_2 cathode	H_2 cathode
$S_2O_8^{2-}$	$OH^{\ominus} \rightarrow O_2$

anode

$S_2O_8^{2-}$ → anode
peroxodisulphate ion

K
Na
Ca
Mg
Al
Zn
Sn
Pb
H
Hg
Cu
Ag

$NO_3^{\ominus}$
$x-6=-1$
$x=+5$

Question

Equivalent conductance $\longrightarrow \Lambda_{eq} \longrightarrow$ conductance due to all the ions produced from 1 gram equivalent electrolyte.

The molar conductance of NaCl, HCl and C_2H_5COONa at infinite dilution are 126.45, 426.16 and 91~~$\text{ohm}^{-1}\,\text{cm}^2$~~. The equivalent conductance of C_2H_5COOH is

$S\,cm^2\,eq^{-1}$

(A) $201.28\,\text{ohm}^{-1}\,\text{cm}^2$

(B) $390.71\,\text{ohm}^{-1}\,\text{cm}^2$ ✓

(C) $698.28\,\text{ohm}^{-1}\,\text{cm}^2$

(D) $540.48\,\text{olm}^{-1}\,\text{cm}^2$

$$C_2H_5COO\boxed{H + Na}Cl \longrightarrow C_2H_5COONa + HCl$$

$$\Lambda + 126.45 = 91 + 426.16$$

Λ Eq of C_2H_5COOH

$$\Lambda_{eq.\,C_2H_5COOH} = 426 + 91 - 126.45$$

Question

$Na^{\oplus}, Cl^{\ominus}$

$\text{Molten NaCl} \xrightarrow{\text{current}} Na + Cl_2 \rightarrow \text{anode}$

Na → cathode deposition

During the electrolysis of brine, by using inert electrodes.

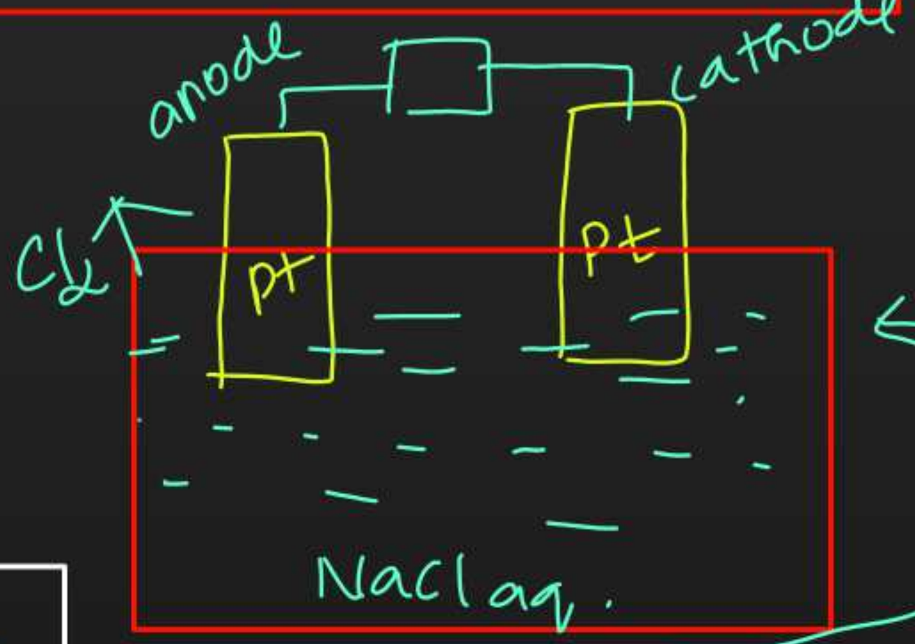
A. O_2 liberates at anode

B. H_2 liberates at andoe

C. Na deposits on cathode

D. Cl_2 liberates at anode

$\boxed{pH > 7}$



$H_2\uparrow$

$2H^{\oplus} + 2e^{\ominus} \longrightarrow H_2\uparrow$

$Na^{\oplus}$ $Cl^{\ominus}$

$H^{\oplus}$ $OH^{\ominus}$

cathode anode

→ chinetically slow

→ Overvoltage potential

bubble voltage

Question

For spontaneity of a cell, which is correct?

- A $\Delta G = -$ ve ✓, $\Delta E = 0$ ✗
- B $\Delta G = -$ ve, $\Delta E = +$ ve
- ~~C~~ $\Delta G = -$ ve
- D $\Delta G = 0$, $\Delta E = 0$

$$\boxed{\Delta G = \Delta H - T\Delta S}$$

$\Delta G = -$ve

$\boxed{\Delta E \rightarrow \text{change in internal energy.}}$

E_{cell} or EMF

$\rightarrow$ +ve.

Question

Specific conductance of 0.1 M HNO_3 is 6.3×10^{-2} ohm^{-1} cm^{-1}. The molar conductance of the solution is

A 315 ohm^{-1}cm^{2}mol^{-1}

B 6.300 ohm^{-1}cm^{2}mol^{-1}

C 63.0 ohm^{-1}cm^{2}mol^{-1}

[✓ D] 630 ohm^{-1}cm^{2}mol^{-1}

$$\Lambda_m = \frac{1000 \times K}{C}$$

$$\Lambda_m = \frac{1000 \times 6.3 \times 10^{-2}}{0.1 \rightarrow 1/10}$$

$$= \cancel{1000} \times 6.3 \times \cancel{10^{-2}} \times \cancel{10^{1}}$$

$$= 6.3 \times 10^{2}$$

$$= 630 \text{ S cm}^{-1}$$

or ohm^{-1} cm^{-1}

Question

In fuel cells ______ are used as catalysts.

A Nicle-cadmium

B Zinc-mercury

C Lead-manganese

[✓ D] Platinum-palladium

Question

The molar conductivity is maximum for the solution of concentration

A 0.002 M

B 0.005 M

[✓ C] 0.001 M

D 0.004 M

→ more dilution

↓

more H_2O

↓

less concentration.

Question

The resistance of 0.01m KCl solution at 298 K is 1500 Ω. If the conductivity of 0.01 m KCl solution at 298 K is 0.1466×10^{-3} S cm^{-1}. The cell constant of the conductivity cell in cm^{-1} is

[✓ A] 0.219

B 0.291

C 0.301

D 0.194

$$K = \frac{1}{R} \times \boxed{\frac{l}{a}}$$

$$0.1466 \times 10^{-3} = \frac{1}{1500} \times \frac{l}{a}$$

$$0.1466 \times 10^{-3} \times 1500 = l/a$$

$$0.147 \times \cancel{10^{-3}} \times 15 \times \cancel{10^{2}}$$

$$0.0147 \times 15 =$$

Question

$$Ag^+ \xrightarrow{e^-} Ag$$

$$Ag^+ \xrightarrow{e^-} Ag$$

$H_2(g) + 2AgCl(s) \rightleftharpoons 2Ag(s) + 2HCl(aq)$

E°_{cell} at 25°C for the cell is 0.22 V. The equilibrium constant at 25°C is

- A 2.8×10^7
- B 5.2×10^8 ✗
- C 2.8×10^5
- D 5.2×10^4 ✗

$$E^\circ_{cell} = \frac{0.0591}{n} \log K_c$$

$$0.22 = \frac{0.059}{2} \log K_c$$

$$\frac{0.22 \times 2}{0.059} = \log K_c$$

$$\frac{\overset{0.22}{\cancel{0.44}}}{\underset{0.03}{\cancel{0.06}}} = \log K_c$$

$$\frac{\overset{7.3}{\cancel{22}}}{\cancel{3}}$$

$$\log K_c = 7.3$$

$$K_c = \underbrace{\text{Antilog}(7.3)}_{10^7}$$

$$\text{Antilog}(7.3)$$

$$\text{Antilog}(7 + 0.3)$$

$$\text{Antilog}(\log 10^7 + \log 2)$$

$$\cancel{\text{Antilog}}\left[\cancel{\log}(2\times10^7)\right]$$

$$= \underline{\underline{2\times10^7}}$$

$$\log 2 = 0.3010$$

$$\log 10^7 = 7\log_{10}10$$

$$= 7\times1$$

$$= 7 \,//$$

$$\log a + \log b = \log ab$$

$$\underline{\log 10^7 + \log 2 = 2\times10^7}$$

Question

The pair of electrolytes that posses same value of the constant (A) in the Debye-Huckel-Onsager equation, $\Lambda_m = \Lambda_m^\circ - A\sqrt{C}$ is

A $\overset{2+}{\mathrm{Mg}}\overset{2-}{\mathrm{SO_4}}, \overset{+}{\mathrm{Na_2}}\overset{2-}{\mathrm{SO_4}}$

