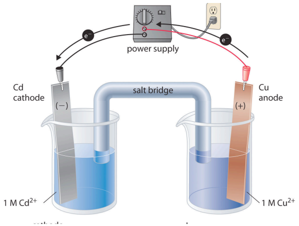
Wie können wir Spannung & Stromstärke bei galvanischen Zellen & Elektrolysen berechnen?

Physikalische Chemie by SciFox

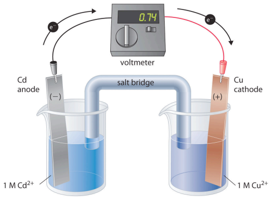
PHYSICAL CHEMISTRY BASICS
PART 12: ELECTRODES

Current and voltage in batteries and electrolysis



$$anode = (+)$$

?



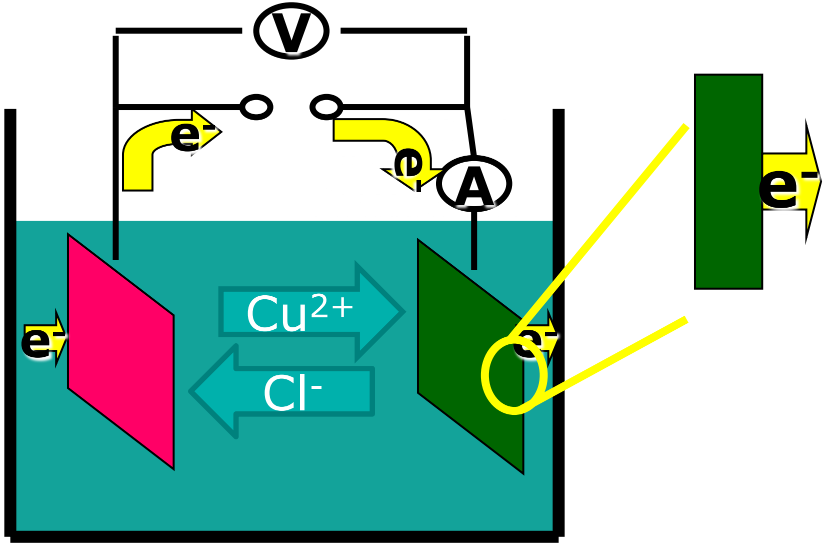
$$anode = (-)$$

$$Q = ?\ mAh$$
$$w = ?\ kWh$$

GALVANICCELL ⊕

Electrolysis ⊖

How does current flow **through** an electrolyte (I)?



#ElectrodeReaction

#Cathode

#Reduction

$$\nu_{Ox}\,[Ox] + \nu_e\, e^- \longrightarrow \nu_{Red}\,[Red] \quad (V-6.1)$$

$$Cu^{2+} + 2\,e^- \rightarrow Cu$$

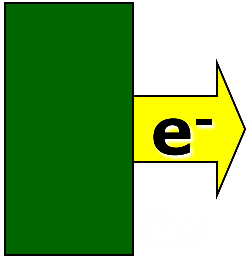
GALVANIC CELL ⊕

Electrolysis ⊖

#Cathode

#Reduction

What reaction happens at the **cathode**?



$$Cu^{2+} + 2\,e^- \rightarrow Cu$$

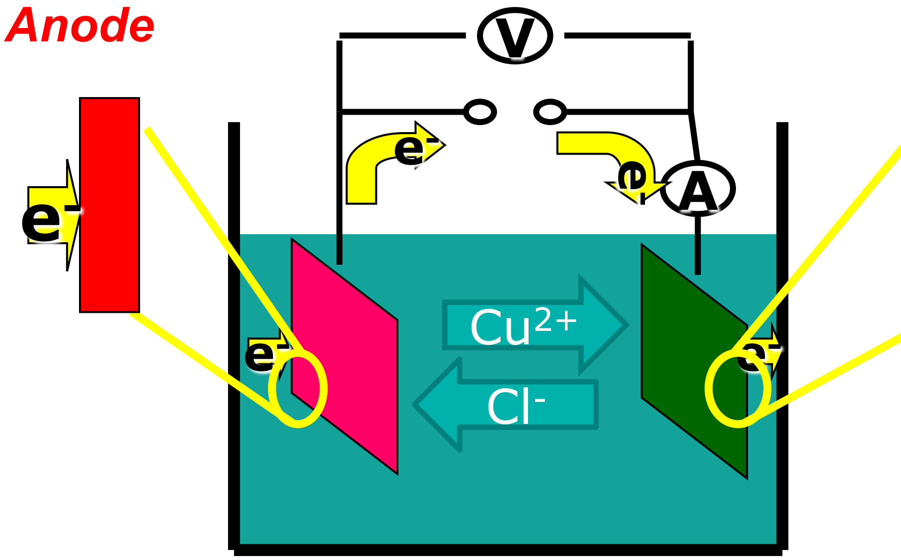
$(V - 6.7)$

$$4\,H^+ + PbO_2 + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2\,H_2O$$

GALVANIC CELL ⊖

Electrolysis ⊕

How does current flow **through** an electrolyte (II)?



$$\nu_{Ox}\,[\mathbf{Ox}] + \nu_e\,\mathbf{e^-} \longleftarrow \nu_{Red}\,[\mathbf{Red}]$$

$(V-6.1)$

$$Cl_2 + 2\,e^- \leftarrow 2\,Cl^-$$

#Anode

#Oxidation

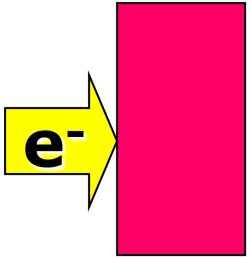
GALVANICCELL ⊖

Electrolysis ⊕

#Anode

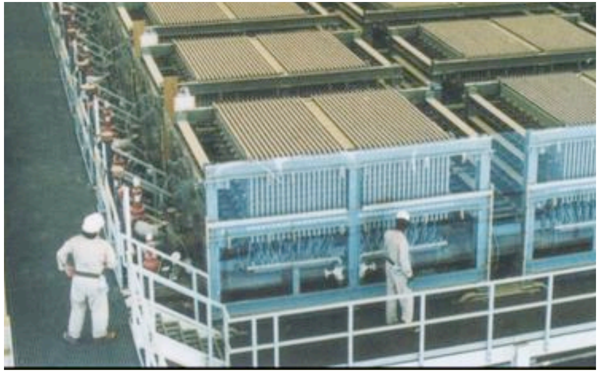
#Oxidation

What reaction happens at the **anode**?



$$Zn^{2+} + 2\,e^{-} \leftarrow Zn \quad (V-6.2)$$

$$Cl_2 + 2\,e^{-} \leftarrow 2\,Cl^{-} \quad (V-6.8)$$



GALVANIC CELL ⊕

Elektrolysis ⊖

#Cathode

#Reduction

GALVANIC CELL ⊖

Electrolysis ⊕

#Anode

#Oxidation

True or False?

