1412 - E3

CH 17 - Electrochemistry

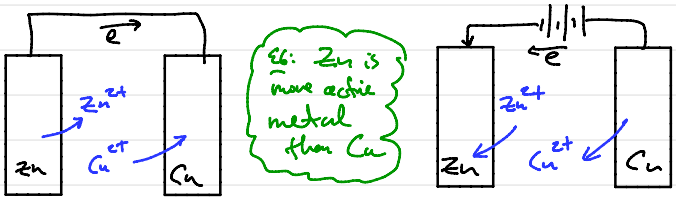
Notes

17

Pole Designations

voltaic

electrolytic



	OIL	RIG	RIG	OIL
RED CAT >	anode	cathode	cathode	anode
	← Cathode is always site of Reduction			
POLE >	⊖	⊕	⊖	⊕

CATHODE/ANODE DESIGNATION… changes wrt direction of electron flow

POLE DESIGNATION… doesn't change; same for Voltaic and Electrolytic

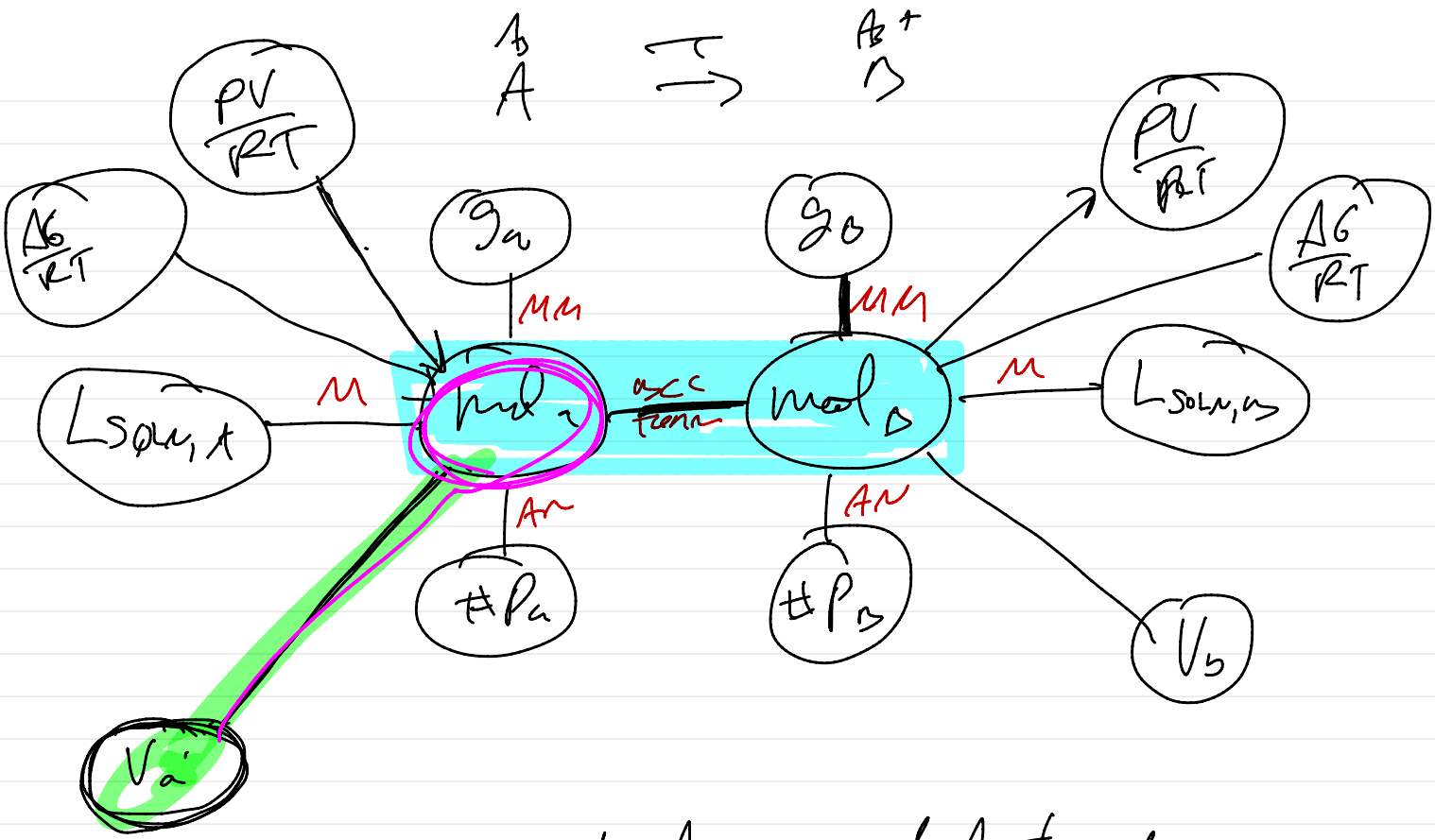
← More active metal = ⊖

… has "more negative" $E^\circ_{reduction}$

… more active metal wants to lose electrons (corrode)

$\Delta G = -nFE^\circ$

… more (−) → ← more positive…



$$1\,Ag \rightarrow 1Ag^{+} + 1e^{-}$$

CHAPTER 17: ELECTROCHEMISTRY

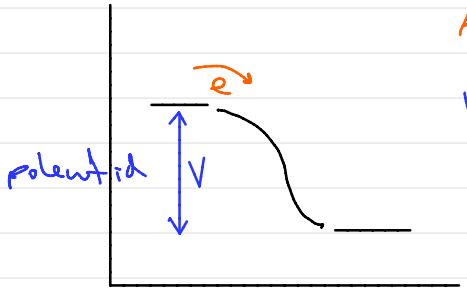
Introduction [6.1]

Intro

Electricity — flow of electrons

Q: ¿where do electrons come from?
A: Atoms
Q: ¿what happens in a REDOX Rxn?
A: Electrons are transferred

Q: ¿Why do electrons flow?
A: To minimize / dispose energy.