~~**B**~~ $\overset{+}{\mathrm{NH_4}}\overset{-1}{\mathrm{Cl}}, \overset{+}{\mathrm{Na}}\overset{-1}{\mathrm{Br}}$

C $\overset{+}{\mathrm{Na}}\overset{-}{\mathrm{Br}}, \overset{+2}{\mathrm{Mg}}\overset{-2}{\mathrm{SO_4}}$ ✗

D $\overset{+}{\mathrm{Na}}\overset{-1}{\mathrm{Cl}}, \overset{\boxed{+2}}{\mathrm{Ca}}\overset{-1}{\mathrm{Cl_2}}$ ✗

Question

$2I^{\ominus} \longrightarrow I_2 \longrightarrow \underline{\underline{2e^{\ominus}}}$

Given $E^{\circ}_{\frac{Fe^{3}}{Fe^{2+}}} = +\underline{0.76\ V}$ and $E^{\circ}_{\frac{I_2}{I^-}} = +\underline{0.55\ V}$. Then equilibrium constant for the reaction taking place in galvanic cell consisting of above two electrodes is $\left[\frac{2.303\ RT}{F} = 0.06\right]$

A $\underline{1 \times 10^{7}}$

B 1×10^{9}

C 3×10^{8}

D 5×10^{12}

$E^{0}_{cell} = E^{0}_{cathode} - E^{0}_{anode}.$

$= 0.76 - 0.55$

$E^{0}_{cell} = \underline{\underline{0.21\ V}}$

$E^{0}_{cell} = \frac{0.06}{n} \log K_c$

$0.21 = \frac{0.06}{2} \log K_c$

$\frac{0.42}{0.06} = \frac{\overset{\cancel{21}\ 7}{\cancel{42}} \times \cancel{100}}{\cancel{100} \times \underset{\cancel{3}}{\cancel{6}}}$

$\boxed{\log K_c = 7}$ $\longrightarrow K_c = \text{Antilog}(7.0)$

1×10^{7}

$\log_{10} 1 = 0 \longrightarrow 10^0 = \underline{1}$ Antilog $0 \longrightarrow \underline{1}$

$\log_{10} 2 = 0.3$

$\log 3 = 0.477$

$\log 4 = 0.60$

$\log 5 = 0.679 \simeq 0.7$

$\log 6 = 0.8$ $\log 9 = 0.95$

$\log 7 = 0.85$ $\log_{10} 10 = \underline{1}$

$\log 8 = 0.9$

Question

$H_2O + F_2 \longrightarrow HF + O_2\uparrow$

F_2 ↳ O.A (strong)

If an aqueous solution of NaF is electrolyzed between inert electrodes, the product obtained at anode is

- (A) F_2
- (B) H_2 ✗
- (C) Na ✗
- (D) O_2 ✓

cathode

↳ $H_2\uparrow$ released.

anode

$OH^{\ominus} \longrightarrow O_2\uparrow$

Quickly accepts ← $H^{\oplus}$

aq. NaF

$H^{\oplus}$ $OH^{\ominus}$

$Na^{\oplus}$ $F^{\ominus}$ ✗

cathode ↓ accepte $e^{\ominus}$

anode ↓ $e^{\ominus}$ donate.

$F^{\ominus} \longrightarrow F + e^{\ominus}$ ✗

Question

For a cell involving two electron changes. $E^\circ_{\text{cell}} = 0.3\ \text{V}$ at 25°C. Then equilibrium constant of the reaction is

A 10^{-10}

B 3×10^{-2}

C 10

[✓ D] 10^{10}

$$E^\circ_{cell} = \frac{0.0591}{n}\log K_c$$

$$0.3 = \frac{0.059}{2}\log K_c$$

$$0.3 \times 2 = 0.06 \log K_c$$

$$\therefore \log K_c = \frac{0.6}{0.06} = \frac{\cancel{6} \times 10\cancel{0}}{10 \times \cancel{6}}$$

$$\log K_c = 10$$

$$K_c = \text{Antilog } 10.0$$

$$\boxed{K_c = 1 \times 10^{10}}$$

Question

anode cathode

In $H_2 - O_2$ fuel cell the reaction occurring at cathode is

A $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

B $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4\overset{\ominus}{OH}(aq)$

C $H^+ + e^- \rightarrow \frac{1}{2}H_2$

D $H^+(aq) + \overset{\ominus}{OH}(aq) \rightarrow H_2O\,(l)$

H_2 O_2

anode cathode

Question

Molar volume → 1 mole of H_2 gas → 22.4L.

$\frac{63.5}{2} = 31.7$

The mass of copper that will be deposited at cathode in electrolysis of 0.2 M solution of copper sulphate when a quantity of electricity equal to that required to liberate 2.24 L of hydrogen from 0.1 M aqueous H_2SO_4 is (At mass of Cu = 63.5)

$\overset{+2}{CuSO_4}$

A 1.59 g

B 318 g

C 6.35 g

D 12.70 g

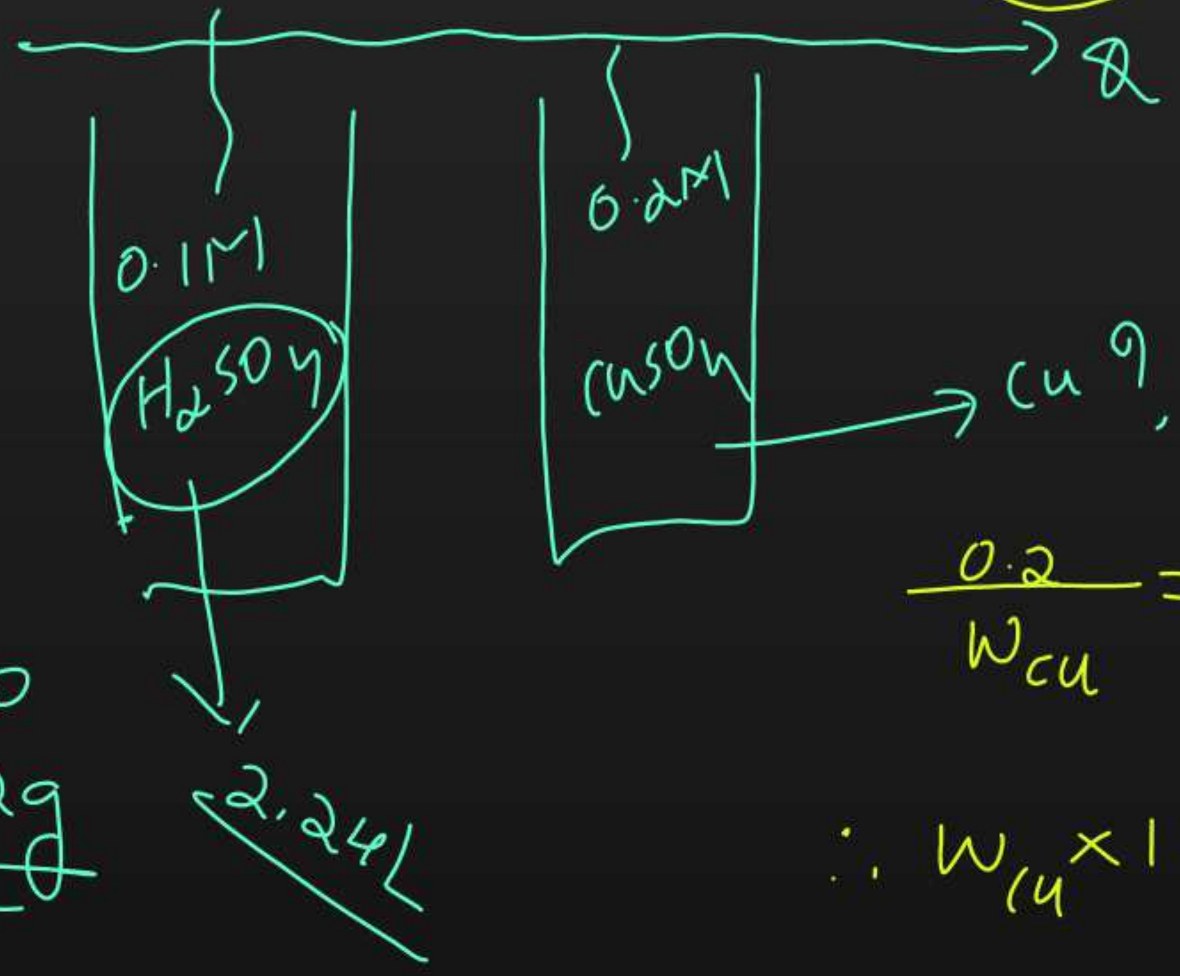
22.4L → 2g (H_2)

2.24L → ?