A: The cathode is always the negative terminal

B: The reduction takes place at the cathode

C: The cathode is always at the more positive potential

D: At a cathode there is a net electron flow into the electrode

GALVANIC CELL ⊕

Elektrolysis ⊖

#Cathode

#Reduction

GALVANIC CELL ⊖

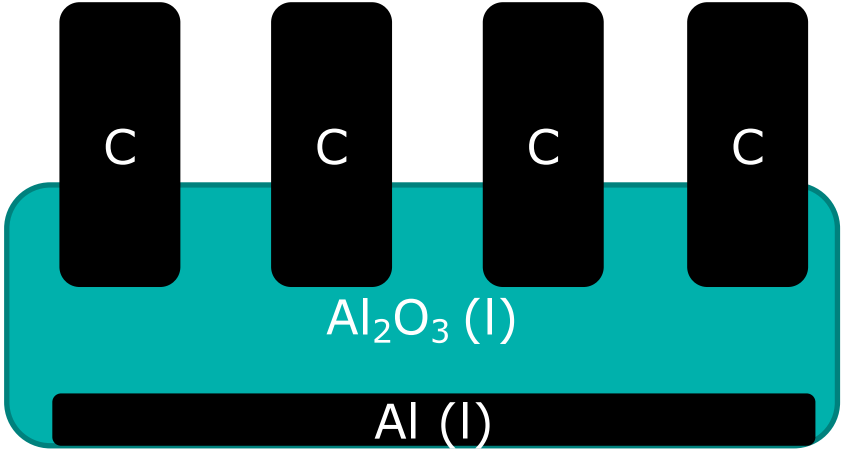
Electrolysis ⊕

#Anode

#Oxidation

Where is the **anode**? Where is the **positive terminal**?

$(V-6.2)$



$(V-6.3)$

M. Faraday
1791-1867

$F = 96\,485 \frac{\text{As}}{\text{mol}}$

What determines the **quantity of substances** produced or consumed at the electrodes?

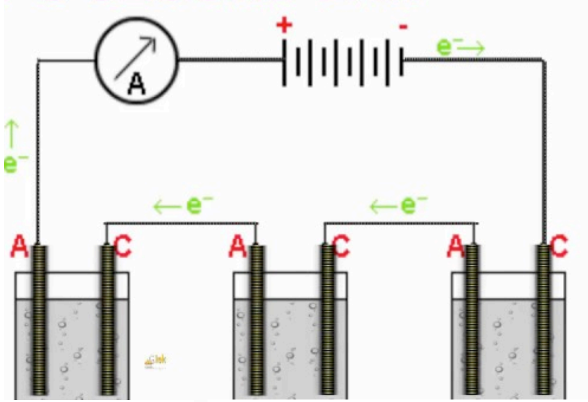
$$\nu_{Ox}\,[Ox] + \nu_e\, e^- \underset{i_{an}}{\overset{i_{cath}}{\rightleftarrows}} \nu_{Red}\,[Red] \qquad (V-6.1)$$

$$n = \frac{m}{M} = \frac{I \cdot t}{\nu_e\, F} \qquad (V-6.4)$$

$I = 1\,A$

$t = 1d$

$n = 0.9\,mol$



#Faraday'sLawOfElectrolysis

$Cu^{2+} + 2\,e^- \rightarrow Cu$ $(V-6.7)$

$Ag^+ + e^- \rightarrow Ag$ $(V-6.6)$

Wow much **mass** is **deposited** during electrolysis (1A, 1d)?

$$n\,\nu_e = \frac{m}{M}\nu_e = \frac{I\ \cdot t}{F}$$

$$= \frac{1\ A\ \cdot 86\,400\ s}{96\,485\ As/mol} = 0.9\ \text{mol} \qquad (V-6.5)$$

$$Cu^{2+} + 2\ e^- \rightarrow Cu \quad (V-6.7) \quad 28.5\ g$$

$$Ag^+ + \ e^- \rightarrow Ag \quad (V-6.6) \quad 96.5\ g$$

$$Al^{3+} + 3\ e^- \rightarrow Al \quad (V-6.3) \quad 8.1\ g$$

$$Cl_2 + 2\ e^- \leftarrow 2\ Cl^- \quad (V-6.8) \quad 31.7\ g$$

How do we measure **electric charge**?

Electrolysis of H_2SO_4

$\approx 100\ mA, \approx 1\ h:\ \frac{I\ t}{F} \approx 3.7\ mmol\ e^-$

$n_{H_2} = 2.0\ mmol$

$\boldsymbol{n_{e^-} = 4.0\ mmol}$

$\boldsymbol{Q = 386\ C}$

$$n_{H_2} = \frac{p_{H_2} V_{H_2}}{R\ T}$$

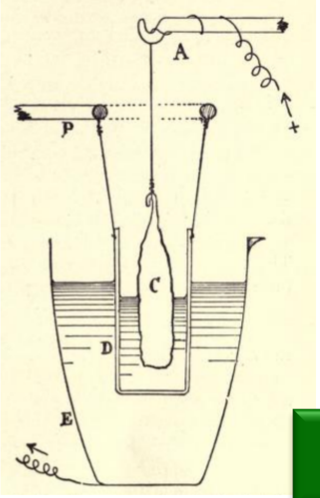
$$p_{H_2} = p_{ex} - p_{H_2O} \pm p_{hydro}$$