A = # electrons over hill per unit of time

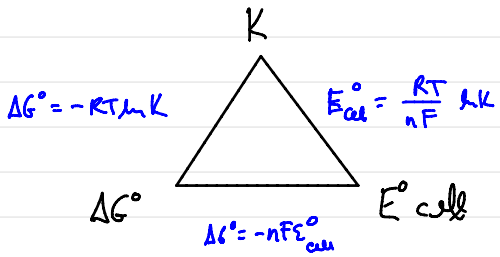
V = potential = hill height

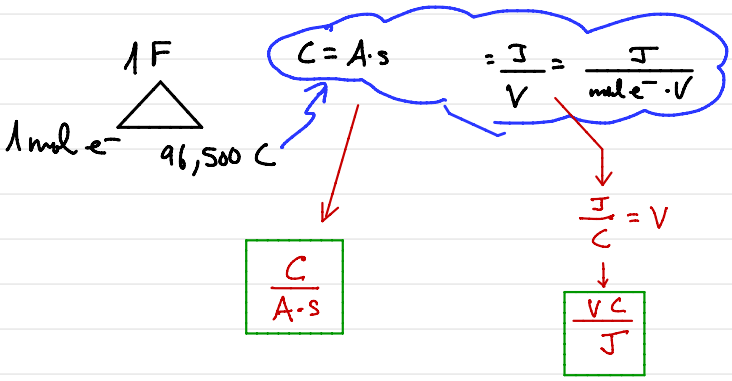
Units & Equations

Common Electrical Terms

Quantity	Definition	Measure or Unit
Electric charge	Charge on a proton	1.602×10^{-19} C
Electric current	The movement of charge	ampere = A = 1 C/s
Electric potential	The force trying to move the charge	volt = V = J/C
Electric field	The force acting upon other charges in the vicinity	

Table 17.1



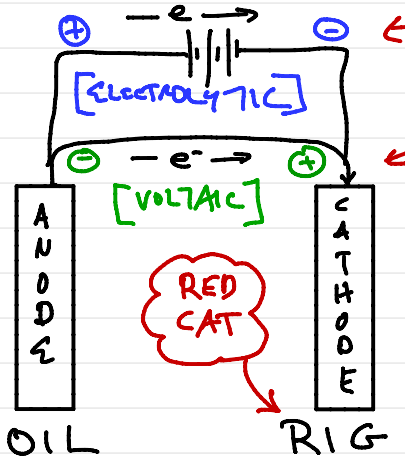


NOTE: Section 17.1 – the portion which deals w/ balancing RedOx Rxns will be covered in the lab

Wed, Nov 6

Galvanic Cells [17.2]

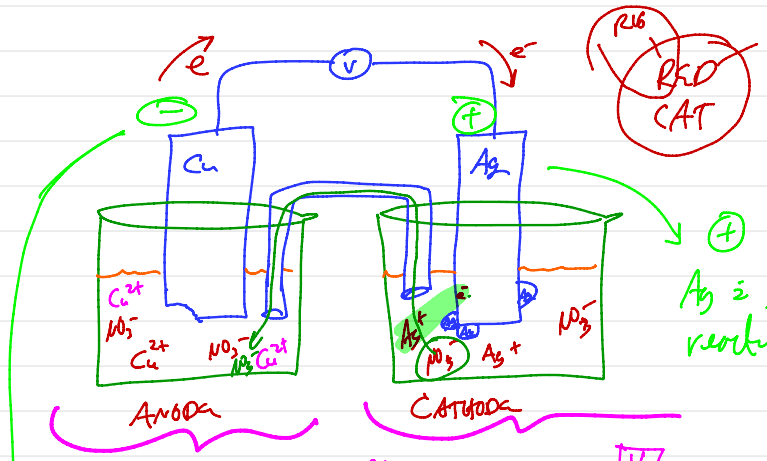
Cells : 2-Types ① Galvanic or Voltaic (Battery) - natural flow of e^-
② Electrolytic - forced flow of e^-



Roll ⊕ or ⊖ is determ by metal activity ... does not change

Zn more active than Cu
Zn = ⊖ roll ; Cu = ⊕ roll

learn how to differentiate
① ANODE vs. CATHODE ← red cat
② ⊕ vs. ⊖ ← ⊖ = more active



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

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

$Cu + 2Ag^+ \rightarrow 2Ag + Cu^{2+}$

Cu is more reactive

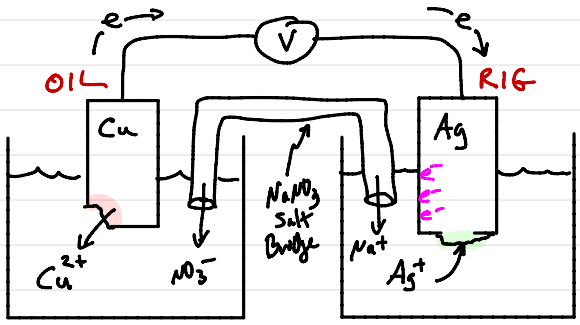
V

Reactive	Unreactive
K	Ag
Na	Au

OIL

RIG

$Cu | Cu^{2+}(aq, 1M) || Ag^+(aq, 1M) | Ag$



copper(II) nitrate soln
Anode
⊖ more reactive cell

silver nitrate soln
Cathode (RED CAT)
⊕ less reactive cell

OIL (ANODE)
RIG (CATHODE)

Half-cell RXNS [OpenStax Chem]

$$Cu \longrightarrow Cu^{2+} + 2e^-$$

$$2e^- + 2Ag^+ \longrightarrow 2Ag$$

$$2Ag^+ + Cu \longrightarrow Cu^{2+} + 2Ag$$

CELL NOTATION ⇓ Shorthand

$$Cu_{(s)} \mid Cu^{2+}_{(aq,\,1M)} \parallel Ag^+_{(aq,\,1M)} \mid Ag_{(s)}$$

anode (OIL) ‖ cathode (RIG)

⟹ ALL "COEFFICIENTS" IGNORED ⟸

(EX) Cell Notation

¿Use cell notation to describe a cell in which copper(II) ions are reduced to copper, and Zn metal is oxidized to zinc ions?

