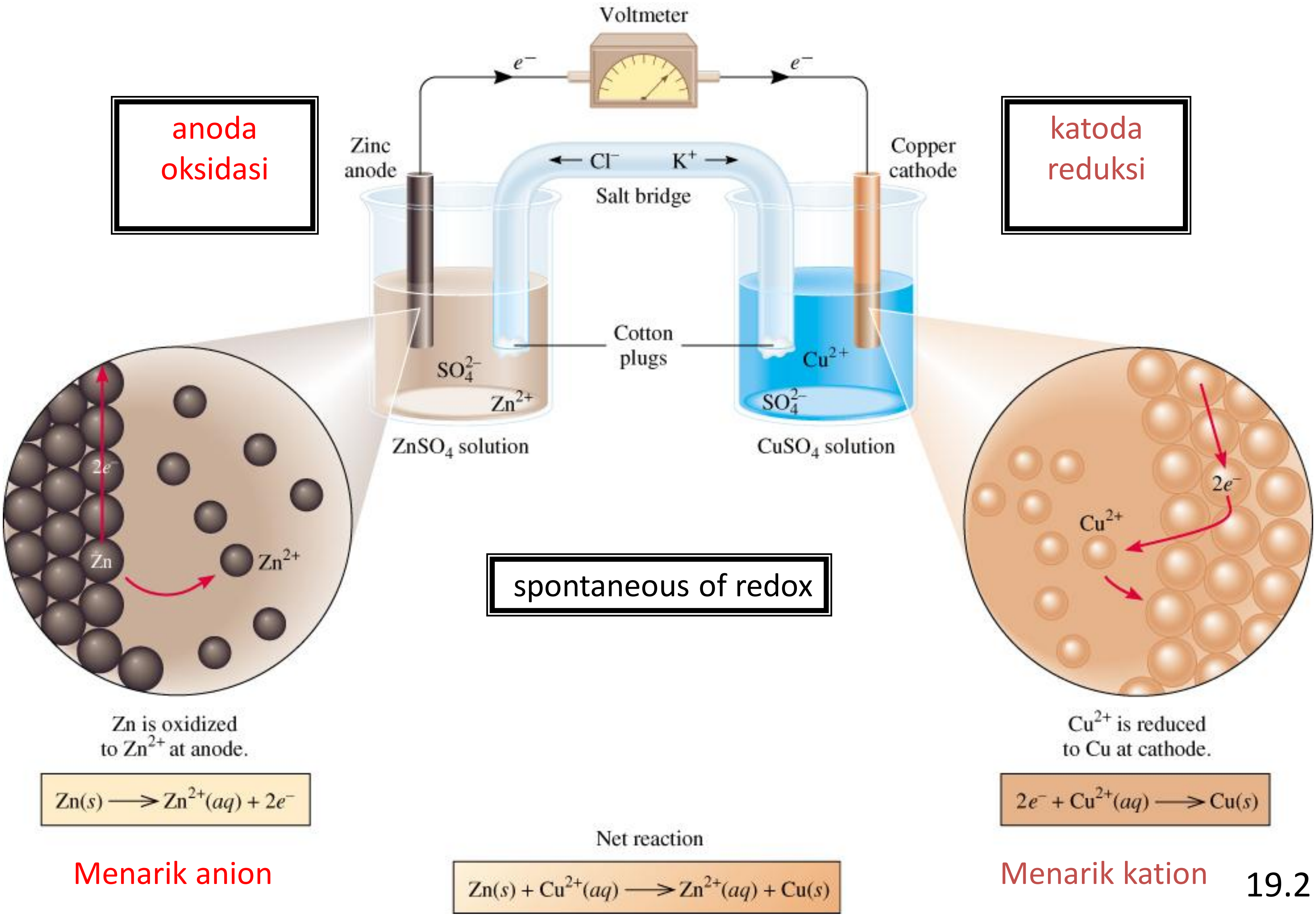
ELECTROCHEMISTRY

Electrochemical Cell



anoda
oksidasi

katoda
reduksi

spontaneous of redox

Zn is oxidized to Zn^{2+} at anode.

$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

Cu^{2+} is reduced to Cu at cathode.

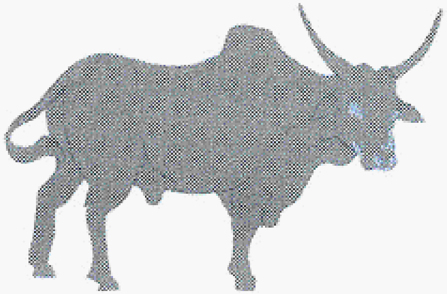
$2e^- + Cu^{2+}(aq) \longrightarrow Cu(s)$

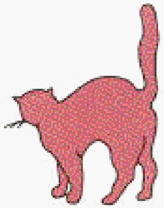
Net reaction

$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

Menarik anion

Menarik kation





Which Half-Reaction Occurs at Which Electrode? If you sometimes forget which half-reaction occurs at which electrode, you're not alone. Here are some memory aids to help:

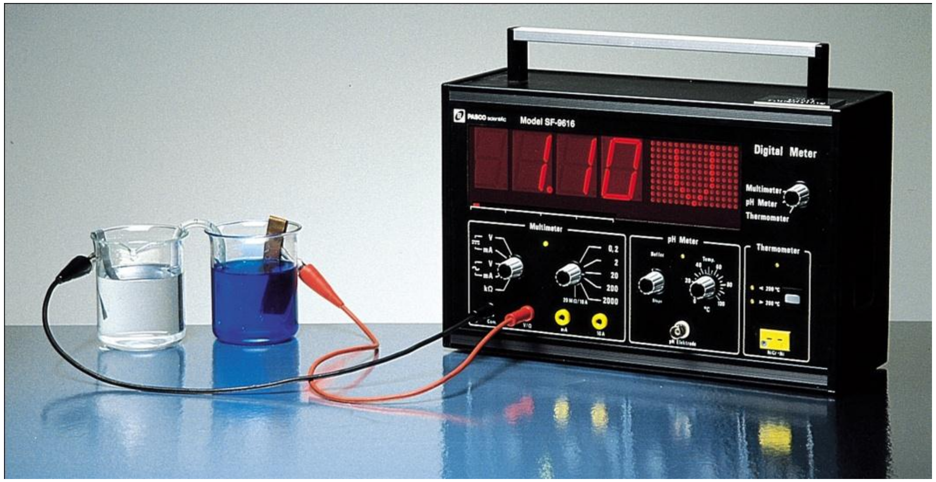
1. The words *anode* and *oxidation* start with vowels; the words *cathode* and *reduction* start with consonants.
2. Alphabetically, the *A* in anode comes before the *C* in cathode, and the *O* in oxidation comes before the *R* in reduction.
3. Look at the first syllables and use your imagination:

ANode, OXidation;
REDuction, CAThode ⇒
AN OX and a RED CAT

ELECTROCHEMISTRY CELL

DisoLute of electric potential betwen two anoda an cathoda call as:

- ***voltage ce;l***
- ***electromotif force (emf)***
- ***potential cell***



Cell Diagram

$$\mathrm{Zn}\ (s) + \mathrm{Cu}^{2+}\ (aq) \longrightarrow \mathrm{Cu}\ (s) + \mathrm{Zn}^{2+}\ (aq)$$

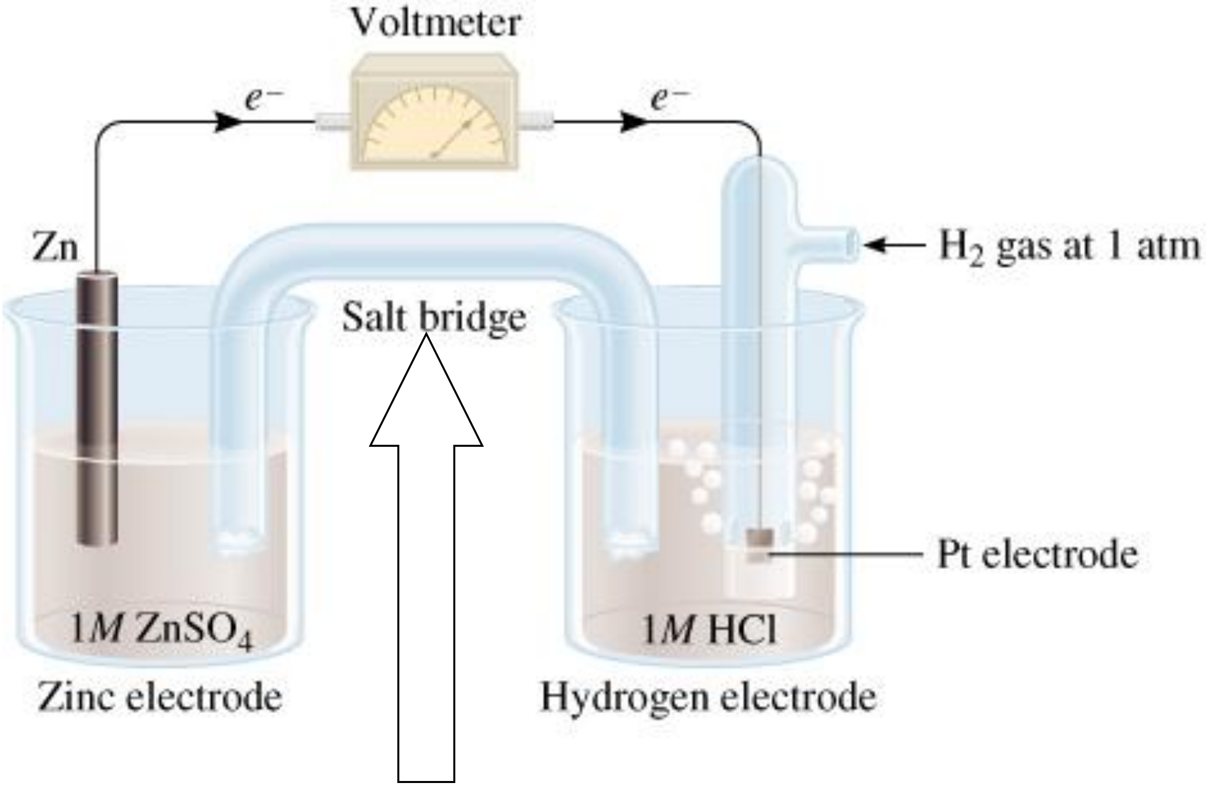
$[\mathrm{Cu}^{2+}] = 1\ M$ & $[\mathrm{Zn}^{2+}] = 1\ M$

anoda **cathoda**

$\mathrm{Zn}\ (s)\ |\ \mathrm{Zn}^{2+}\ (1\ M)\ ||\ \mathrm{Cu}^{2+}\ (1\ M)\ |\ \mathrm{Cu}\ (s)$

Tanda || utk memisahkan setengah sel

Tanda | utk memisahkan reaktan/fasa tiap setengah sel



Related of electricity betwen two half reaction

Tube contain of inert salt solutiom (KNO_3)

Two type of Cell

electrochemical cell - need “DC source” = electron pumpt (example: battery)

- electron is forced to move in the one purpose, independet on spontanity
- Electrical energy is used for nonspontan reaction can occur
- electron is moved to electron, so reduction reaction can occur.