$$p_{hydro} = \rho\ g\ h$$

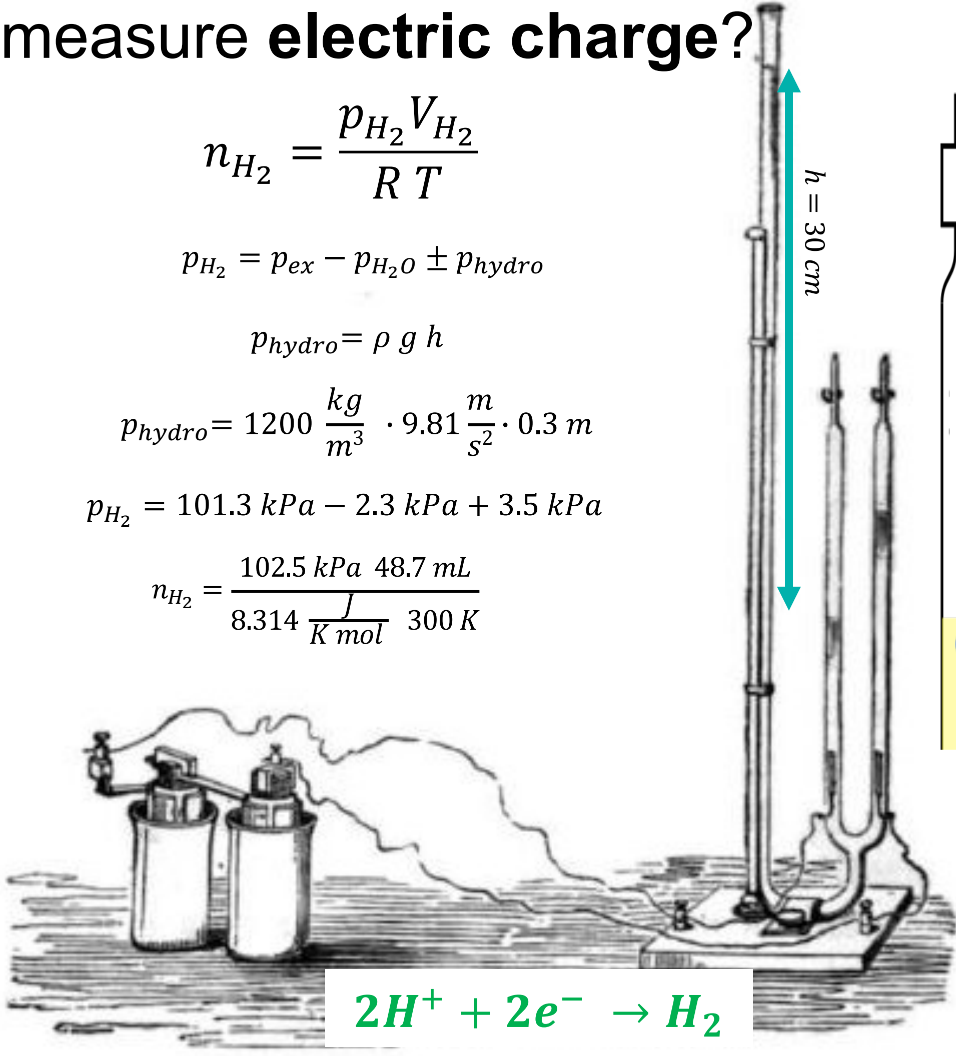
$$p_{hydro} = 1200\ \frac{kg}{m^3} \cdot 9.81\ \frac{m}{s^2} \cdot 0.3\ m$$

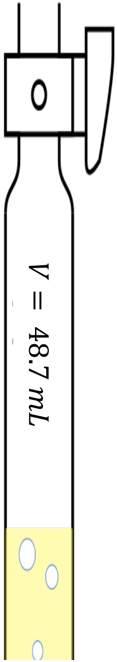
$$p_{H_2} = 101.3\ kPa - 2.3\ kPa + 3.5\ kPa$$

$$n_{H_2} = \frac{102.5\ kPa\ \ 48.7\ mL}{8.314\ \frac{J}{K\ mol}\ \ 300\ K}$$



#Electrometer





$\boldsymbol{2H^+ + 2e^- \rightarrow H_2}$

Discussion **a fuel cell**

6. 1 mol Wasserstoff wird in einer ideal arbeitenden Brennstoffzelle (PEMFC) zu flüssigem Wasser umgesetzt

- Wie viel elektrische Energie w_{el} und wie viel Wärme q werden dabei abgegeben?
- Berechnen Sie den (idealen) Wirkungsgrad η $(= \Delta G/\Delta H)$ dieser Brennstoffzelle.
- Die Zelle liefert einen Strom von 50 mA; Welches Volumen Wasserstoff V_{H2} (p = 1 bar, T = 300 K) wird an der Anode pro Stunde umgesetzt?

$$(\boldsymbol{V - 11.68 - 11.70})$$

PCÜ47 galvanische Zelle als reversible Redoxreaktion - Wirkungsgrad einer Brennstoffzelle

Physikalische Chemie by SciFox

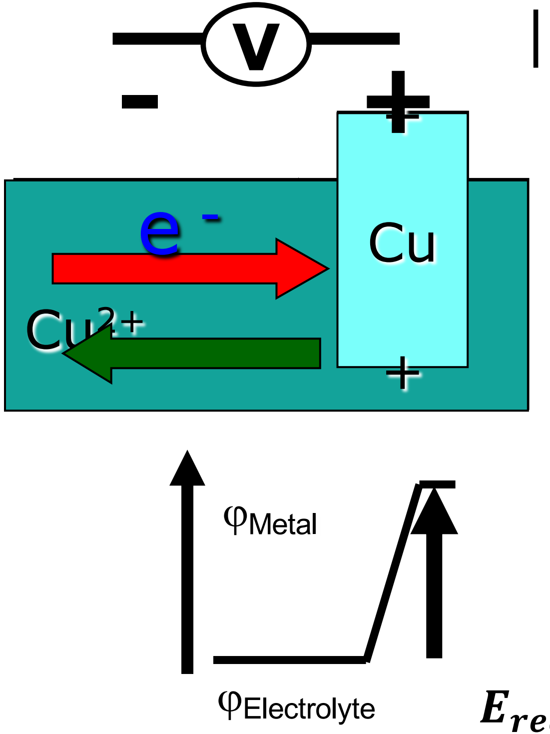
What **potential** appears at the interface electrode/electrolyte?

$E°_{Cu/Cu^{2+}} = +0.34\ \mathrm{V}$

$E°_{Zn/Zn^{2+}} = -0.74\ \mathrm{V}$

$$Cu^{2+} + 2\,e^- \underset{Ox}{\overset{Red}{\rightleftarrows}} Cu \qquad (V-7.2)$$

$$|i_-| = i_+ = i_0 \qquad (V-7.1)$$



$$E_{red/ox} = \varphi_{Me} - \varphi_{El}$$

#ElectrodePotential

What is the structure of the **electrode/electrolyte interface**?

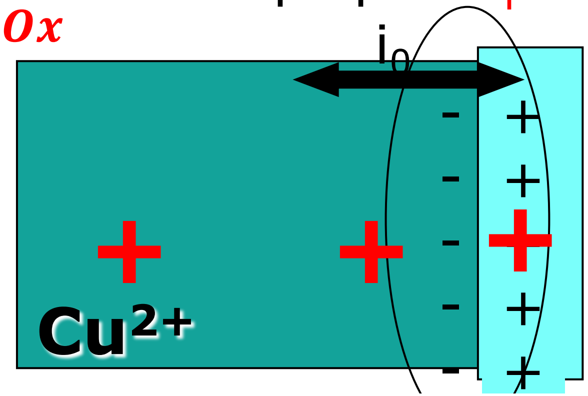
$E°_{Cu/Cu^{2+}} = +0.34\ \mathrm{V}$

$E°_{Zn/Zn^{2+}} = -0.74\ \mathrm{V}$

$$Cu^{2+} + 2\,e^- \rightleftarrows Cu$$

Red / Ox

$$|i_-| = i_+ = i_0$$



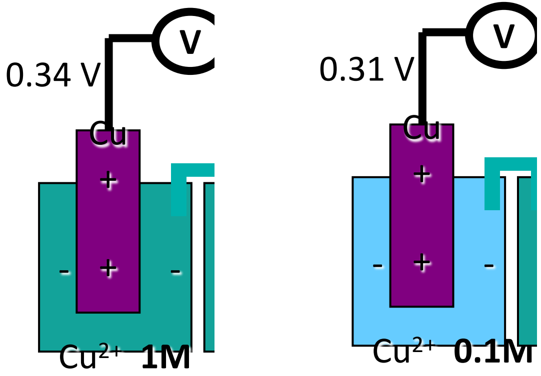
+0.34

#DoubleLayer

$$E_{redox} = \Delta\varphi(i=0) = \varphi_{Me} - \varphi_{El}$$

How does **concentration** affect electrode potential?