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

$Zn \rightarrow Zn^{2+} + 2e$ OIL

OIL RIG

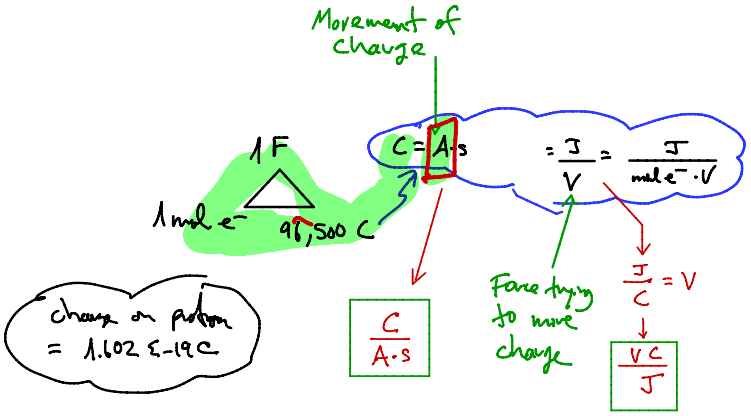
$Zn \,|\, Zn^{2+} \,\|\, Cu^{2+} \,|\, Cu$

anode cathode

OIL RED

$Zn \,|\, Zn^{2+} \,\|\, Cu^{2+} \,|\, Cu$

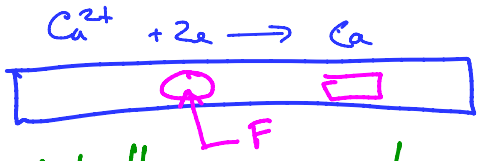
Calculations using Amp and Time and Grams



0.0324F

(EX) AW from Faraday

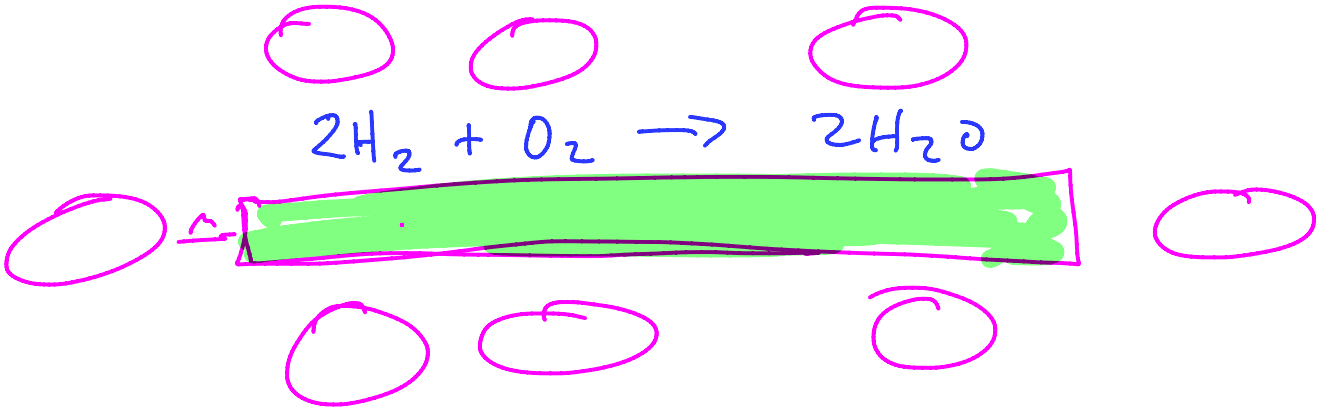
¿Calculate the atomic weight of calcium, knowing 0.0324 F was needed to produce 0.651 g,Ca?



$$\frac{\square\ g\ Ca}{\square\ mol\ Ca} = \frac{0.651\ g}{1} \left| \frac{1}{0.0324\ F} \right| \frac{1\ F}{1\ mol\ e^-} \left| \frac{2\ mol\ e^-}{1\ mol\ Ca} \right. = \frac{40.2\ g}{mol}$$

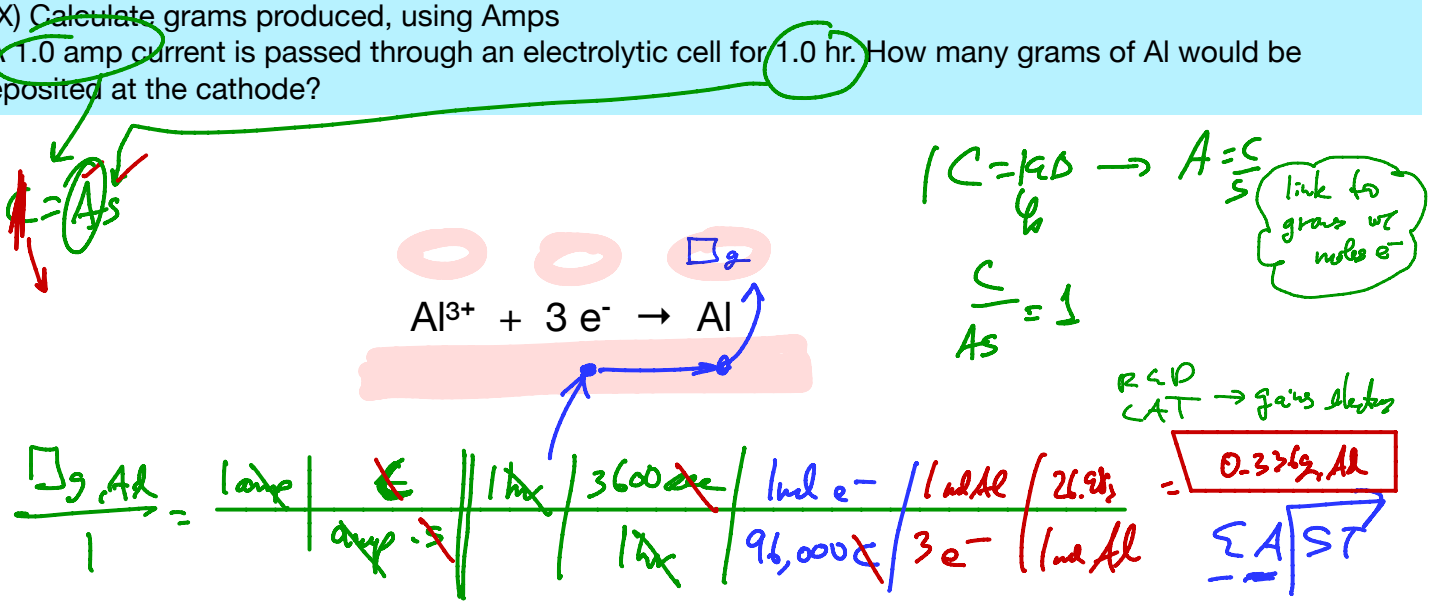
(CF) (gmi)