Two type of Cell

Volta CELL OR GalvaniC Cell – pasive electric
(not need “Dc source”)

- electron can move becouse spontaneous reaction
- Using chemistry to gain energy
- electron was take from cathoda by reduction, cause moving electron to all direction
- Can be used as dc source to electrolytic cell.

How we know that reaction is spontan?

potential electric?

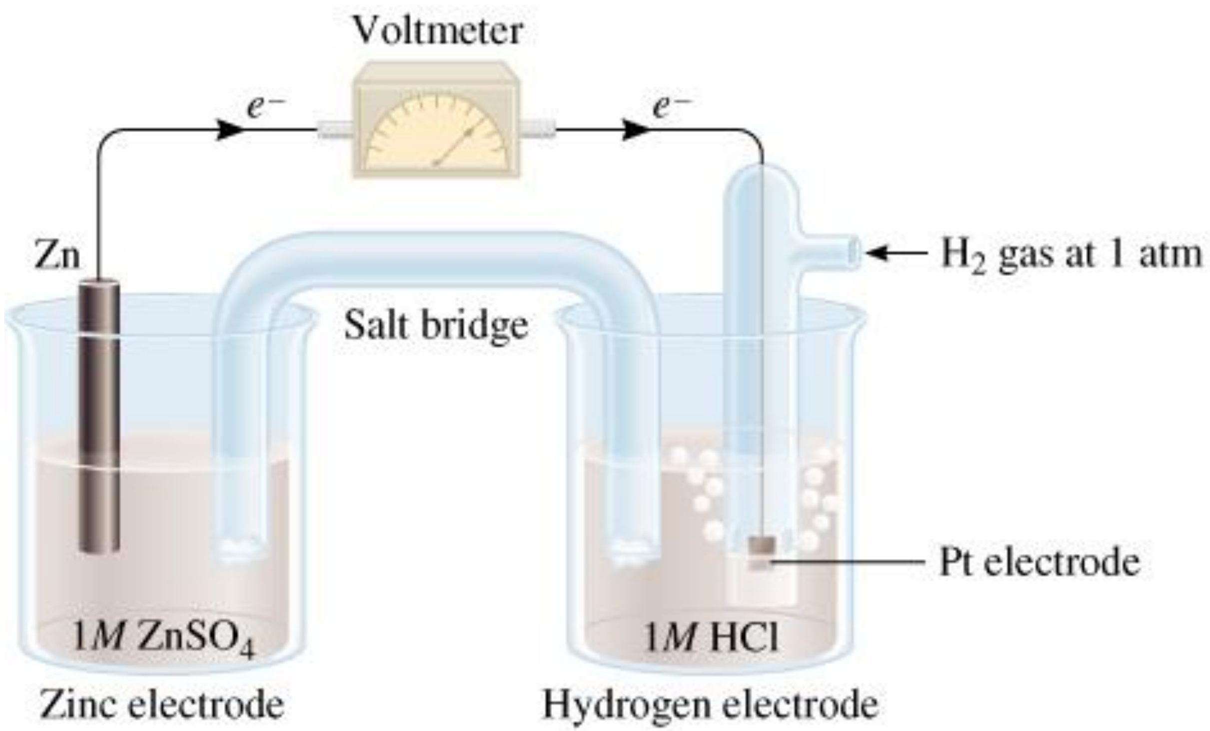
How the consentration effect the process?

electromotif (emf) is cell electricpotenstial

E (emf) $\Rightarrow$ units = volts (V)

emf is dispute of potential betwen anoda dan cathoda

Standard Potential of Electroda



Zn (s) | Zn^{2+} (1 M) || H^+ (1 M) | H_2 (1 atm) | Pt (s)

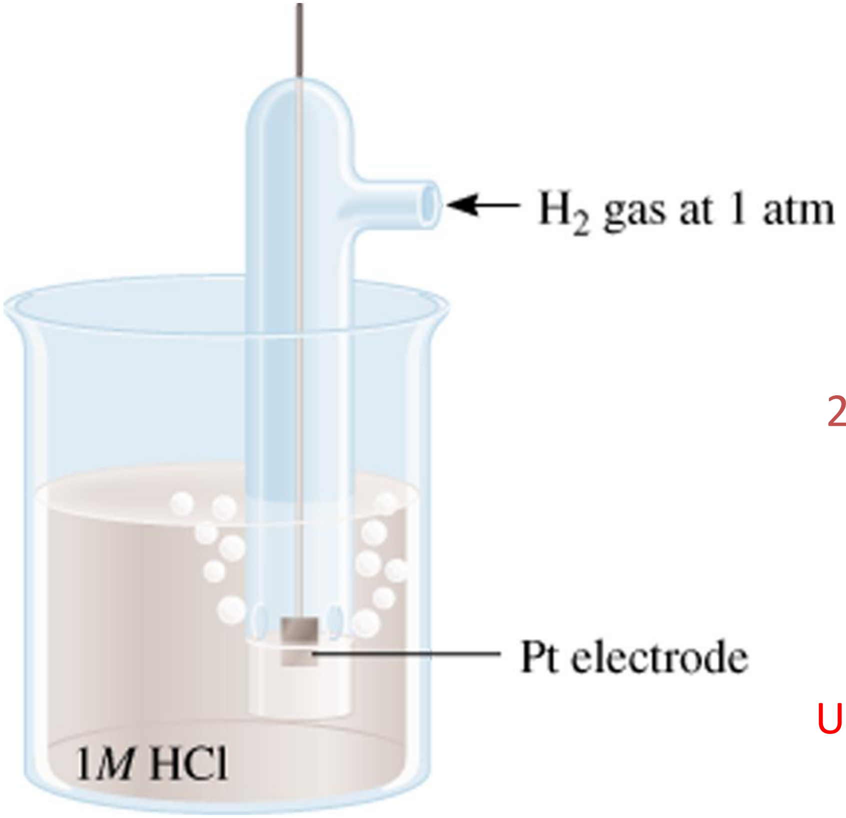
anoda (oxidation): Zn (s) $\longrightarrow$ Zn^{2+} (1 M) + $2e^-$

katoda (reduksi): $2e^-$ + $2H^+$ (1 M) $\longrightarrow$ H_2 (1 atm)

Zn (s) + $2H^+$ (1 M) $\longrightarrow$ Zn^{2+} + H_2 (1 atm)

Standard Potential of Electroda

Potensial reduction standard (E^0) is voltage that relate with **reduction standard at** electroda if consentration all of solute matter 1 *M* and all of gas at 1 atm.



hydrogen standard Electroda

reduction reaction

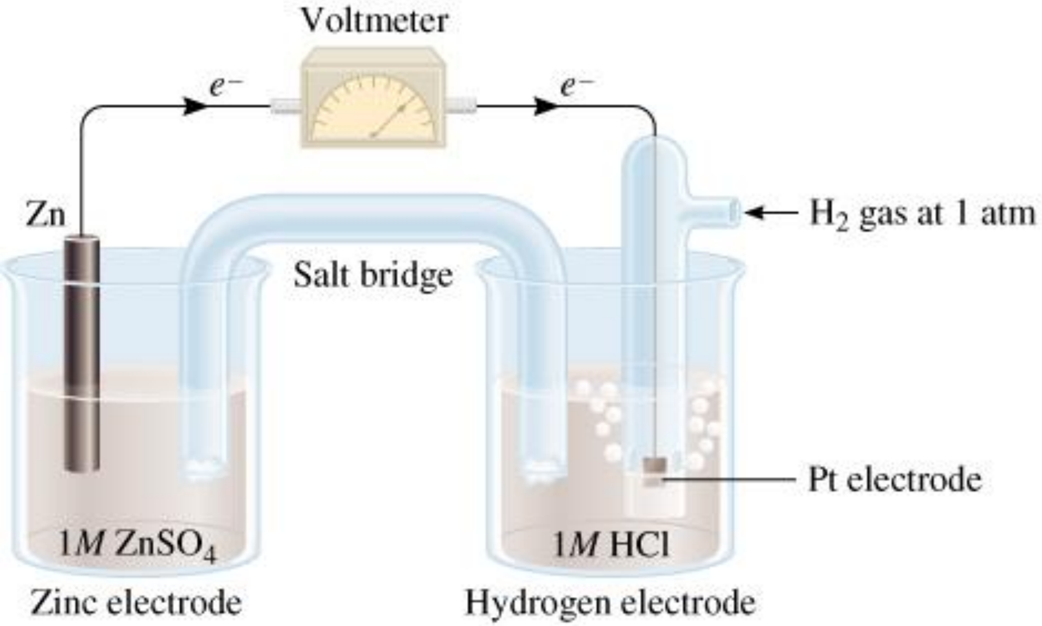
$$2e^- + 2H^+ (1\ M) \longrightarrow H_2\ (1\ atm)$$

$$\boxed{E^0 = 0\ V}$$

Use ase as reference to measure potential other matter

Potensial Elektroda Standar

$E^0_{sel} = 0,76\ V$



emf standar (E^0_{sel})

$$\boxed{E^0_{sel} = E^0_{katoda} - E^0_{anoda}}$$

reduction

oxidaton

Zn (*s*) | Zn^{2+} (1 *M*) || H^+ (1 *M*) | H_2 (1 atm) | Pt (*s*)

$$E^0_{sel} = E^0_{H^+/H_2} - E^0_{Zn^{2+}/Zn}$$

$$0.76\ V = 0 - E^0_{Zn^{2+}/Zn}$$

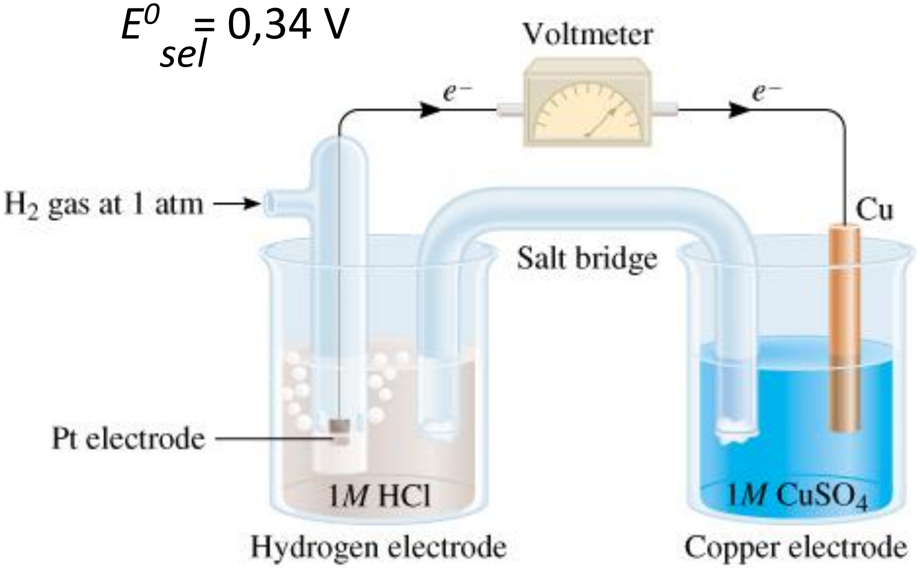
$$E^0_{Zn^{2+}/Zn} = -0,76\ V$$

$$Zn^{2+}\ (1\ M) + 2e^- \longrightarrow Zn \qquad E^0 = -0,76\ V$$

$$\boxed{E° > 0 \text{ reaksi spontan}}$$

Standard Potential Electroda

$E^0_{sel} = 0{,}34$ V



$E^0_{sel} = E_{katoda}{}^0 - E_{anoda}{}^0$

$E_{sel}{}^0 = E_{Cu}{}^0{}_{/Cu^{2+}} - E_{H}{}^0{}_{/H^+}$

$0{,}34 = E_{Cu}{}^0{}_{/Cu^{2+}} - 0$

$E_{Cu}{}^0{}_{/Cu2+} = 0{,}34$ V

Pt (*s*) | H_2 (1 *atm*) | H^+ (1 *M*) || Cu^{2+} (1 *M*) | Cu (*s*)