Walter Nernst
1864-1941



$$Red$$

$$\nu_{Ox} Ox + \nu_e e^- \rightleftarrows \nu_{Red} Red \qquad (V-6.1)$$

$$Ox$$

$$E_{Red/Ox} = E^\circ_{Red/Ox} + \frac{RT}{\nu_e F} \ln \frac{[Ox]^{\nu_{Ox}}}{[Red]^{\nu_{Red}}} \qquad (V-7.3)$$

#NernstEquation

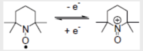
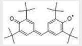
Comparing **Strengths** of Oxidants and Reductants

$MnO_4^-(aq) + 8\ H^+(aq) + 5\ e^- \rightarrow Mn^{2+}(aq) + 4\ H_2O(l)$	+1.51
$Cl_2(g) + 2\ e^- \rightarrow 2\ Cl^-(aq)$	+1.36
$O_2(g) + 4\ H^+(aq) + 4\ e^- \rightarrow 2\ H_2O(l)$	+1.23
$Ag^+(aq) + e- \rightarrow Ag(s)$	+0.80
$O_2(g) + 2\ H_2O(l) + 4\ e^- \rightarrow 4\ OH^-(aq)$	+0.40;
$Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)$	+0.34
$AgCl(s) + e^- \rightarrow Ag(s) + Cl^-(aq)$	+0.22
$2\ H^+(aq) + 2\ e^- \rightarrow H_2(g)$	0.00
$Fe^{3+}(aq) + 3\ e^- \rightarrow Fe(s)$	-0.04
$Zn^{2+}(aq) + 2\ e^- \rightarrow Zn(s)$	-0.76
$2\ H_2O(l) + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$	-0.83
$Mg^{2+}(aq) + 2\ e^- \rightarrow Mg(s)$	-2.36
$Li^+(aq) + e^- \rightarrow Li(s)$	-3.05

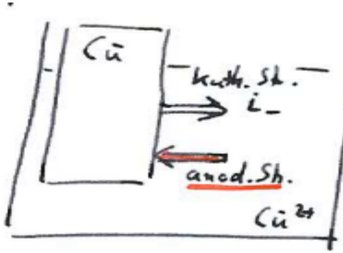
$(V - Tab.7.1)$

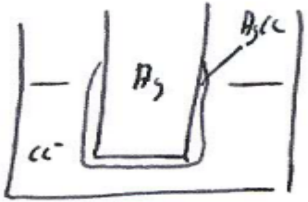
#ElectrochemicalSeries

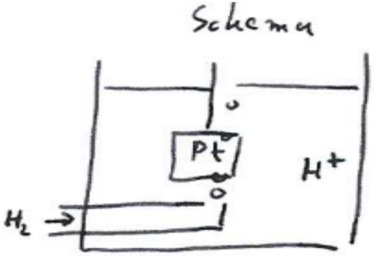
#StandardPotential

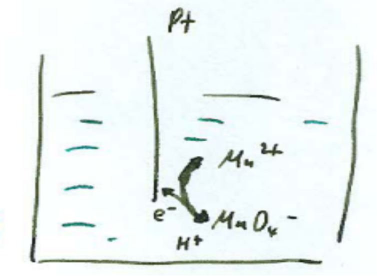
Oxid. Form / red. Form	Durchtrittsreaktion	E_{redox}^0 (V)
Oxidierte Form ist stark oxidierend		
MnO_4^-, H^+/Mn^{2+}, H_2O	$MnO_4^-(aq) + 8\,H^+(aq) + 5\,e^- \rightarrow Mn^{2+}(aq) + 4\,H_2O(l)$	+1,51
Cl_2/Cl^-	$Cl_2(g) + 2\,e^- \rightarrow 2\,Cl^-(aq)$	+1,36
O_2, H^+/H_2O	$O_2(g) + 4\,H^+(aq) + 4\,e^- \rightarrow 2\,H_2O(l)$	+1,23; +0.82 bei pH = 7
TEMPO-Radikal		≈ 0,95
Ag^+/Ag	$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0,80
O_2, H_2O/OH^-	$O_2(g) + 2\,H_2O(l) + 4\,e^- \rightarrow 4\,OH^-(aq)$	+0,40; +0.82 bei pH = 7
Galvinoxyl-Radikal		≈ 0,28
Cu^{2+}/Cu	$Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$	+0,34
$AgCl/Ag$, Cl^-	$AgCl(s) + e^- \rightarrow Ag(s) + Cl^-(aq)$	+0,22
Fe^{3+}/Fe	$Fe^{3+}(aq) + 3\,e^- \rightarrow Fe(s)$	–0,04
Pb^{2+}/Pb	$Pb^{2+}(aq) + 2\,e^- \rightarrow Pb(s)$	–0,13
Sn^{2+}/Sn	$Sn^{2+}(aq) + 2\,e^- \rightarrow Sn(s)$	–0,14
Fe^{2+}/Fe	$Fe^{2+}(aq) + 2\,e^- \rightarrow Fe(s)$	–0,44
Zn^{2+}/Zn	$Zn^{2+}(aq) + 2\,e^- \rightarrow Zn(s)$	–0,76
H_2O/H_2, OH^-	$2\,H_2O(l) + 2\,e^- \rightarrow H_2(g) + 2\,OH^-(aq)$	–0,83; –0.42 bei pH = 7
Al^{3+}/Al	$Al^{3+}(aq) + 3\,e^- \rightarrow Al(s)$	–1,66
Mg^{2+}/Mg	$Mg^{2+}(aq) + 2\,e^- \rightarrow Mg(s)$	–2,36
Li^+/Li	$Li^+(aq) + e^- \rightarrow Li(s)$	–3,05
Reduzierte Form ist stark reduzierend		

What do the **reduced and oxidized species** of the redox couple look like?