$$\frac{\square\ g\ Ca}{mol} = \frac{0.651\ g}{1} \left| \frac{1}{0.0324\ F} \right| \frac{1\ F}{1\ mol\ e^-} \left| \frac{2\ mol\ e^-}{1\ mol\ Ca} \right. = \frac{40.2\ g}{mol}$$


$2H_2 + O_2 \rightarrow 2H_2O$

(EX) Calculate grams produced, using Amps

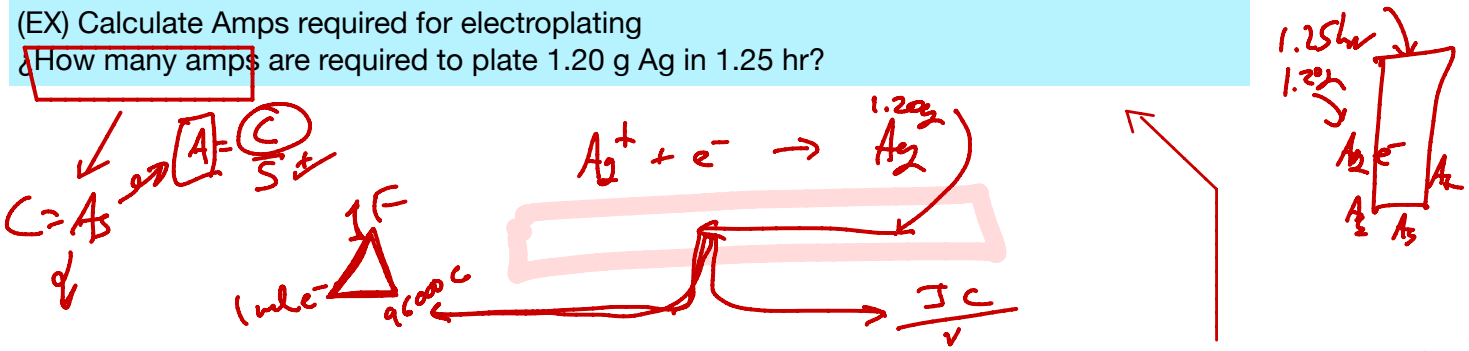
¿A 1.0 amp current is passed through an electrolytic cell for 1.0 hr. How many grams of Al would be deposited at the cathode?



$$\frac{\square\, g_{Al}}{1} = \frac{1\ amp}{1} \mid \frac{1\ C}{1\ amp\ sec} \parallel \frac{1\ hr}{1} \mid \frac{3600\ sec}{1\ hr} \parallel \frac{1\ mol\ e^-}{96000\ C} \mid \frac{1\ mol_{Al}}{3\ mol\ e-} \mid \frac{26.98\ g}{1\ mol_{Al}} = \frac{0.336\ g}{1}$$

(EX) Calculate Amps required for electroplating

¿How many amps are required to plate 1.20 g Ag in 1.25 hr?



$$\frac{\square A}{1} = \frac{\square C}{s} = \frac{1.20\ g_{Al}}{1.25\ hr} \mid \frac{1\ hr}{3600\ sec} \parallel \frac{1\ mol_{Ag}}{107.87\ g_{Al}} \mid \frac{1\ mol\ e-}{1\ mol\ Ag} \mid \frac{96000\ C}{1\ mol\ e-} = \frac{0.239\ C}{s}$$

Mass calculations usually require eventual use of a "1 mole e^-" conversion factor

electrochem data $\Leftrightarrow$ 1 mole$^-$ $\Leftrightarrow$ mass

Standard Reduction Potentials [17.3]

- SHE (Std Hydrogen Electrode) = ground zero for "potential"
- Analogous to °C = zero at F.P. of water.

$2H^+(aq, 1M) + 2e^- \longrightarrow H_2(g, 1\,atm)$ $\quad E^\circ = 0\ V$.

Building the STD Reduction Scale

- For any cell voltage measured in which SHE is one of the ½-cells, then the voltage of the cell is also the voltage of the other ½-cell, given $E^\circ(SHE) = 0$.

SHE

$Pt(s) \mid H_2(g, 1\,atm) \mid H^+(aq, 1M) \parallel Cu^{2+}(aq, 1M) \mid Cu(s)$

$H_2(g) \longrightarrow 2H^+(aq) + 2e^-$ $\quad E^\circ = 0$ (OIL, anode)

$Cu^{2+} + 2e^- \longrightarrow Cu(s)$ $\quad E^\circ = ?$ (RIG, cathode)