anoda (oxidation): $H_2\ (1\ atm) \longrightarrow 2H^+\ (1\ M) + 2e^-$

cathoda (reduction): $2e^- + Cu^{2+}\ (1\ M) \longrightarrow Cu\ (s)$

$H_2\ (1\ atm) + Cu^{2+}\ (1\ M) \longrightarrow Cu\ (s) + 2H^+\ (1\ M)$

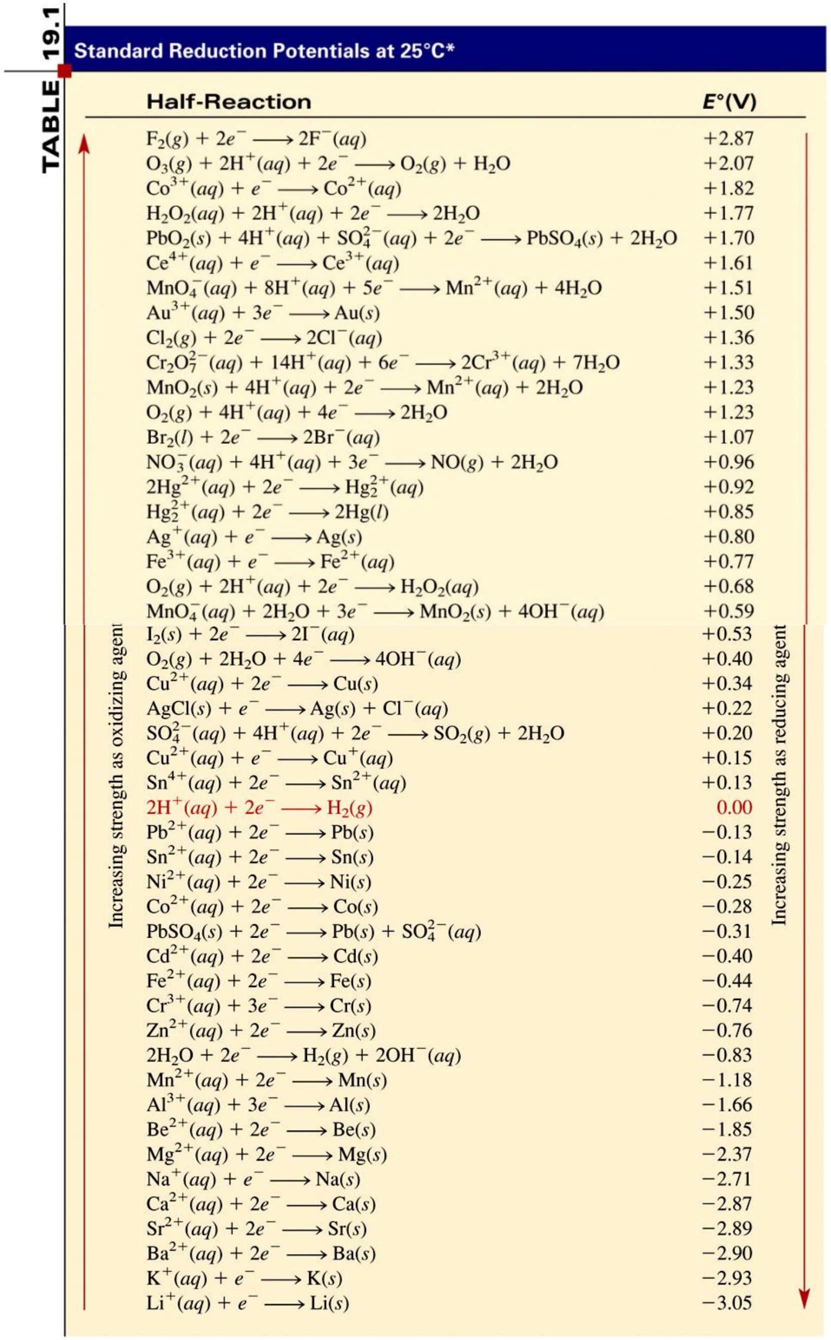
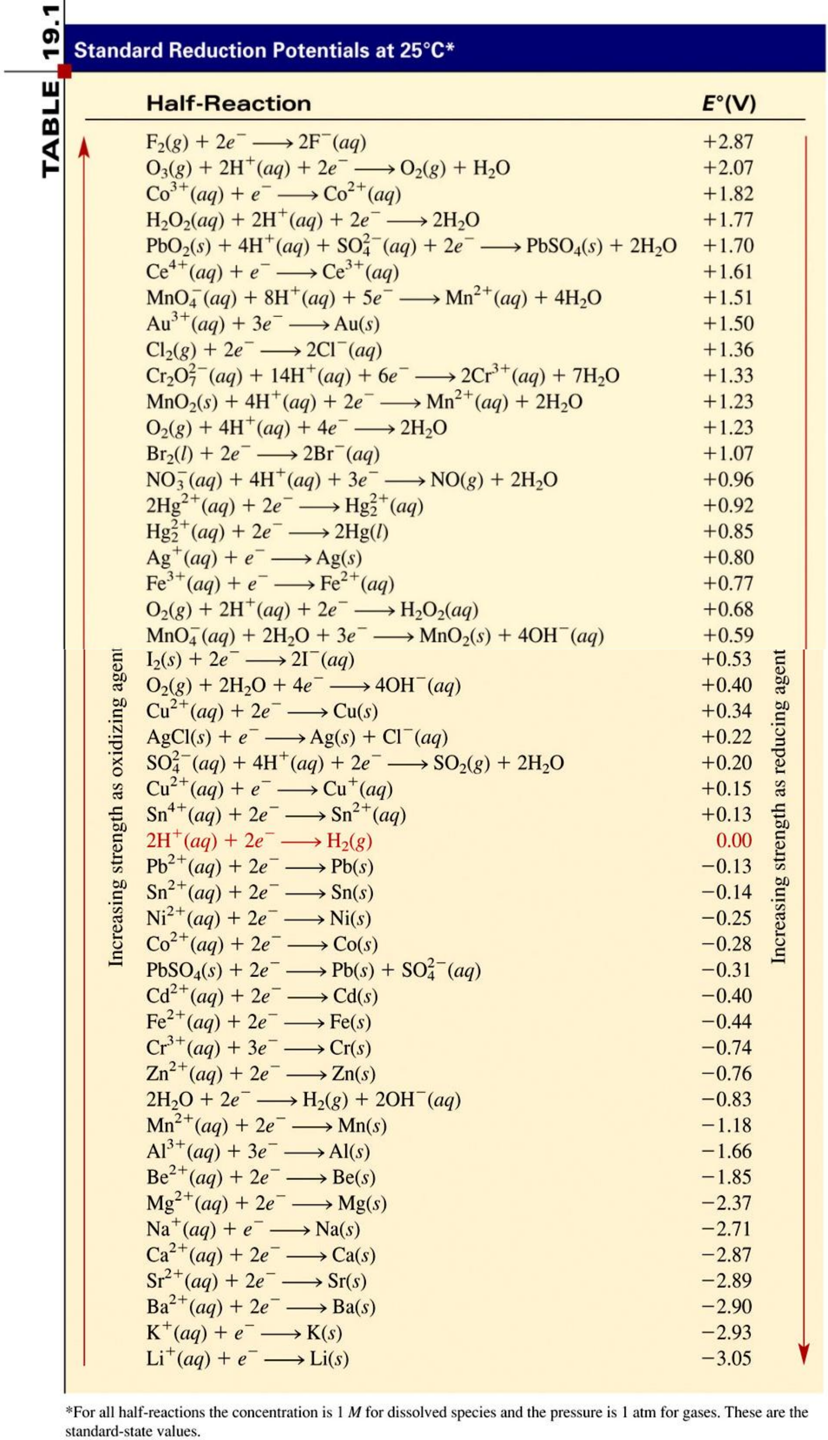