$$\frac{2.24\times2}{22.4}$$

$$= \frac{\cancel{224}\times\cancel{10}\times2}{\cancel{100}\times\cancel{224}} = 2/10 = 0.2g$$

Mass of H_2 is 0.2g.



$$\frac{W_{H_2}}{W_{Cu}} = \frac{\text{Eq.wt of } H_2}{\text{Eq.wt Cu}}$$

$$\frac{0.2}{W_{Cu}} = \frac{1}{31.7g}$$

$$\therefore W_{Cu}\times1 = 0.2\times31.7$$

$$W_{Cu} = \frac{31.7\times2}{10} = \frac{6.34}{10}$$

6.34g

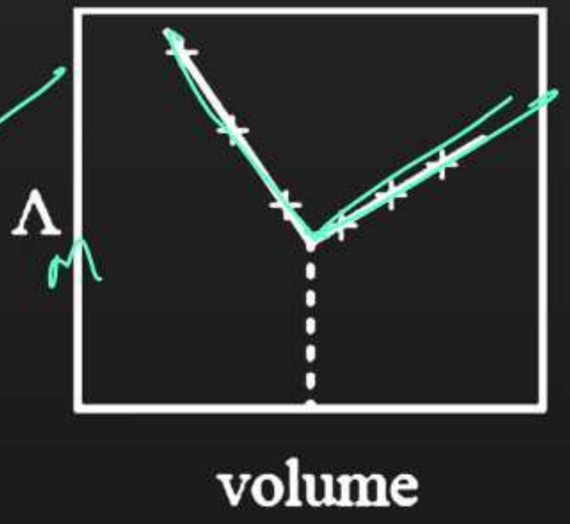
Question

$AgNO_3 + KCl \longrightarrow \underset{(s)}{AgCl}\downarrow + KNO_3$

$K^{\oplus} + NO_3^{\ominus}$

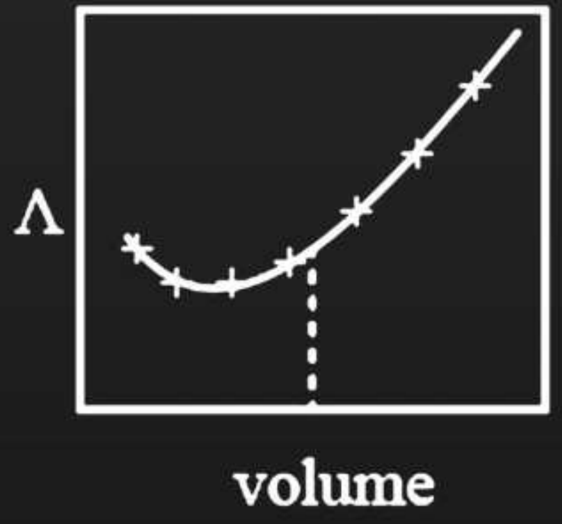
AgNO$_{(aq)}$ was added to an aqueous KCl$_{(aq)}$ solution gradually and the conductivity of the solution was measured. The plot of conductance (Λ) versus the volume of AgNO$_3$ is

A



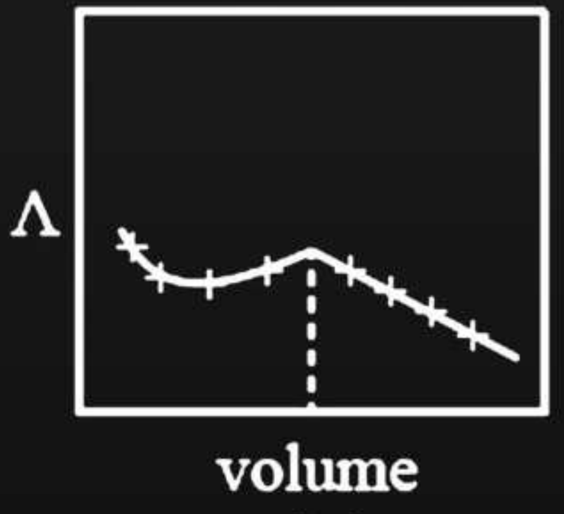
(P)

B



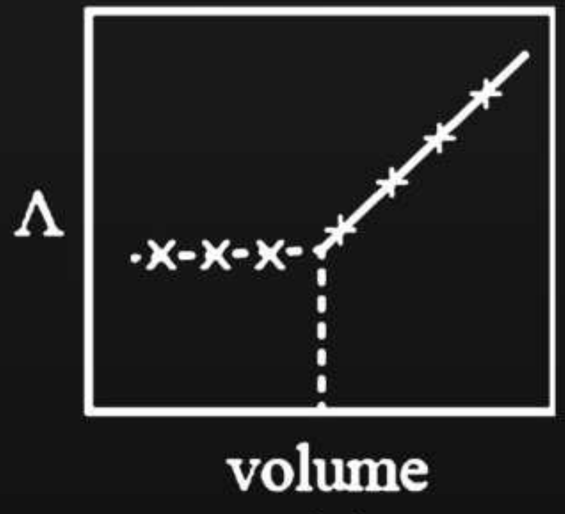
(Q)

C



(R)

D



(S)

hydration

H_2O $K^{\oplus}$ H_2O $NO_3^{\ominus}$ H_2O

H_2O

$KNO_3 + H_2O \rightarrow KOH + HNO_3$

Question

Which of the following is not an example of redox reaction ?

A $\underset{\sim\sim}{\overset{+2}{Cu}O} + \overset{0}{H_2} \rightarrow \underline{\overset{0}{Cu}} + \overset{+1}{H_2}O$ ✓

B $\underset{\sim\sim}{Fe_2O_3} + 3CO \rightarrow \boxed{2Fe} + 3CO_2$ ✓

C $\boxed{\overset{0}{2K}} + \boxed{\overset{0}{F_2}} \rightarrow 2\overset{+}{K}\overset{-1}{F}$ ✓

[✓ D] $\overset{+2}{Ba}\overset{-1}{Cl_2} + \overset{+1}{H_2}\overset{-2}{SO_4} \rightarrow \overset{+2}{Ba}\overset{-2}{SO_4} + 2\overset{+1}{H}\overset{-1}{Cl}$ ✗

Question

Identify disproportionation reaction

A $CH_4 + 2\overset{0}{O_2} \rightarrow \overset{-2}{CO_2} + 2H_2\overset{-2}{O}$ ✗

B $CH_4 + 4Cl_2 \rightarrow C\overset{-1}{\underline{Cl_4}} + 4H\overset{-1}{\underline{Cl}}$ ✗

C $2F_2 + 2OH^- \rightarrow \overset{\ominus}{\underline{2F^-}} + \overset{\ominus}{\underline{OF_2}} + H_2O$ ✗

R

D ✓ $\boxed{2\overset{+4}{NO_2}} + 2\underset{0}{OH^-} \rightarrow \overset{+3}{NO_2^-} + \overset{+5}{NO_3^-} + H_2O$

$x + (-2\times2) = 0$

$x - 4 = 0$

$x = +4$

$x + (-2\times2) = -1$

$x - 4 = -1$

$x = +3$

$x + (-2\times3) = -1$

$x - 6 = -1$

$x = +5$

Question

Which of the following elements does not show disproportionation tendency ?

A Cl

B Br

[✓ C F]

D I

Question

The reducing agent in the given equations +1 [2023]

$4\overset{0}{Ag}(s) + 8CN^-(aq) + 2H_2O(aq) + O_2(g) \rightarrow 4[Ag(CN)_2]^-(aq) + 4OH^-(aq)$ → ①

$2[\underset{+1}{Ag}(CN)_2]^-(aq) + Zn(s) \rightarrow [Zn(CN)_4]^{2-}(aq) + 2Ag(s)$ → ②

$2e^-$

- **A** Zn ✓
- **B** O_2
- **C** H_2O
- **D** CN^-

$$\boxed{Zn + 2Ag^+ \longrightarrow Zn^{2+} + 2Ag}$$

reducing agent

↓

$e^\ominus$ donors

Reductant

Question

$_{16}S \longrightarrow 1s^2, 2s^2, 2p^4, \boxed{3s^2, 3p^4} \longrightarrow \boxed{6e^-}$

What is the oxidation number of S in $H_2S_2O_8$? [2023] → VIMP.

→ peroxodisulphuric acid.