$Cu^{2+} + H_2(g) \longrightarrow 2H^+(aq) + Cu(s)$ $\quad +0.337$

∴ The standard potential, $E^\circ(Cu^{2+}|Cu) = +0.337$

Measured directly

$$E^\circ_{cell} = E^\circ_{CATHODE} - E^\circ_{ANODE}$$

OIL RIG

OIL || RIG

− RIG

E°_{cell}

Selected Standard Reduction Potentials at 25 °C

Half-Reaction	$E°$ (V)
$Au^{3+}(aq) + 3e^- \longrightarrow Au(s)$	+1.498
$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$	+1.35827
$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$	+1.229
$Pt^{2+}(aq) + 2e^- \longrightarrow Pt(s)$	+1.20
$Br_2(aq) + 2e^- \longrightarrow 2Br^-(aq)$	+1.0873
$Ag^+(aq) + e^- \longrightarrow Ag(s)$	+0.7996
$Hg_2^{2+}(aq) + 2e^- \longrightarrow 2Hg(l)$	+0.7973
$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$	+0.771
$MnO_4^-(aq) + 2H_2O(l) + 3e^- \longrightarrow MnO_2(s) + 4OH^-(aq)$	+0.558
$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$	+0.5355
$NiO_2(s) + 2H_2O(l) + 2e^- \longrightarrow Ni(OH)_2(s) + 2OH^-(aq)$	+0.49
$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	+0.337
$Hg_2Cl_2(s) + 2e^- \longrightarrow 2Hg(l) + 2Cl^-(aq)$	+0.26808
$AgCl(s) + 2e^- \longrightarrow Ag(s) + Cl^-(aq)$	+0.22233
$Sn^{4+}(aq) + 2e^- \longrightarrow Sn^{2+}(aq)$	+0.151
$2H^+(aq) + 2e^- \longrightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$	−0.126
$Sn^{2+}(aq) + 2e^- \longrightarrow Sn(s)$	−0.1262
$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$	−0.257
$Co^{2+}(aq) + 2e^- \longrightarrow Co(s)$	−0.28
$PbSO_4(s) + 2e^- \longrightarrow Pb(s) + SO_4^{2-}(aq)$	−0.3505
$Cd^{2+}(aq) + 2e^- \longrightarrow Cd(s)$	−0.4030
$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	−0.447
$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$	−0.744
$Mn^{2+}(aq) + 2e^- \longrightarrow Mn(s)$	−1.185
$Zn(OH)_2(s) + 2e^- \longrightarrow Zn(s) + 2OH^-(aq)$	−1.245

Half-Reaction	$E°$ (V)
$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$	−0.7618
$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$	−1.662
$Mg^{2}(aq) + 2e^- \longrightarrow Mg(s)$	−2.372
$Na^+(aq) + e^- \longrightarrow Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \longrightarrow Ca(s)$	−2.868
$Ba^{2+}(aq) + 2e^- \longrightarrow Ba(s)$	−2.912
$K^+(aq) + e^- \longrightarrow K(s)$	−2.931
$Li^+(aq) + e^- \longrightarrow Li(s)$	−3.04

Table 17.2

LESS ACTIVE END

(+)

more (+) $E°_{red}$

Important rule of thumb

- Target: Largest (+) $E°_{cell}$

∴ swap (chg to OXID) an equation lower on the table.

- Note: Appendix "L" provide extensive list, in alphabetical order

(−)

MORE REACTIVE END

more (−) $E°_{red}$

more / N or M

(EX) Calc E°cell [ex 17.4b]
¿A galvanic cell is composed of an Ag electrode in a 1M AgNO3 solution, and a Mg electrode in 1M Mg(NO3)2 — all at 25°C. What is E°cell?

① Make the metal
② Assign RED potentials to each
③ Reverse the more (−) half-cell
④ Add the half-cell values - such that e's cancel

ΔE°_{RED}

$2Ag^+ + 2e^- \rightarrow 2Ag$ +0.7996 V

~~$Mg^{2+} + 2e^- \rightarrow Mg$ −2.372 V~~

$Mg \rightarrow Mg^{2+} + 2e^-$ +2.372 V

~~$2e^-$~~ $+ 2Ag^+ + Mg \rightarrow 2Ag + Mg^{2+}$ ~~$+ 2e^-$~~ +3.172 V

OIL

RED

$Mg \mid Mg^{2+} \parallel Ag^+ \mid Ag$

(EX) Calc E°cell [ex 17.4b]

¿A galvanic cell is composed of an Ag electrode in a 1M $AgNO_3$ solution, and a Mg electrode in 1M $Mg(NO_3)_2$ — all at 25°C. What is E°cell?

1. Make the metal
2. Assign RED potentials to each
3. Reverse the more (−) half-cell
4. Add the half-cell values

	$E°_{RED}$ (V)
$Ag^+ + 1e^- \rightarrow Ag$	+0.7996
$Mg^{2+} + 2e^- \rightarrow Mg$	−2.372

$$2\times\ Ag^+ + e^- \rightarrow Ag \quad +0.7996$$
$$Mg \rightarrow Mg^{2+} + 2e \quad +2.372$$
$$Mg + 2Ag^+ \rightarrow Mg^{2+} + 2Ag \quad +3.172\ V$$