TABLE 19.1 Standard Reduction Potentials at 25°C*

Half-Reaction	E°(V)
$F_2(g) + 2e^- \longrightarrow 2F^-(aq)$	+2.87
$O_3(g) + 2H^+(aq) + 2e^- \longrightarrow O_2(g) + H_2O$	+2.07
$Co^{3+}(aq) + e^- \longrightarrow Co^{2+}(aq)$	+1.82
$H_2O_2(aq) + 2H^+(aq) + 2e^- \longrightarrow 2H_2O$	+1.77
$PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \longrightarrow PbSO_4(s) + 2H_2O$	+1.70
$Ce^{4+}(aq) + e^- \longrightarrow Ce^{3+}(aq)$	+1.61
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 4H_2O$	+1.51
$Au^{3+}(aq) + 3e^- \longrightarrow Au(s)$	+1.50
$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$	+1.36
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O$	+1.33
$MnO_2(s) + 4H^+(aq) + 2e^- \longrightarrow Mn^{2+}(aq) + 2H_2O$	+1.23
$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O$	+1.23
$Br_2(l) + 2e^- \longrightarrow 2Br^-(aq)$	+1.07
$NO_3^-(aq) + 4H^+(aq) + 3e^- \longrightarrow NO(g) + 2H_2O$	+0.96
$2Hg^{2+}(aq) + 2e^- \longrightarrow Hg_2^{2+}(aq)$	+0.92
$Hg_2^{2+}(aq) + 2e^- \longrightarrow 2Hg(l)$	+0.85
$Ag^+(aq) + e^- \longrightarrow Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \longrightarrow H_2O_2(aq)$	+0.68
$MnO_4^-(aq) + 2H_2O + 3e^- \longrightarrow MnO_2(s) + 4OH^-(aq)$	+0.59
$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$	+0.53
$O_2(g) + 2H_2O + 4e^- \longrightarrow 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	+0.34
$AgCl(s) + e^- \longrightarrow Ag(s) + Cl^-(aq)$	+0.22
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \longrightarrow SO_2(g) + 2H_2O$	+0.20
$Cu^{2+}(aq) + e^- \longrightarrow Cu^+(aq)$	+0.15
$Sn^{4+}(aq) + 2e^- \longrightarrow Sn^{2+}(aq)$	+0.13
$2H^+(aq) + 2e^- \longrightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \longrightarrow Sn(s)$	−0.14
$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$	−0.25
$Co^{2+}(aq) + 2e^- \longrightarrow Co(s)$	−0.28
$PbSO_4(s) + 2e^- \longrightarrow Pb(s) + SO_4^{2-}(aq)$	−0.31
$Cd^{2+}(aq) + 2e^- \longrightarrow Cd(s)$	−0.40
$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	−0.44
$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$	−0.74
$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$	−0.76
$2H_2O + 2e^- \longrightarrow H_2(g) + 2OH^-(aq)$	−0.83
$Mn^{2+}(aq) + 2e^- \longrightarrow Mn(s)$	−1.18
$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$	−1.66
$Be^{2+}(aq) + 2e^- \longrightarrow Be(s)$	−1.85
$Mg^{2+}(aq) + 2e^- \longrightarrow Mg(s)$	−2.37
$Na^+(aq) + e^- \longrightarrow Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \longrightarrow Ca(s)$	−2.87
$Sr^{2+}(aq) + 2e^- \longrightarrow Sr(s)$	−2.89
$Ba^{2+}(aq) + 2e^- \longrightarrow Ba(s)$	−2.90
$K^+(aq) + e^- \longrightarrow K(s)$	−2.93
$Li^+(aq) + e^- \longrightarrow Li(s)$	−3.05

*For all half-reactions the concentration is 1 *M* for dissolved species and the pressure is 1 atm for gases. These are the standard-state values.

- E^0 is for writen reaction
- A half reaction is reversible reaction
- Sign E^0 *change if direction of* reaction was reverse($E°_{red} = -E°_{oks}$)
- Change coefisien stoichiometric of half-sel reaction **not** change value E^0



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*For all half-reactions the concentration is 1 *M* for dissolved species and the pressure is 1 atm for gases. These are the standard-state values.

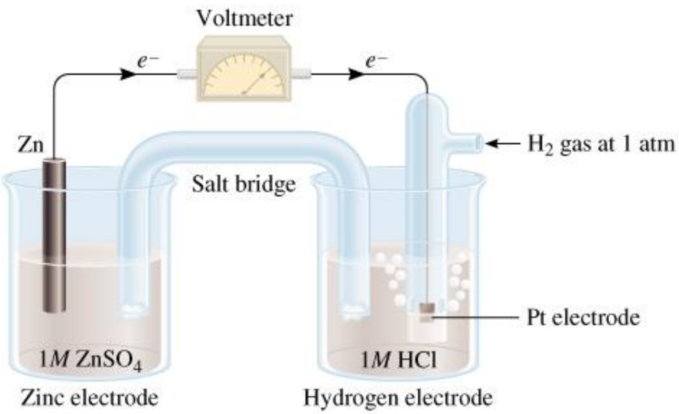
Strongest Oxidizing Agent

- makin positif E^0 makin besar kecendrungan suatu zat mengalami reduksi

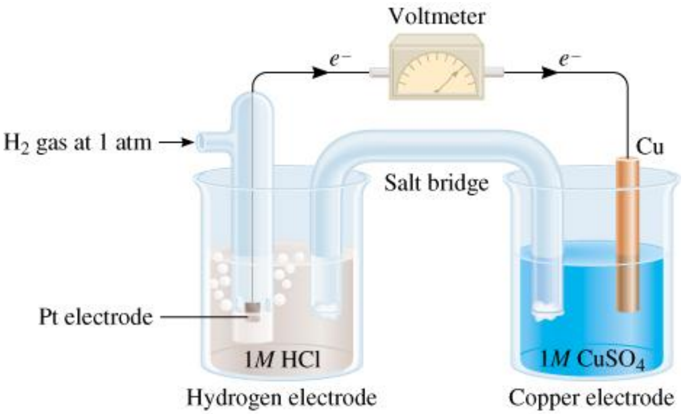
Zerro Refference point

Strongest Reducing Agent

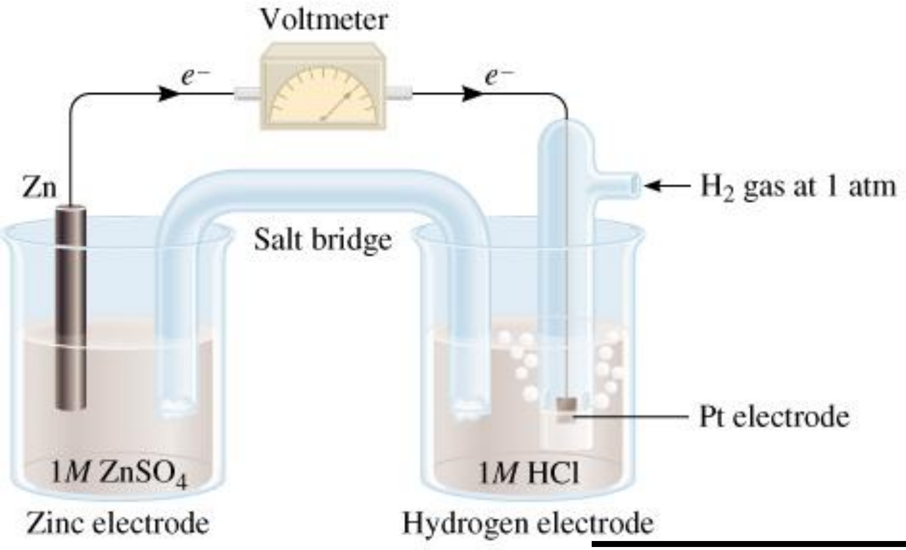
$E^0_{sel} = 0{,}76\ V$



$E^0_{sel} = 0{,}34\ V$



Combine ...!!!

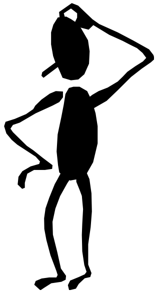


$$Zn\ (s) \longrightarrow Zn^{2+}\ (1\ M) + 2e^-$$

$$2e^- + Cu^{2+}\ (1\ M) \longrightarrow Cu\ (s)$$

$$Zn\ (s) + Cu^{2+}\ (1\ M) \longrightarrow Cu\ (s) + Zn^{2+}\ (1\ M)$$

$E^0_{sel} = 0{,}76\ V + 0{,}34\ V = 1{,}10\ V$



How many emf standard of electrohemistry cell with Cd electroda in the 1,0 *M* $Cd(NO_3)_2$ and electroda Cr in the 1,0 *M* $Cr(NO_3)_3$?

$Cd^{2+}(aq) + 2e^- \longrightarrow Cd(s)$ $E^0 = -0{,}40$ V

Cd oxidator

$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$ $E^0 = -0{,}74$ V

Cd will oxidize Cr

anoda (oxidation): $Cr(s) \longrightarrow Cr^{3+}(1\ M) + 3e^-$ x 2

cathoda (reduction): $2e^- + Cd^{2+}(1\ M) \longrightarrow Cd(s)$ x 3

$2Cr(s) + 3Cd^{2+}(1\ M) \longrightarrow 3Cd(s) + 2Cr^{3+}(1\ M)$

$E^0_{sel} = E_{katoda}^0 - E_{anoda}$ 0

$E^0_{sel} = -0{,}40 - (-0{,}74)$

$E^0_{sel} = 0{,}34$ V $\Rightarrow$ **spontan**

Spontanity of Redox reaction

sponanity $\Delta G < 0$

energy = $Q\,E = -\underbrace{nF}_{\text{Netto of charge}}E_{sel}$

$\Delta G = -nFE_{sel}$ n = number of mol electron in the reaction

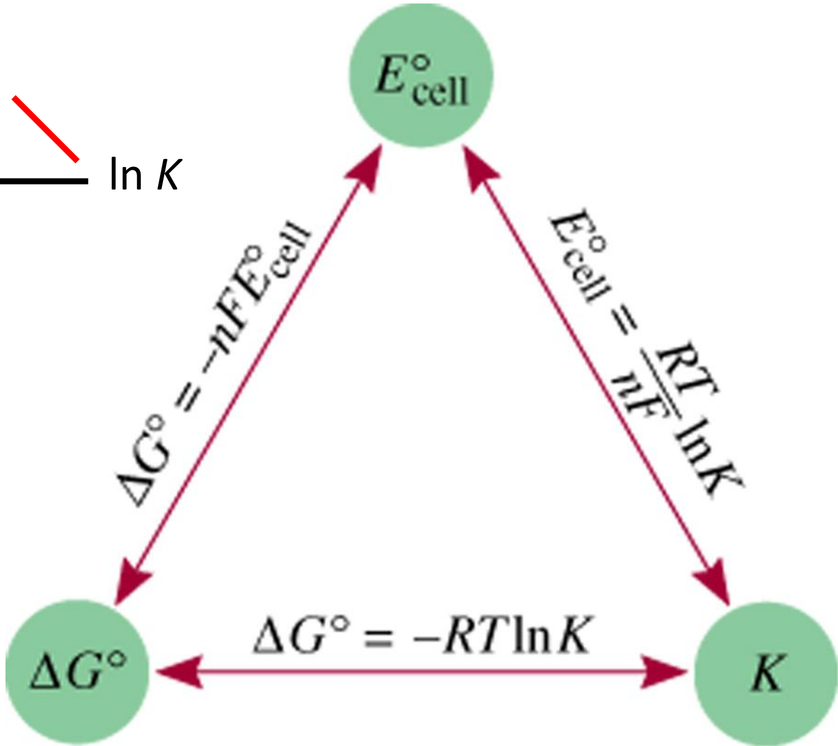
$\Delta G^0 = -nFE_{sel}\ 0$ $F = 96.500 \dfrac{J}{V \bullet mol} = 96.500\ C/mol$