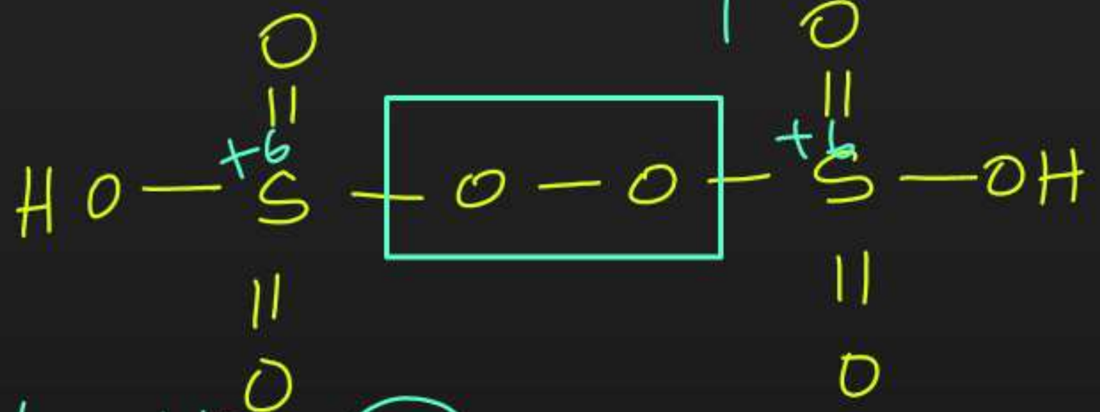
Marshall's acid.

A + 5

B + 4

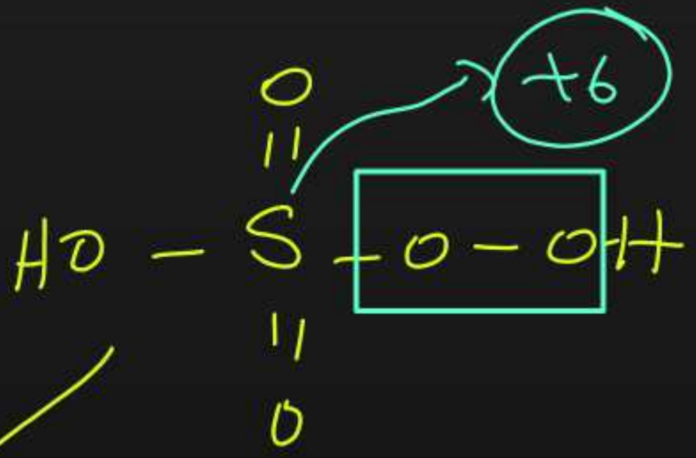
C + 7

[✓ **D**] + 6



$+6 + 6 = \frac{+12}{2} = \boxed{+6}$

Peroxo Mono sulphuric acid. ←



H_2SO_5 → Caro's acid

Question

In which of the following compounds, an element exhibits two different oxidation states ? ⟶ [2022]

(A) ✓ $NH_{\cancel{2}\,4}NO_3$ ⟶ $\overset{-3}{N}H_4^{+}$ $\overset{+5}{N}O_3^{-}$

(B) N_2H_4 ⟶ $\boxed{\overset{-2}{N}H_2-\overset{-2}{N}H_2}$

$\overset{-2}{N} \leftarrow H$

↑

+1

(C) NH_3 ⟶ $\boxed{\overset{-3}{N}H_3}$

(D) NH_2CONH_2 ⟶ $\boxed{\overset{-2}{H_2N} \rightarrow \overset{O}{\overset{\uparrow\uparrow}{C}} \leftarrow \overset{-2}{N}H_2}$

NH_4^{+}

$x + 4 = +1$

$x = +1 - 4$

$x = -3$

$\boxed{NO_3^{-}}$

$x + (-2\times 3) = -1$

$x - 6 = -1$

$x = +5$

Question

For the redox reaction $\overset{\checkmark}{x}MnO_4^- + \overset{\checkmark}{y}H_2C_2O_4 + zH^+ \rightarrow \overset{\checkmark}{m}Mn^{2+} + \overset{\checkmark}{n}CO_2 + pH_2O$. The values of x, y, m and n are [2018]

A 10, 2, 5, 2

B 2, 5, 2, 10 ✓

C 6, 4, 2, 4

D 3, 5, 2, 10

Oxidation no. method ✓

O.N ↓ by 5e⊖

$\overset{+7}{2}MnO_4^{\ominus} + 5C_2O_4^{2-} + H^{\oplus} \longrightarrow 2Mn^{+2} + 5\times 2CO_2$

O.N ↓, by 2e⊖

$\underset{x}{2}MnO_4^{\ominus} + \underset{y}{5}C_2O_4^{2-} + \underset{z}{16}H^{\oplus} \longrightarrow \underset{m}{2}Mn^{2+} + \underset{n}{10}CO_2 + 8H_2O$

Oxidation no. method ✓

OH^-

acidic medium

O.N ↓ $3e^- \times 2 = 6e^-$

$1\times$

$\overset{+6}{\underline{1}Cr_2O_7^{2-}} + 6Fe^{2+} \longrightarrow 2\overset{3+}{Cr} + 6Fe^{3+}$

$1e^-$ O.N ↑

$$Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \longrightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$$

Basic medium.

O.N ↑ by ~~$6e^-$~~ $2e^-$

$$\overset{+7}{MnO_4^-} + \overset{-1}{I^-} \longrightarrow \overset{+4}{MnO_2} + \overset{+5}{IO_3^-}$$

O.NO ↓ by ~~3~~ $1e^-$

$$H_2O + 2MnO_4^- + I^- \longrightarrow 2MnO_2 + IO_3^- + 2OH^-$$

range → 4 to 5

Chemical kinetics → N.P

Chemical kinetics → Max

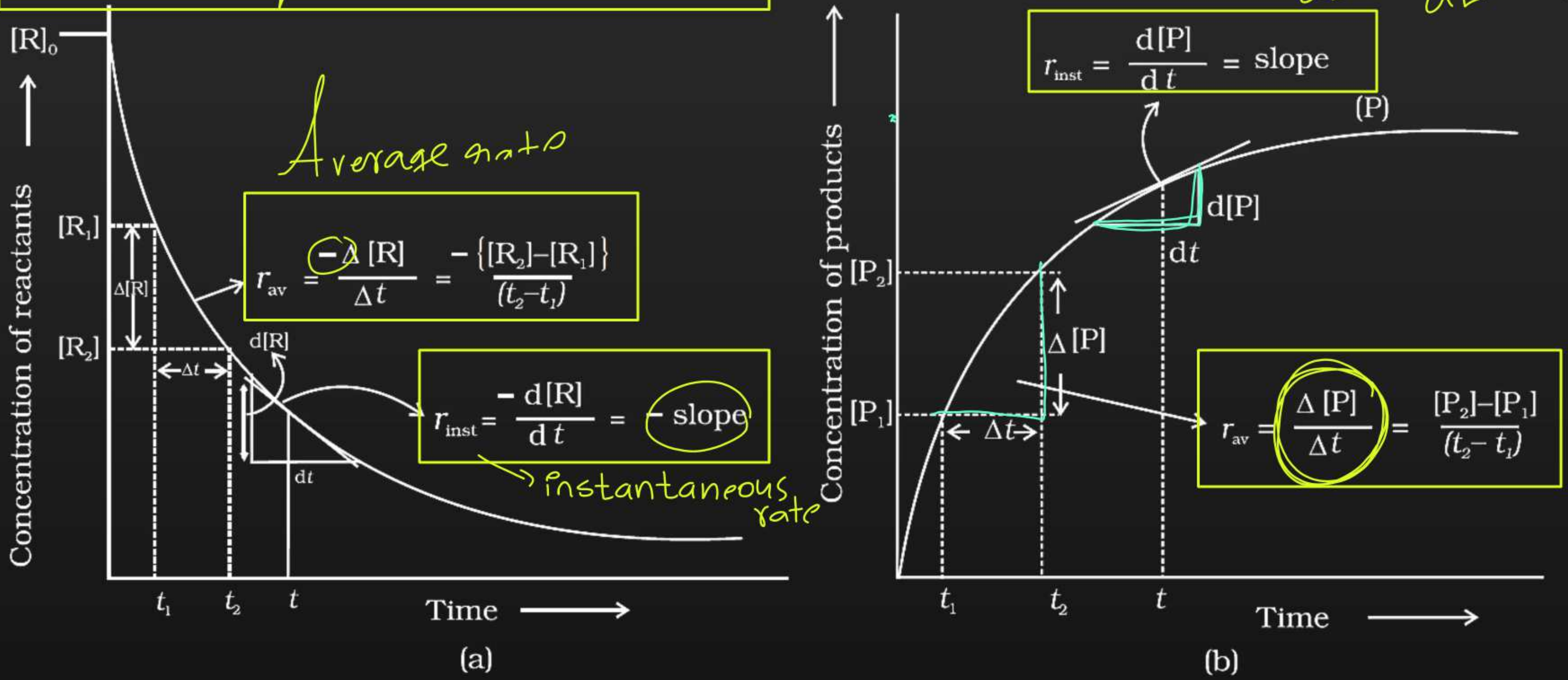
↓

5 Qsn

Rate of the reaction

$R \longrightarrow P$

$$\text{rate} = \frac{-d[R]}{dt} = \frac{d[P]}{dt}$$



Average rate

$$r_{av} = \frac{-\Delta[R]}{\Delta t} = \frac{-\{[R_2]-[R_1]\}}{(t_2-t_1)}$$

$$r_{inst} = \frac{-d[R]}{dt} = -\text{slope}$$

instantaneous rate

$$r_{inst} = \frac{d[P]}{dt} = \text{slope}$$

$$r_{av} = \frac{\Delta[P]}{\Delta t} = \frac{[P_2]-[P_1]}{(t_2-t_1)}$$

Instantaneous and average rate of a reaction

$$5Br^-(aq) + BrO_3^-(aq) + 6H^+(aq) \rightarrow 3Br_2(aq) + 3H_2O(l)$$

$$\text{Rate} = -\frac{1}{5}\frac{\Delta[Br^-]}{\Delta t} = -\frac{\Delta BrO_3^-}{\Delta t} = -\frac{1}{6}\frac{\Delta[H^+]}{\Delta t} = \frac{1}{3}\frac{\Delta[Br_2]}{\Delta t} = \frac{1}{3}\frac{\Delta[H_2O]}{\Delta t}$$