or

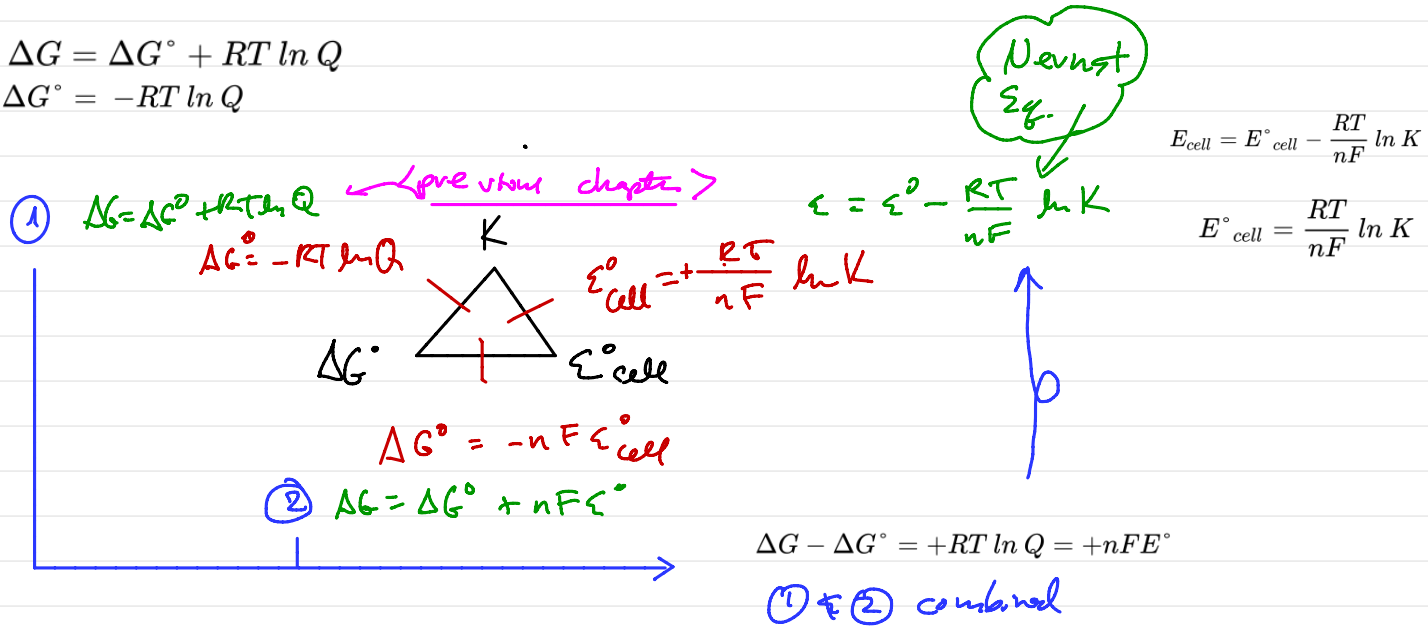
$$Mg\,|\,Mg^{2+}\,\|\,Ag^+\,|\,Ag$$

coefficients NOT written in shorthand

The Nernst Equation [17.4]

$$\Delta G = \Delta G^\circ + RT\,ln\,Q$$

$$\Delta G^\circ = -RT\,ln\,Q$$



$$\Delta G^\circ = -nFE^\circ_{cell}$$

$$\Delta G = \Delta G^\circ + nFE^\circ_{cell}$$

$$E = E^\circ - \frac{RT}{nF}\,ln\,K$$

"RT/nF" Simplified

$$E^\circ_{cell} = \frac{RT}{nF} \ln K = \frac{8.314\ J}{mol\ K} \left| \frac{298.15\ K}{} \right| \left| \frac{}{F} \right| \left| \frac{F\ mole^-}{96{,}500\ C} \right| \left| \frac{C\ V}{J} \right| \ln K$$

unit

$$E^\circ_{cell} = \frac{0.0257\ V}{n} \ln K = \frac{0.0592\ V}{n} \log K$$

log vs. ln

variable

n = # electrons in BCE

$$K = 10^{nE^\circ_{cell}/0.0592\ V}$$

$$E_{cell} = E^\circ_{cell} - \frac{0.0592\ V}{n} \log Q$$

TIP:

$\frac{RT}{F} = 0.025V\ V$

@ 25°C

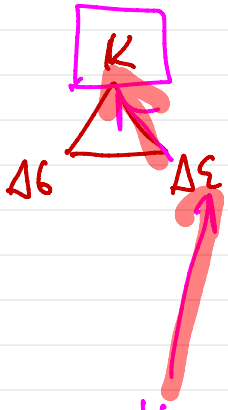
$$E^\circ cell = E^\circ_{CATHODE} - E^\circ_{ANODE}$$

OIL RIG

(EX) Calculate K, from tabulated E° values, for the reaction of iodide ion and bromine?

$K = \frac{PROD}{REACT}$ E° & K Absent

I^+ Br_2



$$2I^- + Br_2 \rightarrow I_2 + 2Br^- \quad +0.545$$

$$2I^- \rightarrow I_2 + 2e^- \quad -0.535$$

$$2e^- + Br_2 \rightarrow 2Br^- \quad +1.08$$

Std. Table $+0.545$ V

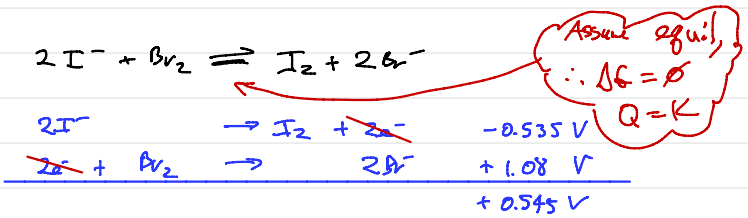
$-RT\ln K = -nF\mathcal{E}$

$$\ln K = \frac{nF\mathcal{E}}{RT} = \frac{2}{1}\left|\frac{96500\ J}{V\ mol}\right|\left|\frac{0.545\ V}{1}\right|\left|\frac{mol\ K}{8.314\ J}\right|\frac{1}{298\ K} = \frac{42.45}{1}$$

$1.08 - 0.535$

$K = e^{42.45} = 2.7 \times 10^{18}$

(EX) Calculate K, from tabulated E° values, for the reaction of iodide ion and bromine?



$\Delta G^\circ = -RT\ln K = -nFE$ ← Nernst

$\frac{8.314\ J}{mol \cdot K}$ $298\ K$ $\frac{96{,}500\ J}{V\ e^-}$ $n = 2$ $+0.545\ V$

$$\ln K = \frac{nFE}{RT} = \frac{2}{1} \times \frac{96500J}{V\ mol} \times \frac{0.545V}{1} \times \frac{mol\ K}{8.134J} \times \frac{1}{298K} = \frac{42.45}{1}$$

$$e^{\ln K} = e^{42.45}$$

$$K = e^{42.45} = 2.7E18$$

Shortcut: recall $\frac{RT}{F}$ @ 25°C $= 0.0257\ V$

$$\ln K = \frac{n\mathcal{E}}{0.0257\ V} = \frac{2 \mid 0.545\ V}{0.0257\ V} = 42.42$$

(EX) Nernst: Calc E(cell)
¿Calculate E(cell) for Fe3+/Fe2+ electrode when Fe2+ is exactly 5-times that of Fe3+?