$\Delta G^0 = -RT \ln K$ $= -nFE^0_{sel}$

$$E^0_{sel} = \frac{RT}{nF} \ln K = \frac{(8{,}314\ \cancel{J}/\cancel{K}\bullet\cancel{mol})(298\ \cancel{K})}{n\,(96.500\ \cancel{J}/V\bullet\cancel{mol})} \ln K$$

$$E^0_{sel} = \frac{0{,}0257\ V}{n} \ln K$$

$$E^0_{sel} = \frac{0{,}0592\ V}{n} \log K$$



spontanity of Redoxon react

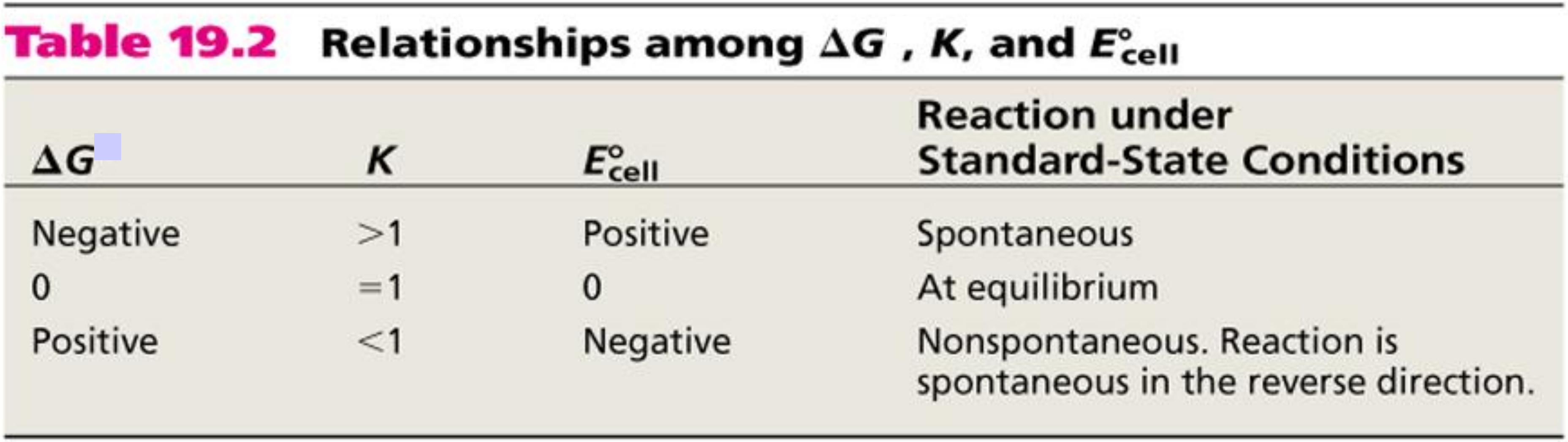
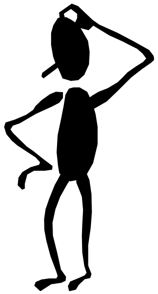


Table 19.2 Relationships among ΔG , K, and E°_{cell}

ΔG	K	E°_{cell}	Reaction under Standard-State Conditions
Negative	>1	Positive	Spontaneous
0	$=1$	0	At equilibrium
Positive	<1	Negative	Nonspontaneous. Reaction is spontaneous in the reverse direction.

$$\Delta G^0 = -RT \ln K$$

$$\Delta G = -nFE_{sel}$$



How many of equilibrium constant for this reacteon at 25^{0}C?

$$Fe^{2+} (aq) + 2Ag (s) \rightleftharpoons Fe (s) + 2Ag^+ (aq)$$

$$E^0_{sel} = \frac{0,0257 \text{ V}}{n} \ln K$$

Oxidatio: $2Ag \rightarrow 2Ag^+ + 2e^-$

reduction: $2e^- + Fe^{2+} \rightarrow Fe$

$n = 2$

$E^0 = E_{Fe^{2+}/Fo} - E_{Ag\ /Ag}^{0}\ _{+}$

0

$E^0 = -0,44 - (0,80)$

$E^0 = -1,24$ V

$$K = \exp\left[\frac{E^0_{sel} \times n}{0,0257 \text{ V}}\right] = \exp\left[\frac{-1,24 \text{ V} \times 2}{0,0257 \text{ V}}\right]$$

$K = 1,23 \times 10^{-42}$

The Efect of consentration to Emf cell

$\Delta G = \Delta G^0 + RT \ln Q$ $\Delta G = -nFE$ $\Delta G^0 = -nFE^0$

$$-nFE = -nFE^0 + RT \ln Q$$

Nernst equation

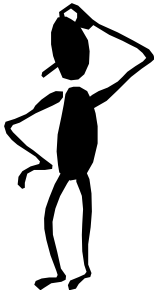
$$E = E^0 - \frac{RT}{nF} \ln Q$$

Can be see at the efect of consentration / non-standar condition

at 298 K

$$E = E^0 - \frac{0{,}0257\ \text{V}}{n} \ln Q$$

$$E = E^0 - \frac{0{,}0592\ \text{V}}{n} \log Q$$



Is this reaction will occur spontanity at 25⁰C if $[Fe^{2+}] = 0,60\ M$ and $[Cd^{2+}] = 0,010\ M$?

$Fe^{2+}(aq) + Cd(s) \rightleftharpoons Fe(s) + Cd^{2+}(aq)$

oxidation: $Cd \longrightarrow Cd^{2+} + 2e^-$

$n = 2$

reduction: $2e^- + Fe^{2+} \longrightarrow 2Fe$

$E^0 = E^0_{Fe^{2+}/Fe} - E^0_{Cd^{2+}/Cd}$

$E^0 = -0,44 - (-0,40)$

$E^0 = -0,04\ V$

$$E = E^0 - \frac{0,0257\ V}{n} \ln Q$$

$$E = -0,04\ V - \frac{0,0257\ V}{2} \ln \frac{0,010}{0,60}$$

$E = 0,013\ V$

$E > 0$ spontan

Consentration cell half the same half-reaction in each cell, bu diffrent in the consentration.

Is this reaction will occur spontan at 25⁰C if $[Ag^+] = 0{,}10\ M$ and $[Ag^+] = 0{,}010\ M$?

$Ag^+\ (aq) + Ag\ (s) \rightleftharpoons Ag\ (s) + Ag^+\ (aq)$

$Ag\ (s)\ |\ Ag^+\ (0.10\ M)\ ||\ Ag^+\ (0.010\ M)\ |\ Ag\ (s)$

oksidasi: $Ag \longrightarrow Ag^+ + e^-$

reduksi: $e^- + Ag^+ \longrightarrow Ag$

$n = 1$

$Ag_{oks} + Ag^+_{red} \longrightarrow Ag_{red} + Ag^+_{oks}$

$E^0 = E^0_{Ag^+/Ag} - E^0_{Ag^+/Ag}$

$E^0 = -0{,}7991\ V - (-0{,}7991\ V)$

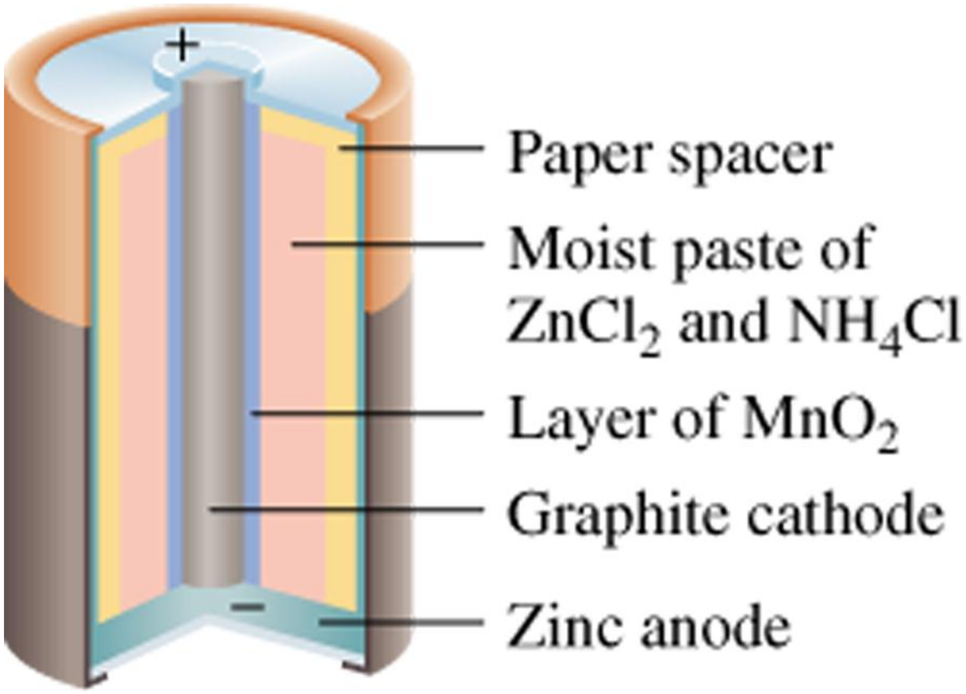
$E^0 = -0{,}000\ V$

$$E = E^0 - \frac{0{,}0257\ V}{n} \ln Q$$

$E < 0$ **Non-spontan**

$$E = -0.000\ V - \frac{0{,}0257\ V}{1} \ln \frac{[Ag^+_{oks}]}{[Ag^+_{red}]} = -\ 0{,}0257\ V \quad \ln \frac{0{,}10}{0{,}010} = -\ 0{,}0592\ V$$

Battery



Dry cell

Sel Leclanché

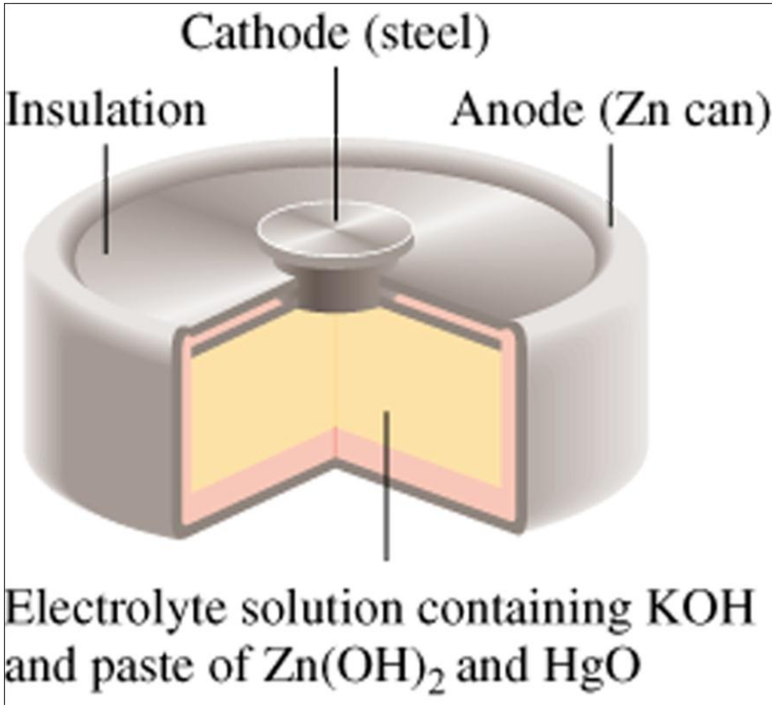
anoda: $Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2e^-$

katoda: $^+$ $2NH_4\ (aq) + 2MnO_2\ (s) + 2e \rightarrow Mn_2O_3\ (s) + 2NH_3\ (aq) + H_2O\ (l)$