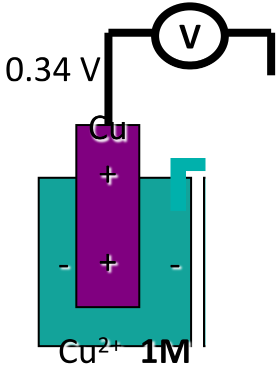
#TypesOfElectrodes

$E^\circ_{Cu/Cu^{2+}} = +0.34\ \text{V}$

$E^\circ_{Zn/Zn^{2+}} = -0.74\ \text{V}$

Making Electrodes (1): **Metal / Ion**

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



$$Cu^{2+}\ (aq) + 2\ e^- \underset{Ox}{\overset{Red}{\rightleftarrows}} Cu\ (s)$$

$$E_{Cu\ /Cu^{2+}} = E^\circ_{Cu\ /Cu^{2+}} + \frac{RT}{F}\ln \frac{\quad}{\quad}$$

#ElectrodeOfThe1st Kind

$(vgl.V-7.8)$

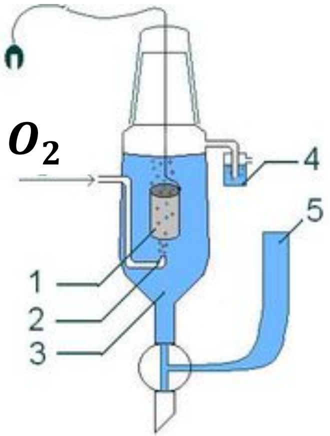
$\frac{RT}{F} ln(10) \approx 59\ mV$

$E°_{H_2/H^+} = +0.00\text{ V}$

$E°_{H_2O/O_2} = +1.23\text{ V}$

Making Electrodes (2): **Gas / Inert Metal / Ion**

$$Pt/O_2(g)/H^+(aq)$$



Red

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

Ox

$(V - 11.52)$

$$E_{O_2/H^+} = E°_{O_2/H^+} + \frac{RT}{F}\ln\frac{\quad}{\quad}$$

1.23 V

$(V - 11.53)$

#GasElectrode

$(vgl.V - 7.6)$

$$\frac{RT}{F}ln(10) \approx 59\,mV$$

$E°_{H_2/H^+} = +0.00\text{ V}$

$E°_{H_2O/O_2} = +1.23\text{ V}$

True or False?

Consider an oxygen electrode

A: The potential is pH independent

B: The potential increases with decreasing pH

C: The potential decreases with decreasing pH

D: The potential becomes smaller as the oxygen partial pressure decreases

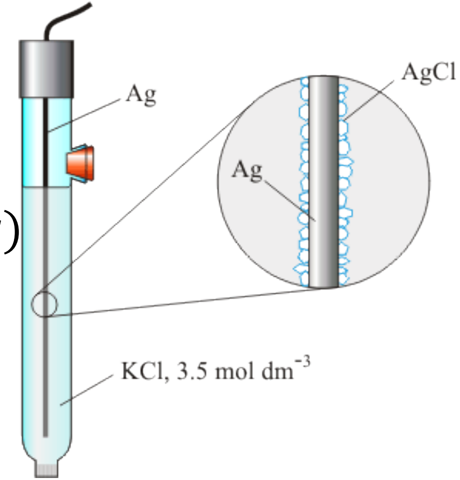
E: The potential increases with decreasing oxygen partial pressure

E: The potential is independent of oxygen pressure

$$E_{O_2/H^+} = E°_{O_2/H^+} + \frac{RT}{4\,F}\ln\frac{[H^+]^4[O_2]}{[H_2O]^2}_{(V - 11.53)}$$

$E^{\circ}_{Ag/AgCl} = +0.22$ V

Making Electrodes (3): **Metal / Insoluble Salt / Ion**



$$Ag\,(s)/AgCl(s)/Cl^{-}(aq)$$

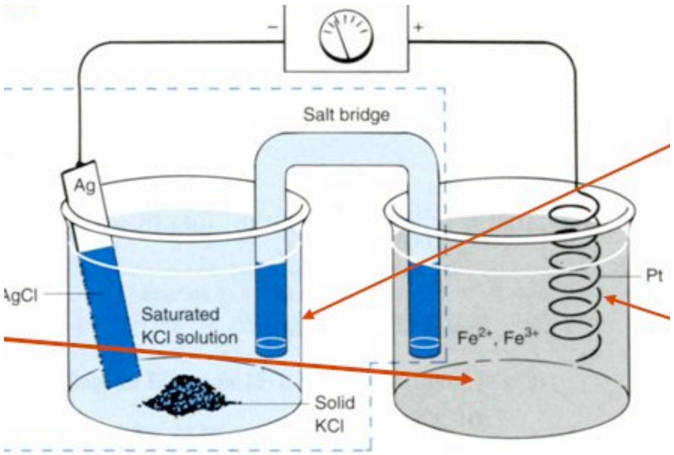
$$AgCl\,(s) + e^{-} \underset{Ox}{\overset{Red}{\rightleftarrows}} Ag\,(s) + Cl^{-}\,(aq) \qquad (V-7.10)$$

$$E_{Ag,Cl^{-}/AgCl} = E^{\circ}_{Ag,Cl^{-}/AgCl} + \frac{RT}{1\,F}\ln\frac{[AgCl]}{[Ag][Cl^{-}]} \qquad (V-7.11)$$

#ElectrodeOfThe2ndKind

$E^{\circ}_{Fe^{2+}/Fe^{3+}} = +0.77\ V$

Making Electrodes (3): **Ion$_{oxidized\ form}$, Ion$_{reduced\ form}$ / Inert Metal**



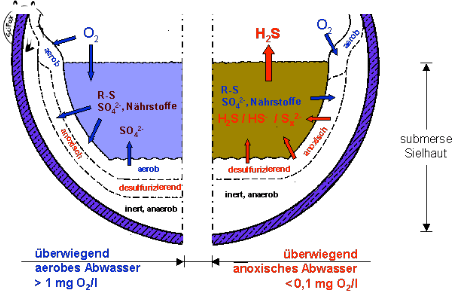
$$Pt/Fe^{2+}, Fe^{3+}(aq)$$

$$Fe^{3+}\ (aq) + 1\ e^{-} \underset{Ox}{\overset{Red}{\rightleftarrows}} Fe^{2+}\ (aq)$$

$$E_{Fe^{2+}/Fe^{3+}} = E^{\circ}_{Fe^{2+}/Fe^{3+}} + \frac{RT}{1\ F} ln \frac{[Fe^{3+}]}{[Fe^{2+}]}$$

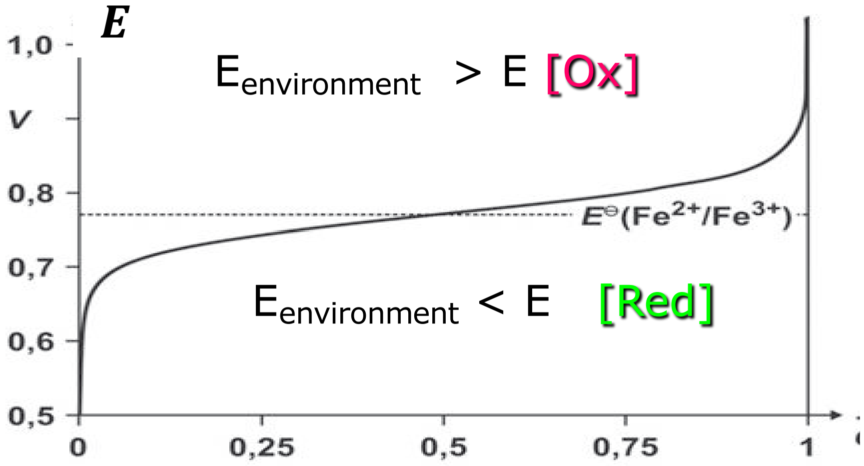
+0.77 V

#RedoxElectrode



Redox couples: which species dominate?