$$H_2 + \frac{1}{2}O_2 \longrightarrow H_2O$$

$$\text{rate} = \frac{-d[H_2]}{dt} = -2\frac{d[O_2]}{dt} = \frac{d[H_2O]}{dt}$$

Question

Which of the following expressions is correct for the rate of reaction given below?

$5Br^{-}_{(aq)} + BrO^{-}_{3(aq)} + 6H^{+}_{(aq)} \rightarrow 3Br_{2(aq)} + 3H_2O_{(l)}$

A $\frac{\Delta[Br^-]}{\Delta t} = 5\frac{\Delta[H^+]}{\Delta t}$ ✗

B $\frac{\Delta[Br^-]}{\Delta t} = \frac{6}{5}\frac{\Delta[H^+]}{\Delta t}$ ✗

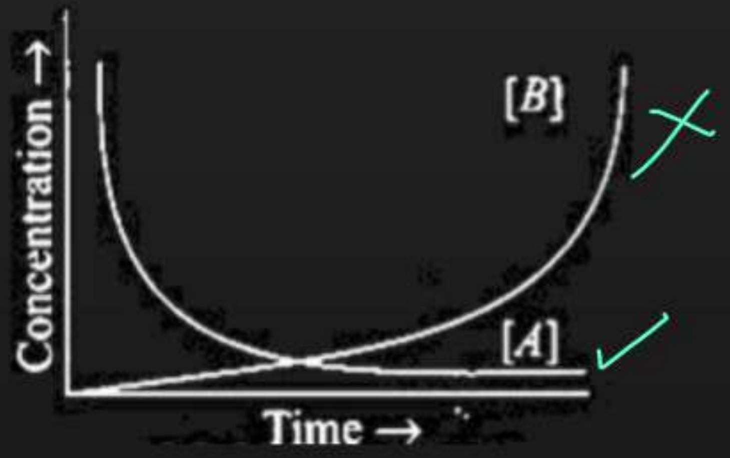
[✓ C] $\frac{\Delta[Br^-]}{\Delta t} = \frac{5}{6}\frac{\Delta[H^+]}{\Delta t}$

D $\frac{\Delta[Br^-]}{\Delta t} = 6\frac{\Delta[H^+]}{\Delta t}$

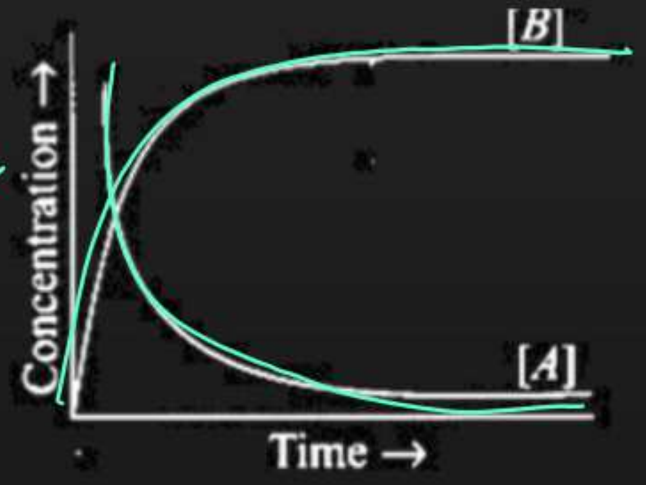
Question

Consider the reaction A → B. The concentration of both the reactants and the products varies exponentially with time. Which of the following figures correctly describes the change in concentration of reactants and products with time?

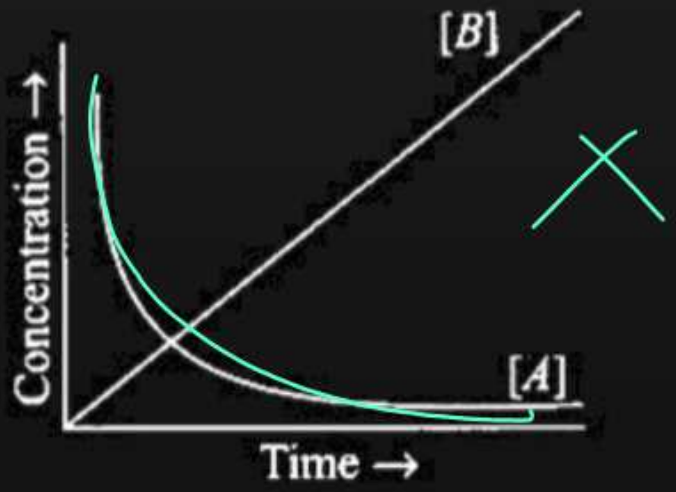
A



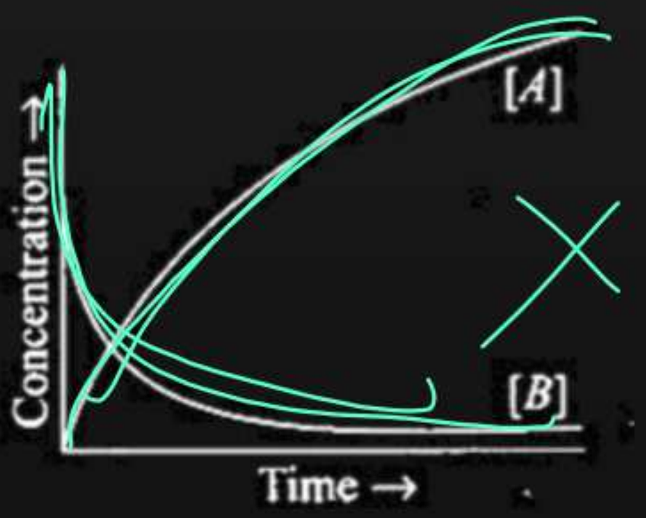
B

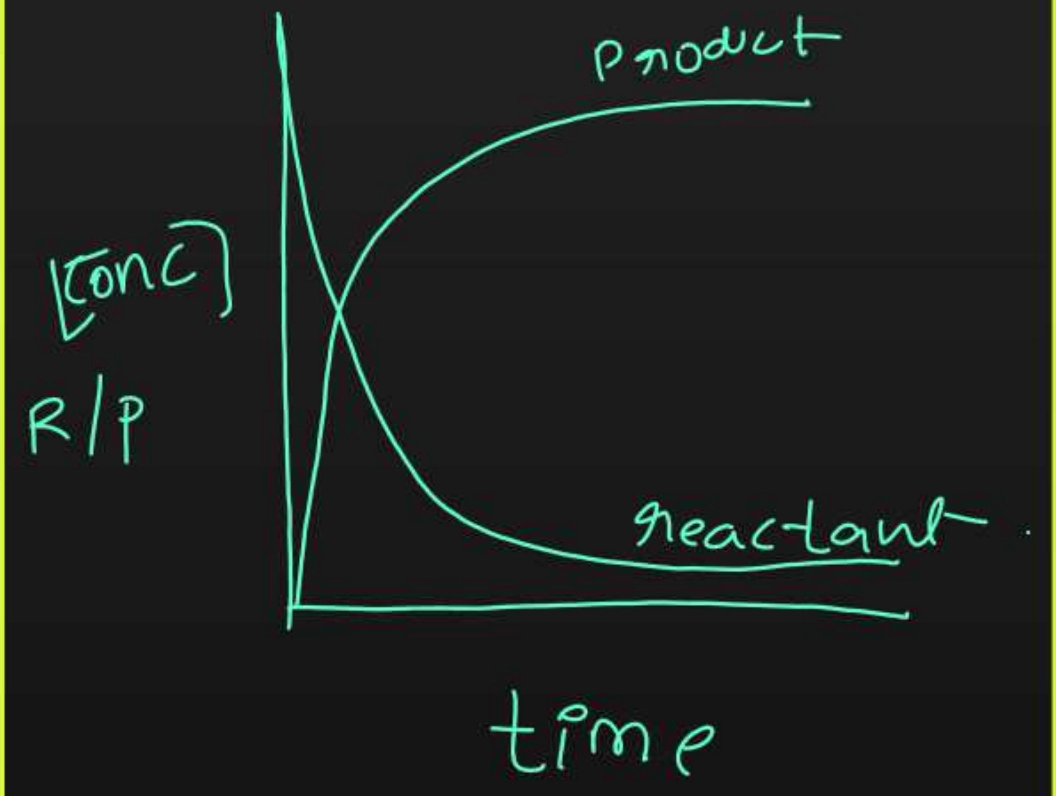


C



D





Law of mass action.