$Zn\ (s) + 2NH_4\ (aq) + 2MnO_2\ (s) \rightarrow Zn^{2+} (aq) + 2NH_3\ (aq) + H_2O\ (l) + Mn_2O_3\ (s)$

Battery

Merkuri Battery



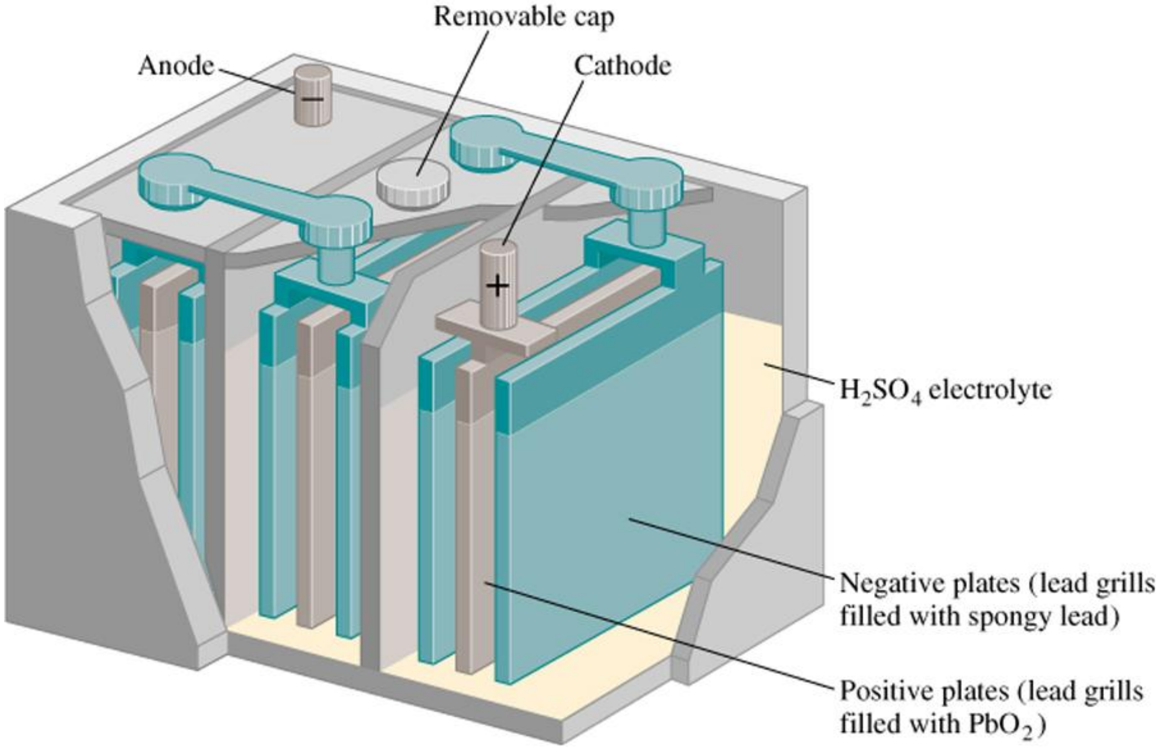
anoda: $Zn(Hg) + 2OH^- (aq) \longrightarrow ZnO (s) + H_2O (l) + 2e^-$

cathoda: $HgO (s) + H_2O (l) + 2e^- \longrightarrow Hg (l) + 2OH^- (aq)$

$Zn(Hg) + HgO (s) \longrightarrow ZnO (s) + Hg (l)$

Battery

Lead Battery
(Aki)

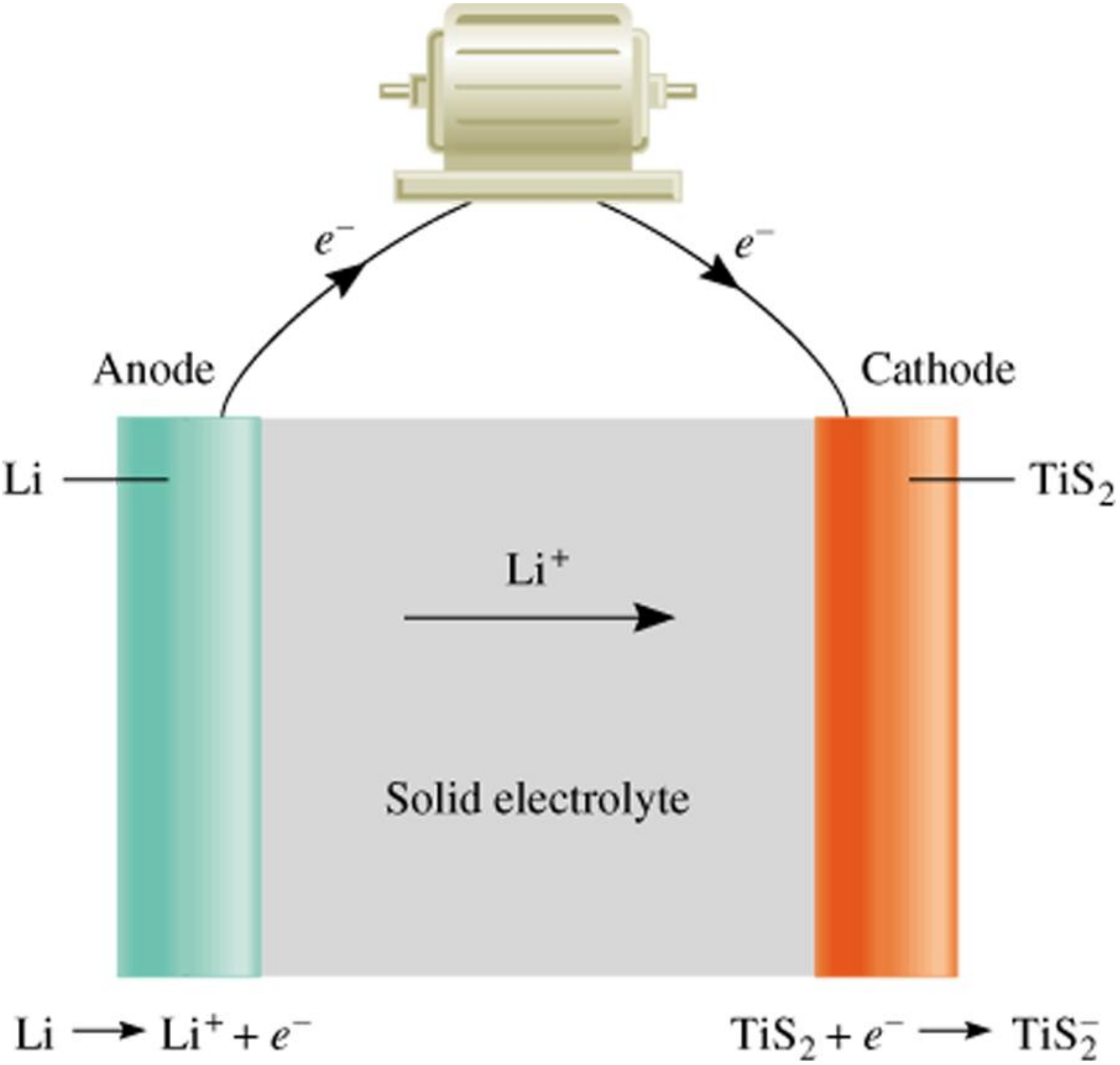


anoda: $Pb\ (s) + SO_4^{2-}\ (aq) \longrightarrow PbSO_4\ (s) + 2e^-$

cathoda: $PbO_2\ (s) + 4H^+\ (aq) + SO_4^{2-}\ (aq) + 2e^- \longrightarrow PbSO_4\ (s) + 2H_2O\ (l)$

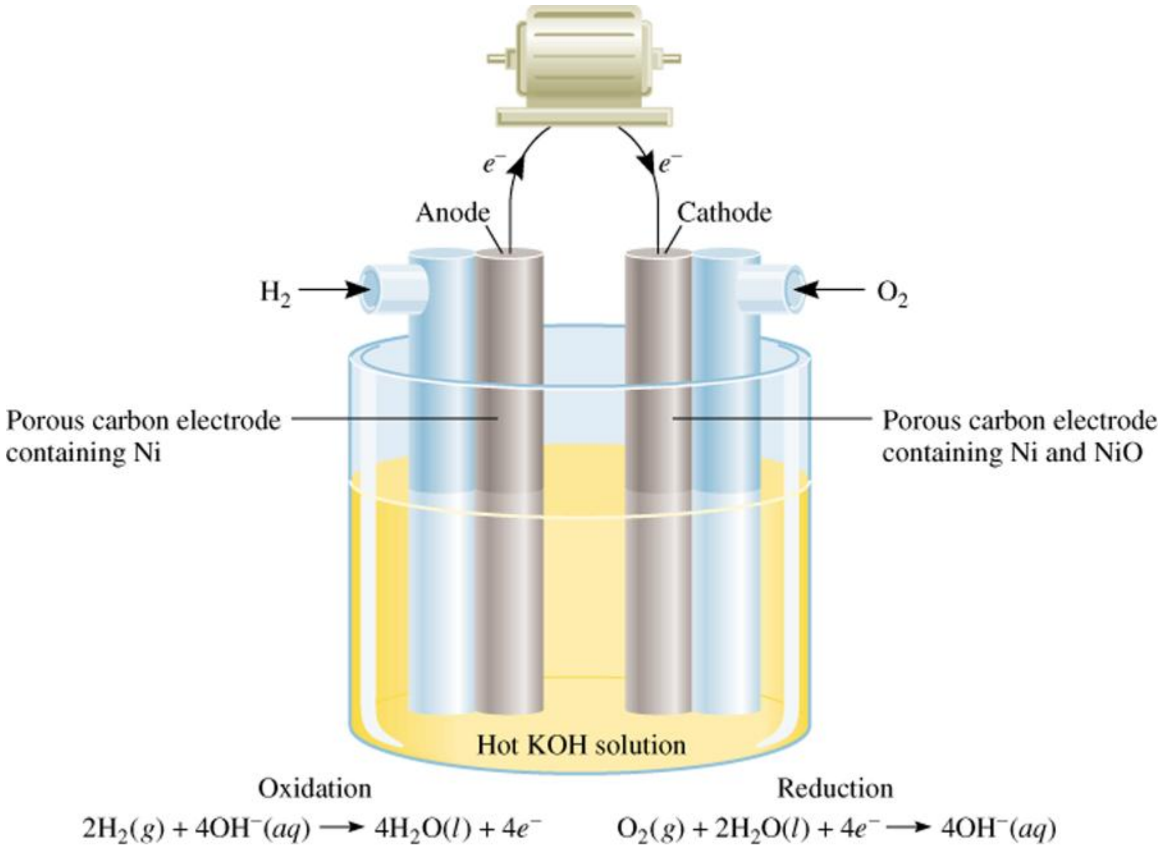
$Pb\ (s) + PbO_2\ (s) + 4H^+\ (aq) + 2SO^{2-}\ (aq_4) \longrightarrow 2PbSO_4s) + 2H_2O\ (l)$