$$E_{Fe^{2+}/Fe^{3+}} = E°_{Fe^{2+}/Fe^{3+}} + \frac{RT}{1\,F} \ln \frac{[Fe^{3+}]}{[Fe^{2+}]}$$



$$\frac{[Fe^{3+}]}{[Fe^{2+}] + [Fe^{3+}]}$$

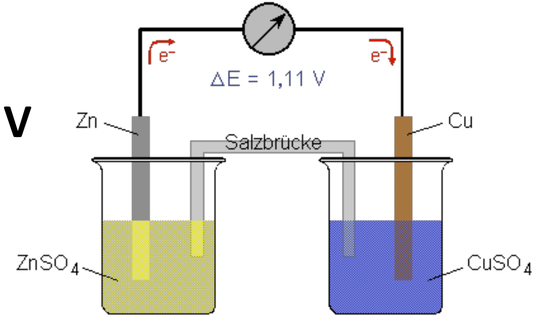
#RedoxEnvironment

$E°_{Cu/Cu^{2+}} = +0.34\ \text{V}$

$E°_{Zn/Zn^{2+}} = -0.74\ \text{V}$

Combination of electrodes: What **voltage** is produced?

John Daniell
1790 - 1845



-0.76 V **+0.34 V**

$Zn^{2+} + 2\ e^- \leftarrow Zn$ **Ox** $(V-8.4)$

Red $Cu^{2+} + 2\ e^- \rightarrow Cu$ $(V-8.5)$

$$Zn + Cu^{2+} \rightarrow Cu + Zn^{2+} \quad (V-8.9)$$

$$\Delta E_{Galv} = E_{Cath} - E_{An}$$

$(V-8.3)$

T	*E*
20°C	1,0990 V
25°C	1,0986 V
30°c	1,0982 V

$(V-Tab. 8.1)$

#OpenCircuitVoltage

#GalvanicCell

#ElectroMotiveForce (EMF)

$E°_{Cu/Cu^{2+}} = +0.34\ \text{V}$

$E°_{Zn/Zn^{2+}} = -0.74\ \text{V}$

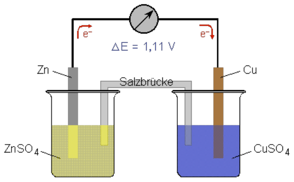
Redox reactions: **spontaneous or reversible**?

e^-

$$Cu^{2+}(aq) + Zn\,(s) \rightarrow Zn^{2+}(aq) + Cu\,(s) \quad (V-8.9)$$

$$q_{p,spon.} = \Delta_R H \qquad \Delta_R H = -217\,\frac{\text{kJ}}{\text{mol}} \quad (V-8.14)$$

$$w_{p,spon.} = 0$$



$$Zn^{2+} + 2\,e^- \leftarrow Zn \qquad Cu^{2+} + 2\,e^- \rightarrow Cu$$

Ox — Red

$$q_{p,\,rev.} = T \cdot \Delta_R S \qquad T\,\Delta_R S = -6\,\frac{\text{kJ}}{\text{mol}}$$

$$w_{el.,rev.} = \Delta_R G \qquad \Delta_R G = -212\,\frac{\text{kJ}}{\text{mol}} \quad (V-8.10)$$

#Heat **#Work**

#Reversibility

$E^\circ_{Br^-/Br_2} = +1.1\ \text{V}$

$E^\circ_{Cl^-/Cl_2} = +1.4\ \text{V}$

A chlorine electrode (E = 1.4V) and a bromine electrode (E = 1.1V) are combinde.

True or False?

A: The chlorine electrode is cathode and negative terminal

B: The chlorine electrode is cathode and positive termial

C: The chlorine electrode is anode and a negative termial

D: The chlorine electrode is anode and positive terminal

E: The reduction takes place at the chlorine electrode

E: The oxidation takes place at the chlorine electrode

$$Br_2\,(l) + 2\,e^- \underset{Ox}{\overset{Red}{\rightleftarrows}} 2\,Br^-\,(aq)$$

$$Cl_2\,(g) + 2\,e^- \underset{Ox}{\overset{Red}{\rightleftarrows}} 2\,Cl^-\,(aq)$$

$E°_{Ag/Ag^+} = +0.80\ \text{V}$

$E°_{H_2O/O_2} = +1.23\ \text{V}$

Discussing a **galvanic cell**

5. Eine Sauerstoffelektrode (pH = 7, p(O_2) = 1 bar) und eine Silberelektrode (c(Ag^+) = 10^{-4} mol/L) werden bei 298–K zusammengeschaltet.

- Berechnen Sie die Leerlaufspannung *E* dieser galvanischen Zelle
- Welche Elektrode ist die Anode; welche Elektrode ist der +-Pol der Zelle?
- Welche Redoxreaktion läuft insgesamt ab, wenn Strom durch die Zelle fließt?

$(\boldsymbol{V - 11.49 - 11.59})$

PCÜ46 Zusammenschaltung von Elektroden - Leerlaufspannung einer galvanischen Zelle

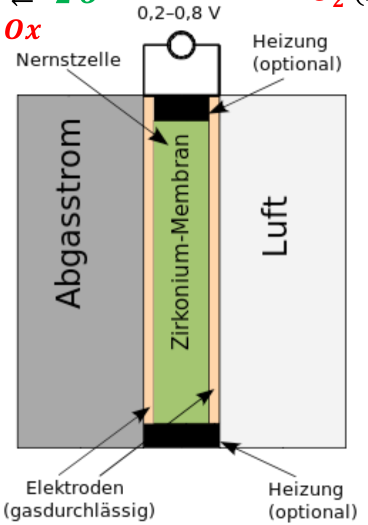
Physikalische Chemie by SciFox

Ag^+/Ag	$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0,80
O_2, H_2O/OH^-	$O_2(g) + 2\ H_2O(l) + 4\ e^- \rightarrow 4\ OH^-(aq)$	+0,40; +0.82 bei pH = 7

Two oxygen electrodes combined = **lambda sensor**

$$O_2\,(<1Pa) + 4\,e^- \underset{Ox}{\overset{Red}{\rightleftarrows}} 2\,O^{2-}$$

$$O_2\,(21\,kPa) + 4\,e^- \underset{Ox}{\overset{Red}{\rightleftarrows}} 2\,O^{2-}$$



$$E(O_2) = E°(O_2) + \frac{RT}{4\,F}\ln\frac{[O_2]^{<1Pa}}{[O^{2-}]^2} \qquad (V-8.16)$$

$$E(O_2) = E°(O_2) + \frac{RT}{4\,F}\ln\frac{[O_2]^{21\,kPa}}{[O^{2-}]^2} \qquad (V-8.15)$$