ratio

$Fe^{3+} + 1e^- \rightarrow Fe^{2+}$ $\quad E^\circ_{cell}$ 0.771 V ← from Red table

$E = E^\circ - \frac{0.0257}{n} \ln Q$ ← $\frac{[prod]}{[reac]}$

$E = 0.771 - \frac{0.0257}{1} \ln \frac{[Fe^{2+}]^5}{[Fe^{3+}]^1}$

$E = 0.771 - 0.0257 \ln 5$

$E = 0.730$ V

★ E°_{cal} → pdt → E_{cell}

$5[Fe^{3+}] = 1[Fe^{2+}]$

$\frac{Fe^{2+}}{Fe^{3+}} = \frac{1}{5}$

(EX) Nernst: Calc E(cell)
¿Calculate E(cell) for Fe3+/Fe2+ electrode when Fe2+ is exactly 5-times that of Fe3+?

① Used equation, to det'n [PROD] & [REACT] for Nernst:

$Fe^{3+} + e^- \rightarrow Fe^{2+} \quad E^\circ = 0.771\ V$

② Next, flesh-out Nernst Eq:

$$E = E^\circ - \frac{0.0257}{n} \ln Q$$

$$E = 0.771 - \frac{0.0257}{1} \ln \frac{[Fe^{2+}]}{[Fe^{3+}]}$$

Translation of text: is "=" 5 times "5"

$Fe^{2+} = 5[Fe^{3+}] \rightarrow \frac{Fe^{2+}}{Fe^{3+}}$

$$E = 0.771 - 0.0257 \ln 5$$

$$E = 0.771 - 0.041 = \boxed{0.730\ V}$$

Batteries and Fuel Cells [17.5]

Types

Primary — non re-chargealbe

- Dry Cell

$$2MnO_2(s) + 2NH_4Cl(aq) + Zn(s) \rightarrow Zn(aq) + Mn_2O_3(s) + 2NH_3(aq) + H_2O(l) + 2Cl^-$$

- Alkaline

$$Zn(s) + 2MnO_2(s) \rightarrow ZnO(s) + Mn_2O_3(s) \qquad E^\circ_{Cell} = +1.43\,V$$

Secondary — re-chargealbe

- Ni-Cad

$$Cd(s) + NiO_2(s) + 2H_2O(l) = Cd(OH)_2(s) + Ni(OH)_2(s)$$

- Lithium Ion

$$LiCoO_2 + x\,C_6 \rightleftharpoons Li_{x-1}CoO_2 + x\,LiC_6$$

- Lead Acid

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

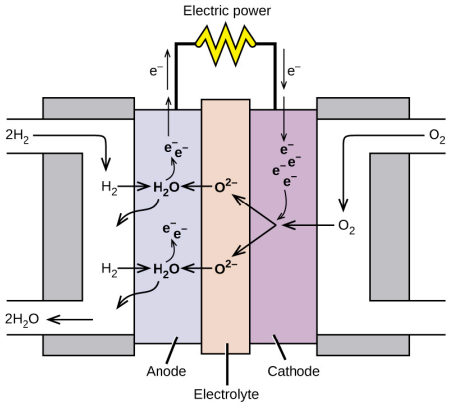
Fuel Cell — requires continuous source of fuel

A: $2H_2 + 2O^{2-} \rightarrow 2H_2O + 4e^-$

C: $4e^- + O_2 \rightarrow 2O^{2-}$

cell: $2H_2^{0} + O_2^{0} \rightarrow 2H_2O^{+1\,-2}$ $E^\circ = 0.9V$

OIL
ANODE



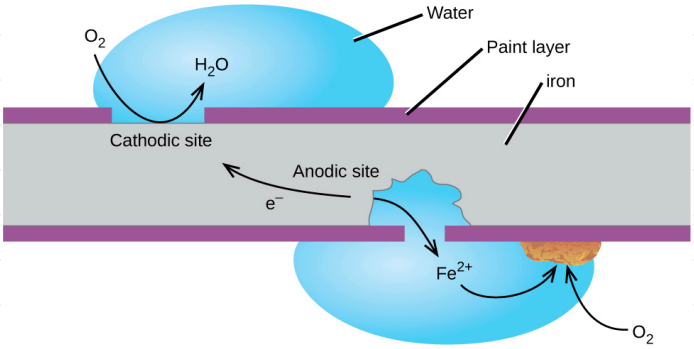
RIG
CATHODE

Corrosion [17.6]

def: degradation of metals due to electrochemical processes.

Corrosion of Iron

$$\begin{array}{lll} & & E^\circ \\ A: & Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^- & +0.44\ V \\ C: & 4e^- + O_2(g) + 2H^+(aq) \longrightarrow 2Fe^{2+}(aq) + 2H_2O(l) & +1.23 \\ \hline & 2Fe(s) + O_2(g) + 4H^+(aq) \longrightarrow 2Fe_2O_3 \cdot xH_2O(s) + 8H^+(aq) & +1.67\ V \end{array}$$

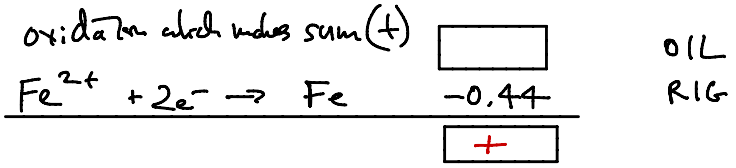


Methods of Prevention

① Cathode Protection — connect to sacrificial anode made of a more active metal
② Galvanization (zinc plated)
  — Zn > Fe, therefore, Zn will rust
  — see on some ships
③ Painting
④ Neutralizer
⑤ Filmers

Sacrificial Anode

- oxidation does the damage
- want more important metal to NOT corrode, so, pair it with a more active (and hopefully cheap, and structurally unimportant) metal



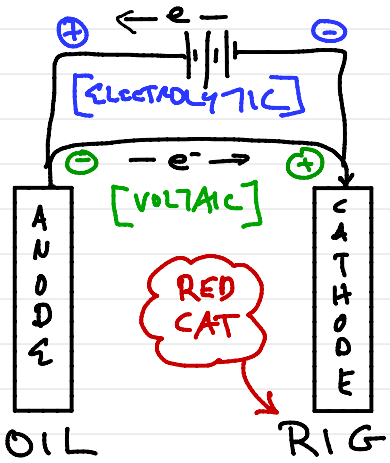
OIL
RIG

Typical : Zn -0.763 V reducing potential
Mg -2.37 V

the will rust, instead of iron

Electrolysis and Electrolytic Cells [17.7]

- Opposite of Galvanic: here, electrical energy causes chem reaction (and not the other way around).