$$aA + bB \underset{R_b}{\overset{K_f}{\rightleftharpoons}} cC + dD$$

$$\left.\begin{aligned} R_f &\propto [A]^a[B]^b \\ R_b &\propto [C]^c[D]^d \end{aligned}\right\} \longrightarrow \begin{aligned} R_f &= K_f[A]^a[B]^b \\ R_b &= K_b[C]^c[D]^d. \end{aligned}$$

Law of mass action

$$R_f = R_b \longrightarrow K_c$$

$$K_f[A]^a[B]^b = K_b[C]^c[D]^d \longrightarrow \text{Theoretical}$$

$$\therefore \frac{K_f}{K_b} = K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}.$$

Question

$\text{K} \rightarrow$ only on temp

$\rightarrow$ Ea, P.

$$\text{Rate} = K[A]^x[B]^y$$

Rate law for the reaction 'B' is doubled, keeping the concentration of 'A' constant, the value of rate constant will be $\longrightarrow$

A The same

B Doubled

C Quadrupled

D Halved

$$A + B \longrightarrow \text{product}$$

$$r_1 = \text{rate} = k[A][B] \longrightarrow ①$$

$$r_2 = K[A][2B] \longrightarrow ②$$

$$r_2 = 2k[A][B] \rightarrow ③.$$

$$\frac{③}{①} \quad \frac{r_2}{r_1} = \frac{2k\cancel{[A]}\cancel{[B]}}{k\cancel{[A]}\cancel{[B]}}$$

rate $\rightarrow$ doubled

rate constant $\rightarrow$ same.

$$\boxed{r_2 = 2r_1}$$

Order → zero order
first order
2nd order
nth order.
fractional order.

Rate law – EXAMPLES

Reaction	Experimental rate expression
$CHCl_3 + Cl_2 \rightarrow CCl_4 + HCl$	Rate $= k[CHCl_3]^{1}[Cl_2]^{1/2} \rightarrow 1 + \frac{1}{2} = 3/2.$
$CH_3COOC_2\overset{1}{H_5} + H_2\overset{0}{O} \rightarrow CH_3COOH + C_2H_5OH$	Rate $= k[CH_3COOC_2H_5]^1[H_2O]^0$

In these reactions, the exponents of the concentration terms are not the same as their stoichiometric coefficients.

Rate law for any reaction cannot be predicted by merely looking at the balanced chemical equation, i.e., theoretically but must be determined experimentally.

Order

Some Common Examples of Reactions of Different Orders.

Rate $= k[H_2]^0[Cl_2]^0$

(i) * Photosynthesis of HCl $H_2(g) + Cl_2(g) \xrightarrow{hv} 2HCl$ ✓

Experimentally, it is found that the rate of reaction is independent of the concentration of H_2 and Cl_2. Hence, it is a reaction of the zero order. Evidently, for such reactions, Rate $= k$

(ii) * Dissociation of N_2O_5 $\boxed{2N_2O_5 \longrightarrow 4NO_2 + O_2}$ → VIMP

Experimentally, it is found that : Rate of reaction $\propto [N_2O_5]^1$

i.e., it depends only on one concentration term. Hence, the reaction is said to be of first order.

(iii) Combination between NO_2 and F_2 to form → $\boxed{\text{Rate} = k[NO_2][F_2]}$

[redacted] $\boxed{2NO_2(g) + F_2(g) \longrightarrow 2NO_2F(g)}$

Experimentally, it is found that : Rate of reaction $\propto [NO_2][F_2]$

i.e., it depends on one concentration term of NO_2 and one concentration term of F_2. Hence, the order of reaction is 2 , i.e., it is a reaction of the second order.

Further, it may be mentioned here that the order of a reaction may not always be a whole number. For example, at 723 K , acetaldehyde decomposes as :

$$\boxed{CH_3CHO \overset{723\ K}{\rightarrow} CH_4 + CO}$$

3/2.

Experimentally, it is found that : Rate of reaction $\propto [CH_3CHO]^{1.5}$. Hence, the order of reaction is 1.5.

Some More Examples of Reactions of Different Orders.

(a) Reactions of zero order :

Some enzyme catalysed reactions and reactions occurring on metal surfaces are reactions of zero order. For example,

(i) Decomposition of ammonia on a hot platinum surface at high pressure

$$\boxed{2NH_3 \xrightarrow[1130\ K]{Pt} N_2 + 3H_2}$$

The rate law equation is: Rate $= k[NH_3]^\circ = k^*$

(ii) Thermal decomposition of HI on gold surface $2HI \xrightarrow[\Delta]{Au} H_2 + I_2$

In these reactions, the concentration of the gaseous reactant on the metal surface remains constant because as the reactants react off on the surface, more molecules are adsorbed on the surface. Hence, rate remains constant throughout.

(b) Reactions of first order:

(i) Decomposition of ammonium nitrite : $NH_4NO_2 \longrightarrow N_2 + 2H_2O$
The rate law equation is : $\text{Rate} = k[NH_4NO_2]$

(ii) Decomposition of hydrogen peroxide : $H_2O_2 \longrightarrow H_2O + \frac{1}{2}O_2$ *
The rate law equation is : $\text{Rate} = k[H_2O_2]^{1}$

(iii) Decomposition of SO_2Cl_2: $SO_2Cl_2 \longrightarrow SO_2 + Cl_2$
The rate law equation is : $\text{Rate} = k[SO_2Cl_2]$

(iv) Hydrogenation of ethene : $C_2H_4(g) + H_2(g) \longrightarrow C_2H_6(g)$
The rate law equation is : $\text{Rate} = k[C_2H_4]$

(v) Decomposition of N_2O_5: $2\,N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)$ *
The rate law equation is : $\text{Rate} = k[N_2O_5]$

(vi) Decomposition of N_2O: $2\,N_2O \xrightarrow{600^\circ C} 2\,N_2 + O_2$

The rate law equation is : Rate $= k[N_2O]$

(vii) Natural and artificial decay of unstable nuclei, e.g., ${}^{226}_{88}Ra \longrightarrow {}^{222}_{86}Rn + {}^{4}_{2}He$

Rate $= k[Ra]$ ✓

(c) Reactions of second order :

(i) Dissociation of HI: $2HI \longrightarrow H_2 + I_2$

The rate law equation is : Rate $= k[HI]^2$ → 2nd order

(ii) Saponification of ethyl acetate:

$CH_3COOC_2H_5 + NaOH \longrightarrow CH_3COONa + C_2H_5OH$ or

$CH_3COOC_2H_5 + OH^- \longrightarrow CH_3COO^- + C_2H_5OH$

VIMP

1 1

The rate law equation is : Rate $= k[CH_3COOC_2H_5][NaOH]$

(iii) Reaction between NO and O_3: $NO + O_3 \longrightarrow NO_2 + O_2$

The rate law equation is : Rate $= k[NO][O_3]$

(d) Reactions of third order:

(i) Combination of NO and O_2 to form NO_2: $2NO + O_2 \longrightarrow 2NO_2$

The rate law equation is : Rate $= k[NO]^2[O_2]$

(ii) Combination of NO and Br_2 to form NOBr: $2NO + Br_2 \longrightarrow 2NOBr$

The rate law equation is : Rate $= k[NO]^2[Br_2]$

Extra info

avoid

(e) Reactions of fractional order

(i) Reaction between H_2 and Br_2 to form HBr: $H_2 + Br_2 \longrightarrow 2HBr$

The rate law equation is : Rate $= k[H_2][Br_2]^{1/2}$, order $= 1\frac{1}{2} = 3/2$

(ii) Combination between CO and Cl_2: $CO + Cl_2 \longrightarrow COCl_2$

The rate law equation is : Rate $= k[CO]^2[Cl_2]^{1/2}$, order $= 2\frac{1}{2}$

(iii) Dissociation of $COCl_2$: $\boxed{COCl_2 \longrightarrow CO + Cl_2}$

The rate law equation is : Rate $= k[COCl_2]^{3/2}$, order $= 1\frac{1}{2}$

VIMP

Note – 1

Molecularity of a Reaction

The number of reacting species (atoms, ions or molecules) taking part in an elementary reaction, which must collide simultaneously in order to bring about a chemical reaction is called molecularity of a reaction.