Battery



Solid- Lithium Battery

Battery



Fuel Cell is electrochemistry cell that need continue reactan to can be use

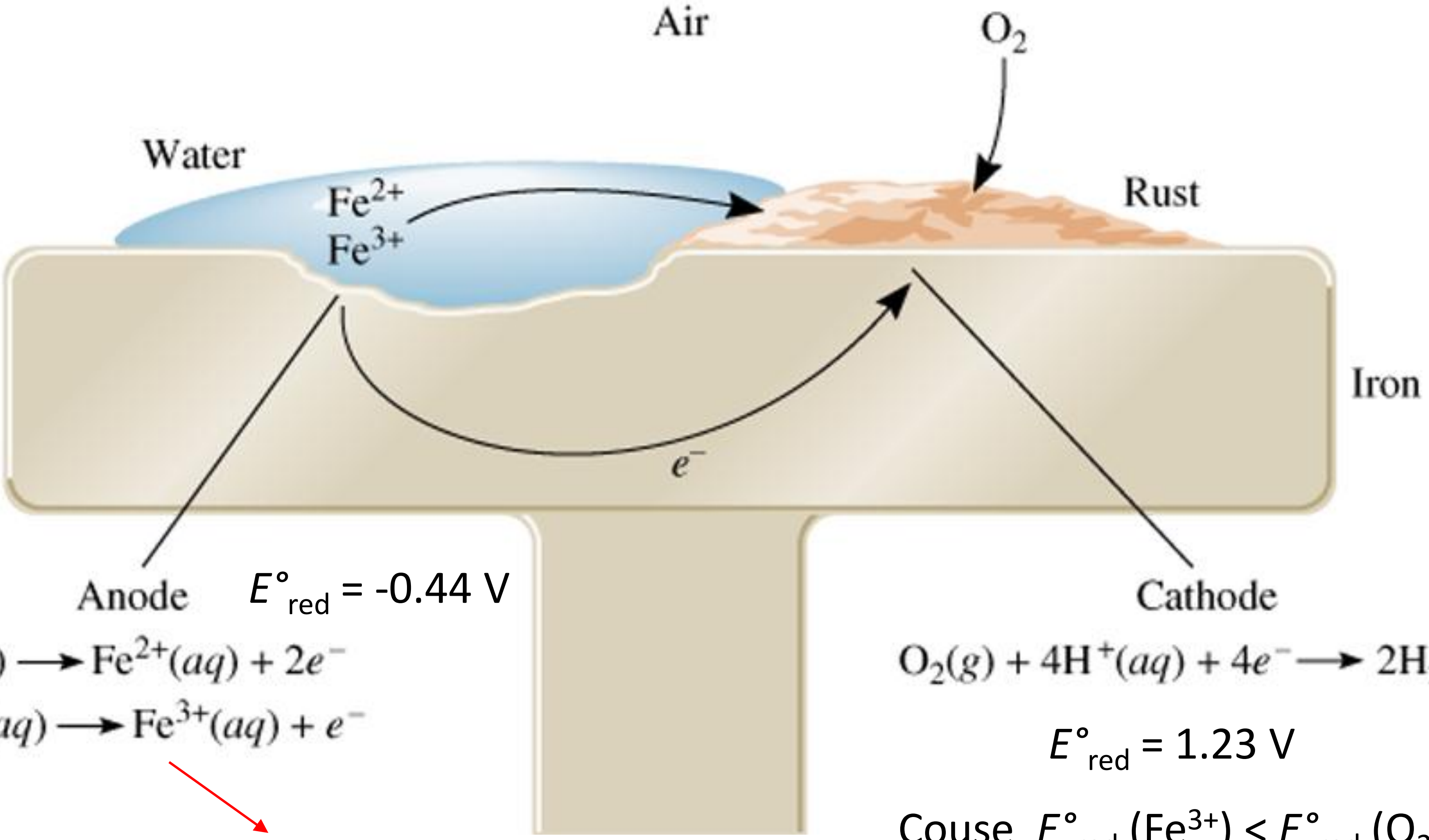
anoda: $2H_2\ (g) + 4OH^-\ (aq) \longrightarrow 4H_2O\ (l) + 4e^-$

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

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

Corrosive

Oxygen dilute in the water cause oxidation



$E°_{red}$ = -0.44 V

$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$

$Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^-$

Corrotion: Fe_2O_3

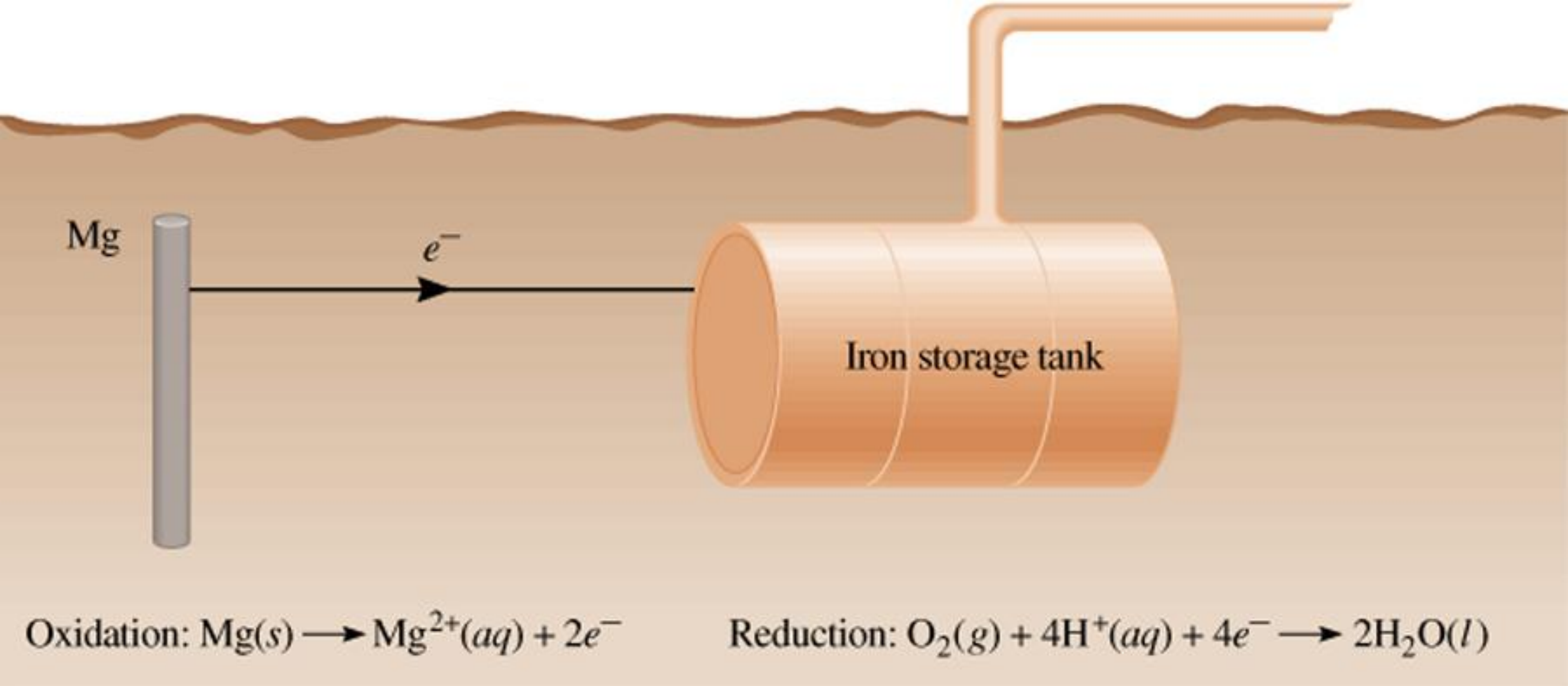
$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$

$E°_{red}$ = 1.23 V

Couse $E°_{red}$ $(Fe^{3+}) < E°_{red} (O_2)$

Fe can be oxydized by oxygen

Perlindungan Katodik Tangki Besi



Oxidation: $Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$

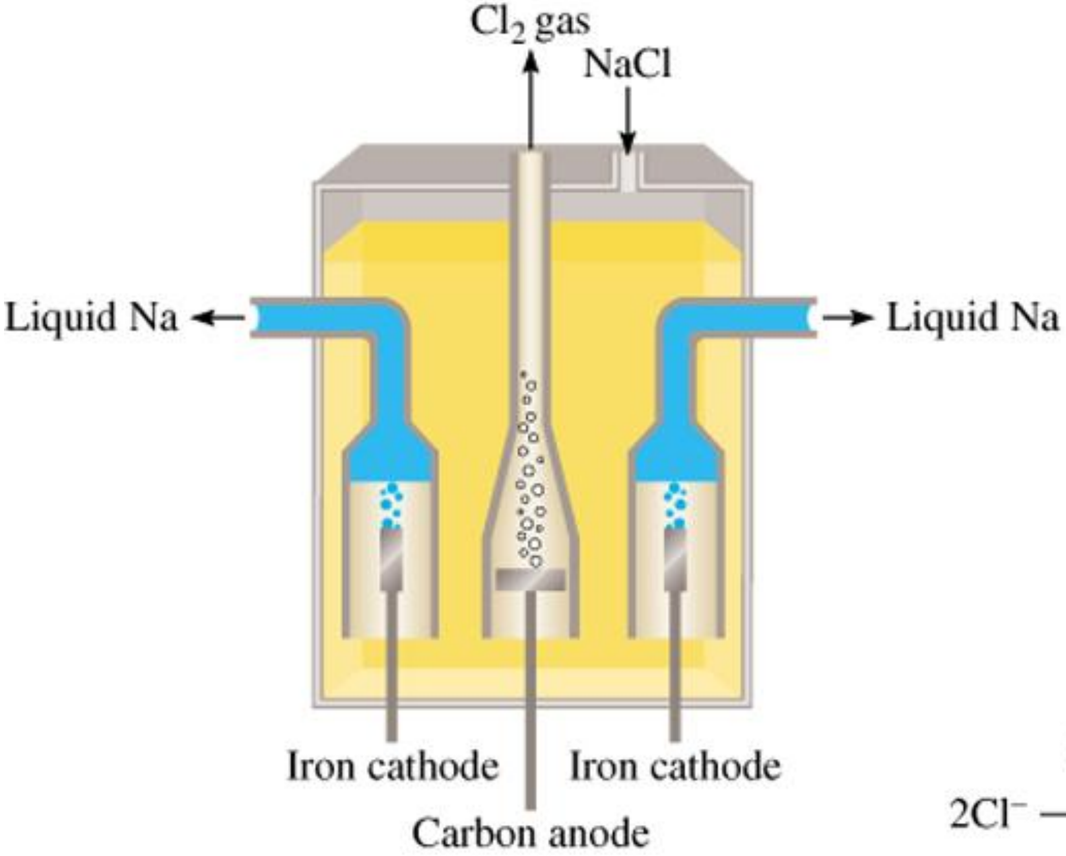
Reduction: $O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$