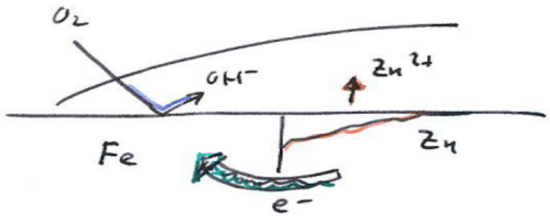
#ConcentrationCell

$$\Delta E = \frac{RT}{4\,F}\ln\frac{[O_2]^{21\,kPa}}{[O_2]^{<1Pa}} \qquad (V-8.17)$$

$E°_{Fe/Fe^{2+}} = -0.44$ V

$E°_{H_2O/O_2} = +1.23$ V

Redox reaction & electrochemical series: does iron **corrode** in air at pH = 7?



$$O_2 + 4H^+ + 4\,e^- \rightarrow 2\,H_2O \quad (V-8.20)$$

$E(OH^-/O_2) = +0.4\,V$ @ $pH\ 7$

$$O_2 + 4H^+ + 2Fe \rightarrow 2\,Fe^{2+} + 2H_2O \quad (V-8.18)$$

$E(Fe\,/Fe^{2+}) = -0.44\,V$

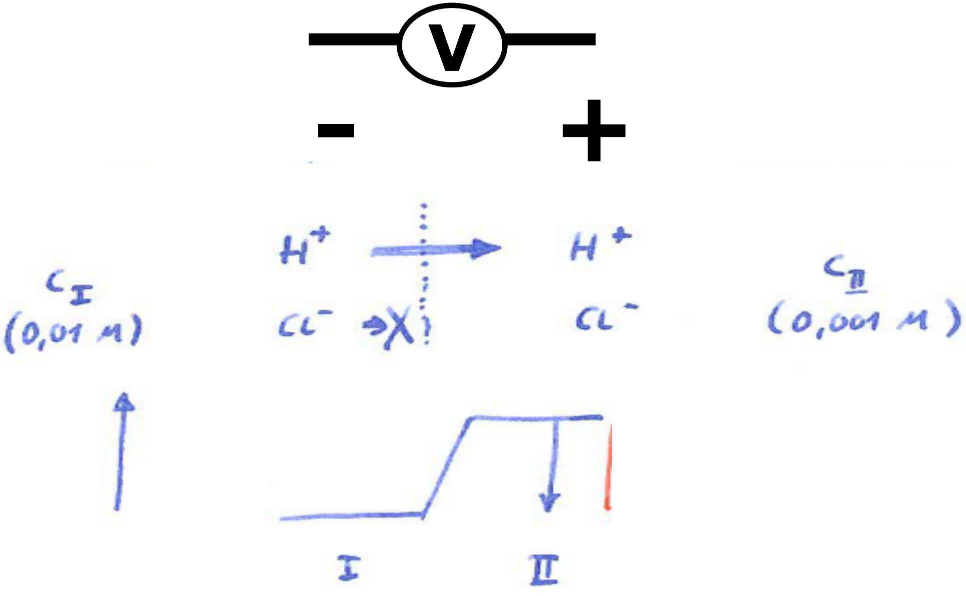
$$2\,Fe^{2+} + 4\,e^- \leftarrow 2\,Fe \quad (V-8.19)$$

$$E°(Red) > E°(Ox)$$

#AffinityRedoxReaction

#Corrosion

Contacting **two electrolytes**



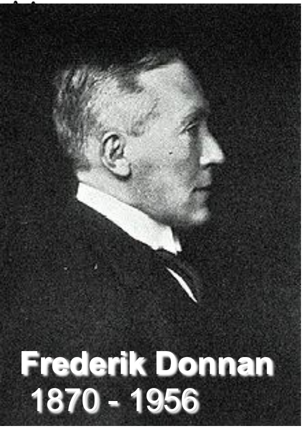
$$\Delta_{Mem}\varphi = \varphi(II) - \varphi(I) = -\frac{RT}{z_i\,F}\ln\frac{[i]^{II}}{[i]^{I}}$$

(V − 9.2)

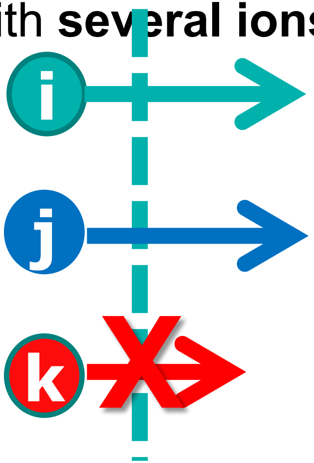
$$\frac{RT}{F}ln(10) \approx 59\ mV$$

#LiquidJunctionPotential

#MembranePotential



Electric potential between electrolytes with **several ions involved**



$$\Delta_{Donnan}\varphi = -\frac{RT}{z_i\,F}\ln\frac{[i]^{II}}{[i]^{I}} = -\frac{RT}{z_j\,F}\ln\frac{[j]^{II}}{[j]^{I}}$$

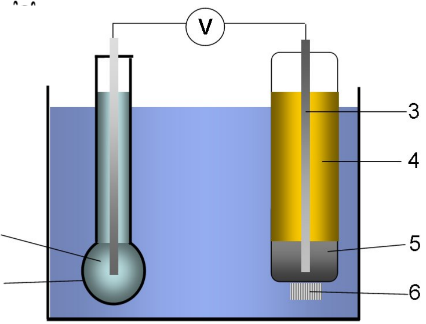
$(V-9.3)$

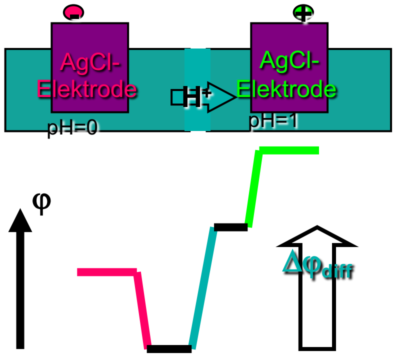
$$z_i = z_j = 1$$

$$[i]^{I}[j]^{I} = [i]^{II}[j]^{II}$$

$(V-9.5)$

#DonnanPotential





$$\Delta_{Mem}\varphi = \varphi(II) - \varphi(I) = -\frac{RT}{z_i\,F}\ln\frac{[i]^{II}}{[i]^{I}}$$

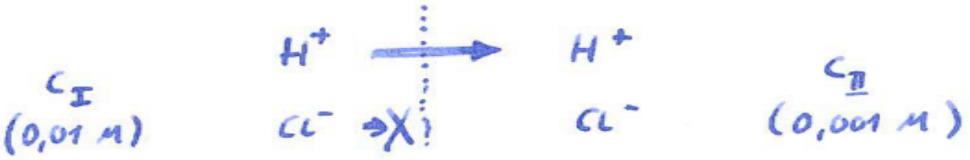
(V – 9.2)

Two solutions with different pH are separated by a semi-permeable membrane.