The reaction can be unimolecular when one reacting species is involved, for example, decomposition of ammonium nitrite.

$$\boxed{NH_4NO_2 \rightarrow N_2 + 2H_2O}$$

Bimolecular reactions involve simultaneous collision between two species, for example, dissociation of hydrogen iodide.

$$2HI \rightarrow H_2 + I_2$$

Trimolecular or termolecular reactions involve simultaneous collision between three reacting species, for example,

$$2NO + O_2 \rightarrow 2NO_2$$

Note – 2

Complex reaction :Reaction which takes place in many steps.

EXAMPLE :

$KClO_3 + 6FeSO_4 + 3H_2SO_4 \rightarrow KCl + 3Fe_2(SO_4)_3 + 3H_2O$ → NCERT

~~The rate of the above reaction is – 2~~

Order → 2.

Complex reaction → single steps → slow step → r.d.s → Molecularity

Note – 3

Molecularity and order for complex reaction

$$2H_2O_2 \xrightarrow[\text{Alkaline medium}]{I^-} 2H_2O + O_2$$

The rate equation for this reaction is found to be

$$\text{Rate} = \frac{-d[H_2O_2]}{dt} = \boxed{k\overset{1}{[H_2O_2]}\overset{1}{[I^-]}}$$

This reaction is first order with respect to both H_2O_2 and I^-. Evidences suggest that this reaction takes place in two steps

$$\boxed{\begin{array}{l} (1)\ H_2O_2 + I^- \rightarrow H_2O + \cancel{IO^-} \\ (2)\ H_2O_2 + \cancel{IO^-} \rightarrow H_2O + I^- + O_2 \end{array}}$$

Both the steps are bimolecular elementary reactions. Species IO^- is called as an intermediate since it is formed during the course of the reaction but not in the overall balanced equation.

The first step, being slow, is the rate determining step.

Integrated zero order & Integrated first order reaction.

$$k = \frac{[R]_0 - [R]}{t}$$

$$k = \frac{[R]_0 - \frac{[R]_0}{2}}{t}$$

$$k = \frac{2[R]_0 - [R]_0}{2t}$$

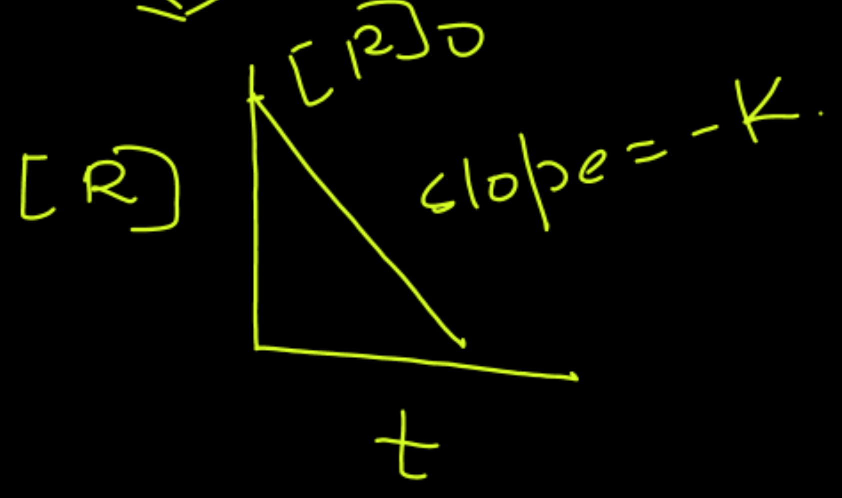
$$K = \frac{[R]_0}{2t_{1/2}}$$

$$t_{1/2} = \frac{[R]_0}{2K}.$$

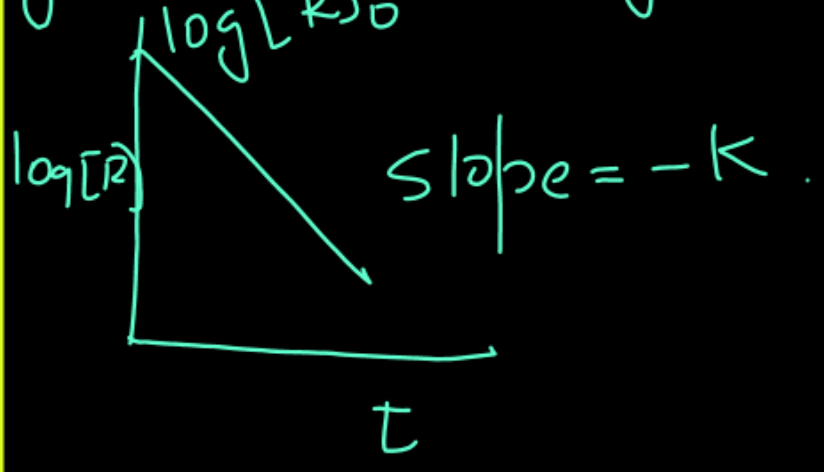
$$Kt = [R]_0 - [R]$$

$$[R] = -Kt + [R]_0$$

zero order



$$\log[R] = -Kt + \log[R]_0$$ first ord



Question

Exemplar qsn .

A graph of volume of hydrogen released vs time for the reaction between zinc and dil. HCl is shown below. On the basis of this, mark the correct option.

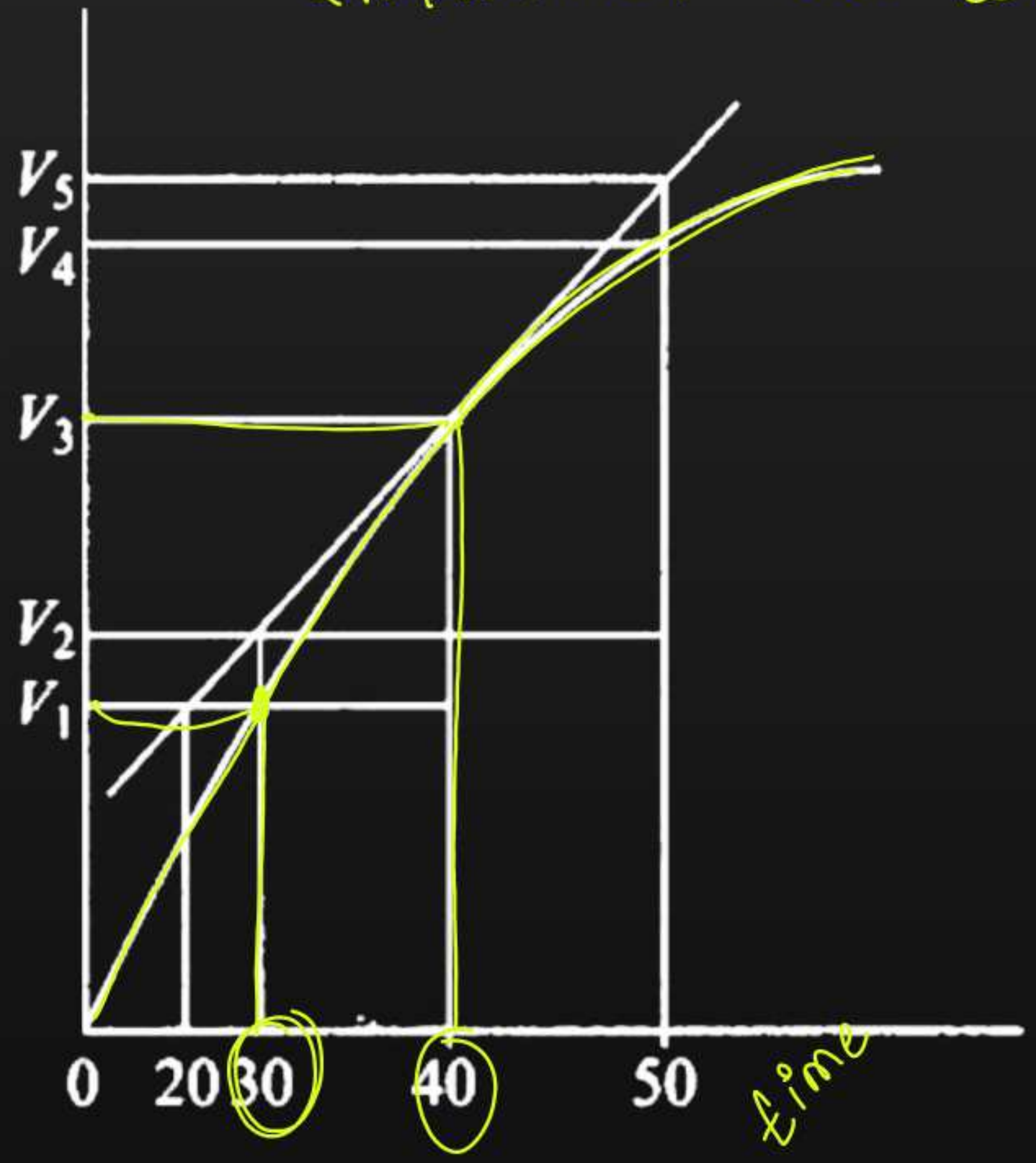
$Zn + dil\ HCl \longrightarrow ZnCl_2 + H_2\uparrow$