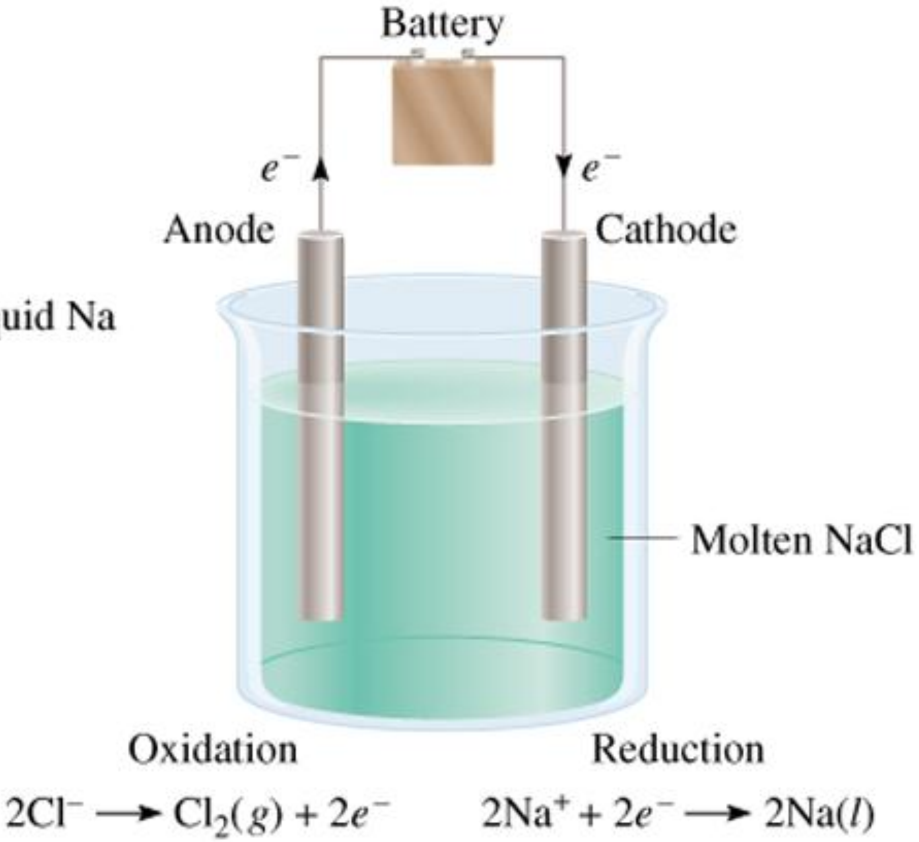
$E°_{red}$ = -2.37 V

$E°_{red}$ = 1.23 V

Mg more easy to Oxidize than Fe

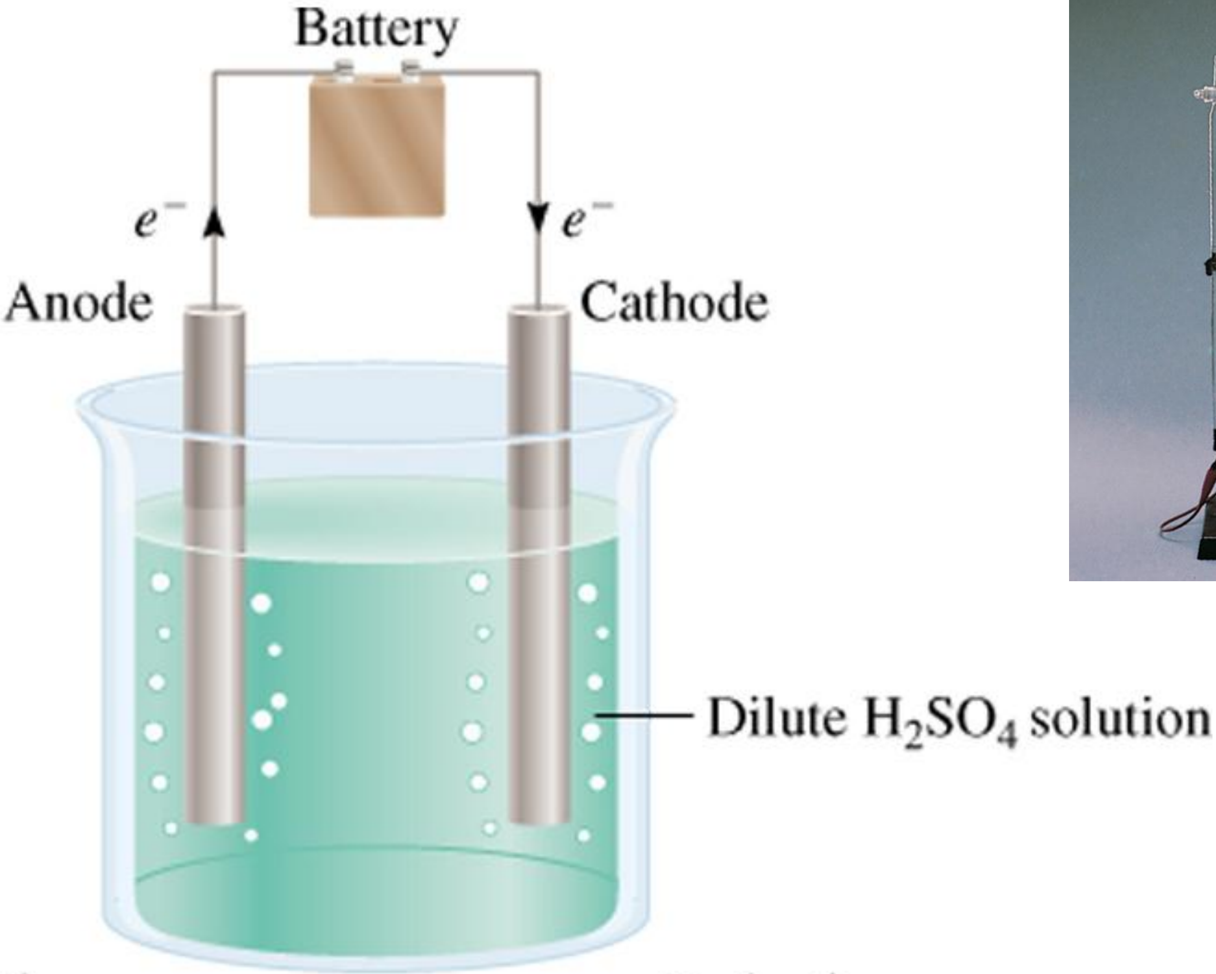
Electrolysis is a process where electric energy be used so that a **nonspontan** chemical reaction can occur.





19.8

Water Elektrolysis



Oxidation

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

Reduction

$4H^+(aq) + 4e^- \longrightarrow 2H_2(g)$

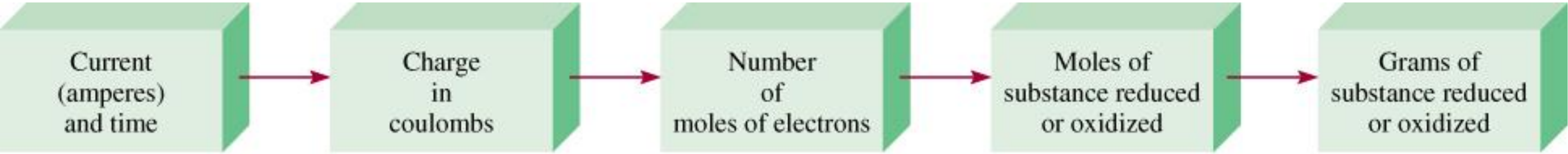
Electrolysis and Massa change

Kuantitatif analysis

How much current/ampere?

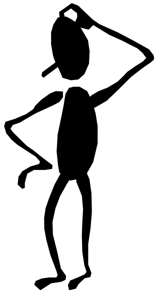
time ?

product?



Charge (C) = Ampere (A) x time (s)

1 mol e^- = 96.500 C



How much Ca that be resulted from electrolytic cell from molten $CaCl_2$ if there are current 0,452 A was flowed pass through cell for 1,5 hours?

anoda: $2Cl^- (l) \longrightarrow Cl_2 (g) + 2e^-$

cathoda: $Ca^{2+} (l) + 2e^- \longrightarrow Ca (s)$

$$Ca^{2+} (l) + 2Cl^- (l) \longrightarrow Ca (s) + Cl_2 (g)$$

2 mol e^- = 1 mol Ca

$$\text{mol Ca} = 0,452 \frac{\cancel{C}}{\cancel{s}} \times 1,5 \text{ jam} \times \cancel{3600} \frac{\cancel{s}}{\cancel{\text{jam}}} \times \frac{1 \cancel{\text{mol}}\ e^-}{96.500 \cancel{C}} \times \frac{1 \text{ mol Ca}}{2 \cancel{\text{mol}}\ e^-}$$

= 0,0126 mol Ca

= 0,50 g Ca

Stoichiometric from a electrolytic cell:

cathoda: $e^- + Ag^+ \longrightarrow Ag$

anoda: $2\ H_2O \longrightarrow O_2\ (g) + 4\ H^+ + 4\ e^-$

How many current (ampere) that be needed to change 0,100 mol Ag^+ become Ag for 10,0 menit?

1 mol elektron = 1 F

A = C/second

Q = nF

I = Q / t

Cari Q

$$Q = 0,10\ \cancel{mol\ Ag}\ \frac{1\ \cancel{mol\ elektron}}{\cancel{mol\ Ag}}\ \frac{1\ \cancel{F}}{\cancel{mol\ elektron}}\ \frac{96.500\ C}{\cancel{F}} =$$

Q = 9.650 C

$$t = 10\ menit\ \frac{60\ second}{1\ menit} = 600\ second$$

I = 9.6500 C / 600 s = 16 C/s = 16 A

Stoichiometric Product in the diffrent Electroda

How much pH from anoda half-sel (exp: volume 0,100 L) after 6,00 g Ag was puted at the cathoda?

cathoda: $e^- + Ag^+ \longrightarrow Ag$

anoda: $2\ H_2O \longrightarrow O_2\,(g) + 4\ H^+ + 4\ e^-$

Look for [H⁺]

$$6{,}00\ \cancel{\text{g Ag}}\ \frac{1\ \cancel{\text{mol Ag}}}{\cancel{107{,}9\ \text{g Ag}}}\ \frac{1\ \cancel{\text{mol elektron}}}{1\ \cancel{\text{mol Ag}}}\ \frac{4\ \text{mol H}^+}{1\ \cancel{\text{mol elektron}}}$$

$$= 0{,}05567\ \text{mol H}^+$$

$$[H^+] = \frac{0{,}0556\ \text{mol}}{0{,}10\ \text{L}} = 0{,}56\ \text{M} \longrightarrow \boxed{\text{pH} = 0{,}25}$$