A: The solution with the higher pH has the more positive potential

B: The solution with the lower pH has the more positive potential

C: Both solutions have the same potential

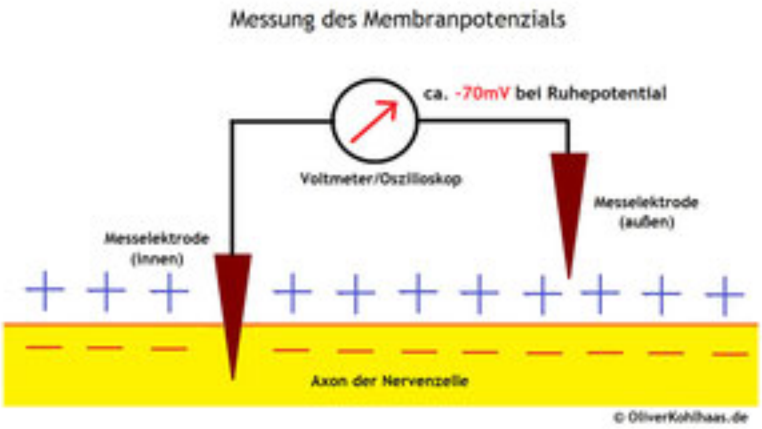


Discussing electric potential on a **biological membrane**

Fragen

8. Eine semipermeable Zellmembran (durchlässig für Kaliumionen, undurchlässig für alle anderen Ionen, d. h. $t_+ = 1$) trennt zwei Lösungen der Kaliumkonzentrationen 155 mmol/L (Zellinneres) und 4 mmol/L (Zelläußeres)

- Berechnen Sie das Diffusionspotenzial $\Delta\varphi_{diff}$ zwischen den Lösungen bei 37 °C.
- Welche Lösung hat das positivere Potenzial?



PCÜ48 Membranpotential an einer Nervenzelle - Kaliumionen im Donnan-Gleichgewicht

Physikalische Chemie by SciFox

GALVANICCELL ⊕ Why does the **cathode** of a galvanic cell have a different polarity than the cathode of an electrolysis cell?

Elektrolysis ⊖

#Cathode

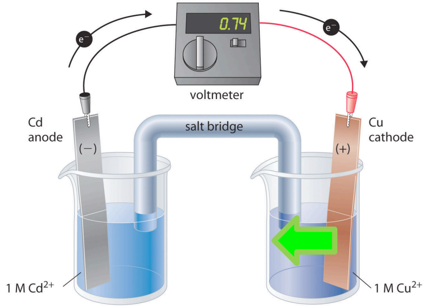
#Reduction

GALVANICCELL ⊖

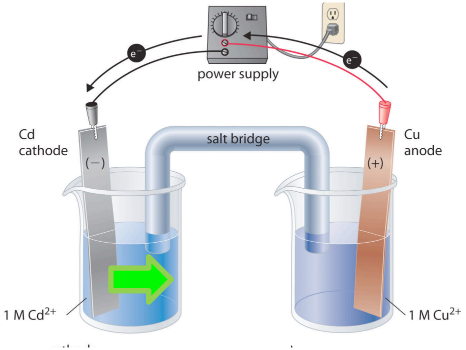
Electrolysis ⊕

#Anode

#Oxidation



$(V - 6.7)$

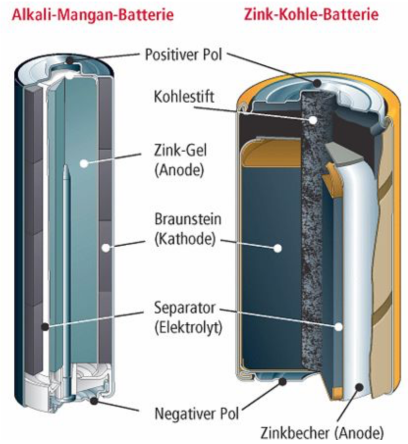


#UnsolicitedElectronFlow

#EnforcedElectronFlow

$E^\circ_{Zn/Zn^{2+}} = -0.74\ \mathrm{V}$

How much **energy** is stored n a zinc-carbon battery?



$$3.6\ g\ Zn$$

$$Zn^{2+} + 2\ e^- \leftarrow Zn$$

$$n = \frac{m}{M} = \frac{I \cdot t}{\nu_e F}$$

$$3.6\ g\ Zn \cong 56\ mmol$$
$$\cong 0.1\ mol\ e \cong 10000\ As$$
$$= 3000\ mAh$$
$$1.5\ V * 10000 As = 15\ kJ$$

REFERENCES:

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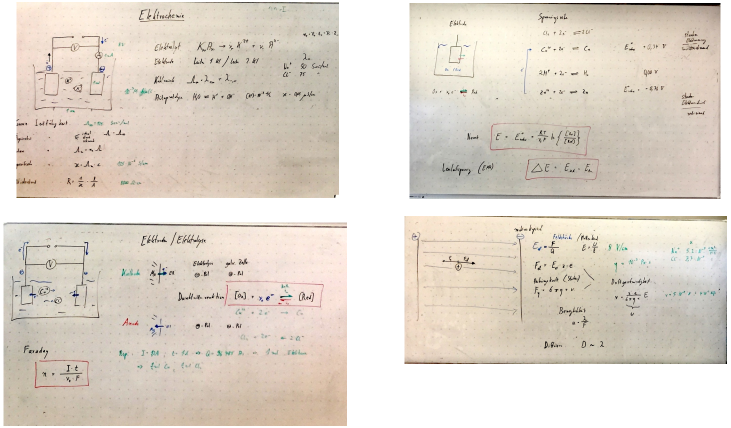
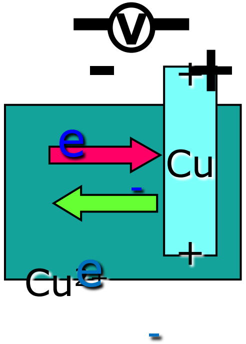
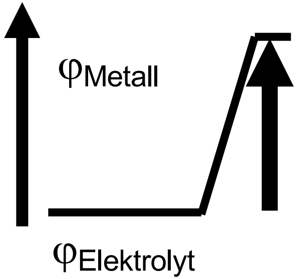


PHYSICAL CHEMISTRY BASICS
PART 12: ELECTRODES



$$E_{red/ox} = \varphi_{Me} - \varphi_{El}$$



$$\nu_{Ox}\,[Ox] + \nu_e\, e^- \underset{i_{an}}{\overset{i_{cath}}{\rightleftarrows}} \nu_{Red}[Red]$$

$$n = \frac{m}{M} = \frac{I \cdot t}{\nu_e F}$$

$$E_{red/ox} = E°_{red/ox} + \frac{R\,T}{\nu_e\,F} \ln \frac{[Ox\,]^{\nu_{Ox}}}{[Red\,]^{\nu_{Red}}}$$

$$\Delta E_{Galv} = E_{Cath} - E_{An}$$