A Average rate up to 40 seconds is $\frac{V_3 - V_2}{40}$ ✗

B Average rate up to 40 seconds is $\frac{V_3 - V_2}{40 - 30}$ ✗

[✓ **C**] Average rate up to 40 seconds is $\frac{V_3}{40}$

D Average rate up to 40 seconds is $\frac{V_3 - V_1}{40 - 20}$ ✗



Question

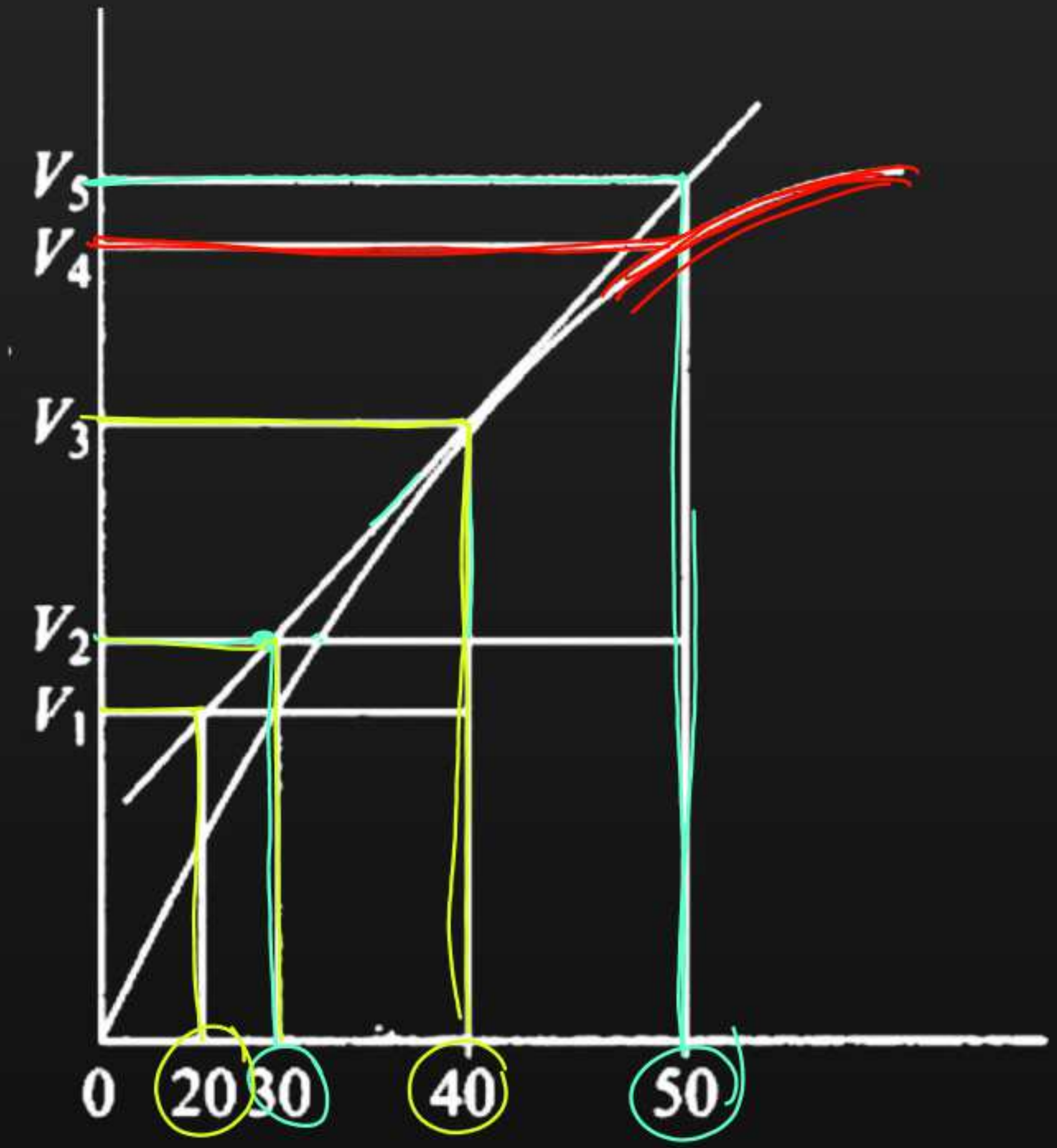
Consider the graph given Which of the following option does not show instantaneous rate of reaction at 40th second?

(A) $\frac{V_5 - V_2}{50 - 30}$

(B) $\frac{V_4 - V_2}{50 - 30}$

(C) $\frac{V_3 - V_2}{40 - 30}$

(D) $\frac{V_3 - V_1}{40 - 20}$



QUESTION

$r_{av} = -\frac{\Delta R}{\Delta t} = -\frac{[R_2 - R_1]}{t_2 - t_1}$ r_{av} for product $= \frac{\Delta P}{\Delta t}$.

Nitrogen dioxide decomposes to nitric oxide and molecular oxygen as

$2NO_{2(g)} \rightarrow 2NO_{(g)} + O_2(g)$

The concentration-time data for the consumption of NO_2 at 300°C are as follows:

Time/s	0	100	200	300
$[NO_2]/M$	8.4×10^{-3}	5.6×10^{-3}	4.3×10^{-3}	3.0×10^{-3}

$\text{rate} = -\frac{[3.0\times10^{-3}] - [4.3\times10^{-3}]}{300 - 200}$

Calculate the average rate of decomposition of NO_2.

A -1.3×10^{-5} M/s

B 1.3×10^{-5} M/s

C 2.5×10^{-5} M/s

D -2.5×10^{-5} M/s

QUESTION

The rate of a gaseous reaction is given by k [A] [B]. If the volume occupied by the reacting gases is suddenly reduced to half the original volume, the rate of the reaction relative to the initial rate will be

A 1/8

B 1/4

C 8/1

D 4

$$rate = k[A]^x[B]^y$$

$$rate_2 = k[2A][2B]$$

$$\boxed{r_2 = 4k[A][B]}$$

QUESTION

The rate law of a gaseous reaction is given by k [A] [B]. If the volume occupied by the reacting gases is suddenly reduced to ~~half the~~ 1/4th original volume, the rate of the reaction relative to the initial rate will be

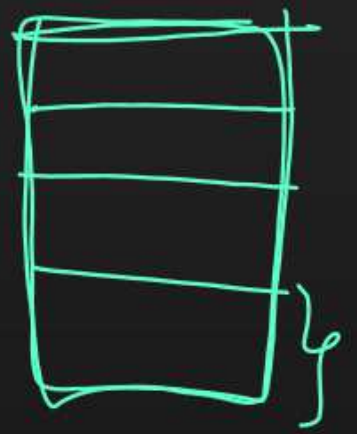
- ~~A~~ 16
- B $\frac{1}{16}$
- C 8
- D 4

$$\boxed{\text{rate} = k[A][B]}$$

$$= k[4A][4B]$$

$$= 16\,k[A][B]$$

$$\boxed{r_2 = r_1\,16}$$



QUESTION

Order of reaction is

A never zero

B never fractional

C always equal to number of molecules taking part in reaction

[✓ D] an experimentally determined quantity.

QUESTION

For a reaction $\boxed{I + OCl^- \rightarrow IO^- + Cl^-}$, in an aqueous medium, the rate of the reaction is given by $\frac{d[IO^-]}{dt} = k\frac{[I^-]^1[OCl^-]^1}{[OH^-]^1}$ The overall order of the reaction is

A −1

[✓ B 1]

C 0

D 2

$$\text{rate} = k\,\frac{[I^-]^1\,[ClO^-]^1}{[OH^-]^1}$$

$$k[I^-]^1[ClO^-]^1[OH^-]^{-1}$$

Thank

You