PROMPT: If Na is so reactive, where is it found?

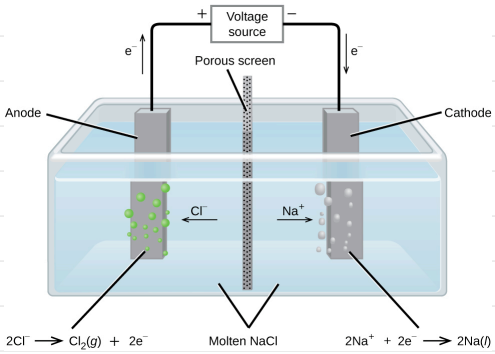
ANSWER:

Electrolysis of Molten Sodium Chloride

		E°
A.	$2Cl^-(l) \rightarrow Cl_2 + 2e^-$	$-1.3\ V$
C	$e^- + Na^+(l) \rightarrow Na(l)$	$-2.7\ V$
	$2Na^+(l) + 2Cl^-(l) \rightarrow 2Na(l) + Cl_2(g)$	$-4.0\ V$

To get rxn to go, must supply minimum of 4 V

Down's Cell →



(EX) Reaction Spontaneity

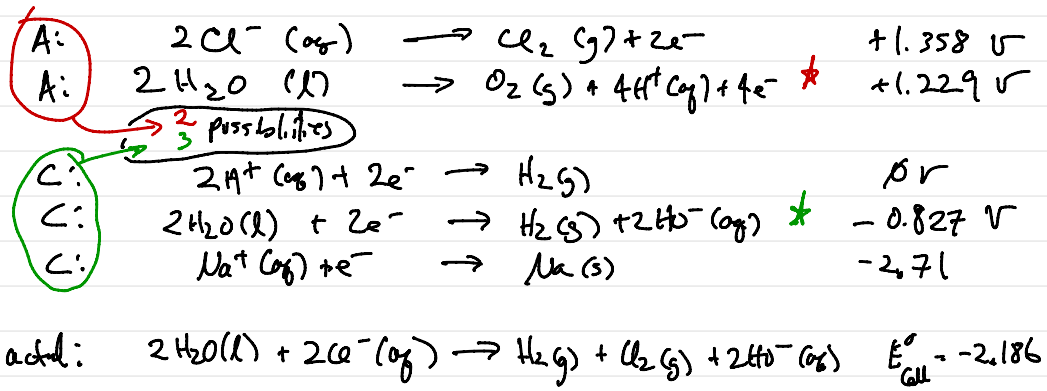
¿During the manufacture of Na metal, how many volts of electricity are produced?

ANSWER: NONE. The calculation shows a net negative (−) voltage, so 4 volts has to be applied to get the reaction to go.

(As you know, Na does not spontaneously form.)

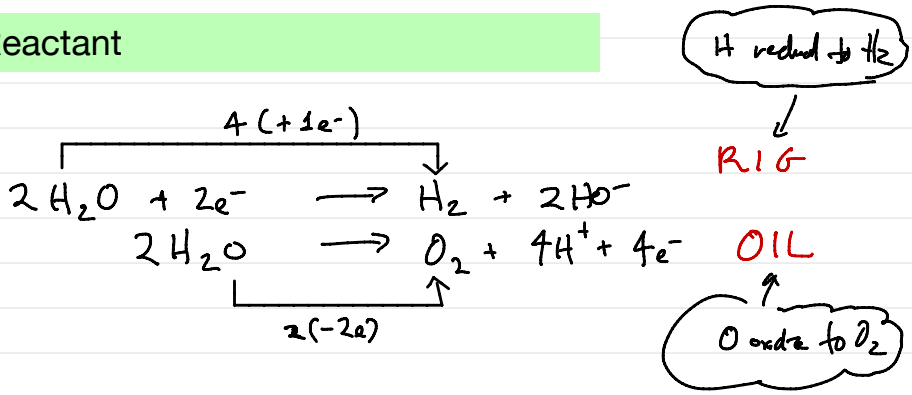
SKIP

Electrolysis of Aqueous NaCl

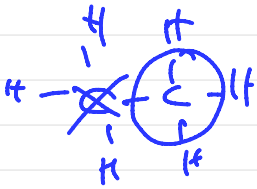


SKIP

Water as Reactant



Oxid–Red of Carbon (examples of Oxidation Numbers)



-4	-3	-2	-1	0	+2	+4
CH_4	C_2H_6	C_2H_4	C_2H_2	CH_2O	CO	O=C=O

General Observations

- OXID # of C increases with increasing # of OXYGENS (more oxygens, higher ON)
- OXID # of C increases with increasing bond multiplocity (single < double < triple)

TIP: metals are reduced; and their ions are oxidized

(EX) Calculate # e's
¿Calculate the number of electrons that have a total charge of 1 coulomb?

$$\frac{\square\, e}{1} = \frac{1\, C}{1} \mid \frac{1\, mol\, e^-}{96000\, C} \mid \frac{6.02E23\, e^-}{1\, mol\, e^-} = \frac{6.24E18\, e^-}{1}$$

(EX) Nernst

¿Given standard reduction potentials, H+/H2(g) = 0 V and O2(g),H+/H2O(l) = 1.229 V, calculate E(cell) when pressure of H2 is 3.00 atm, and O2 is 1.25 atm?

non. STP

E° (V)

$2\left(H_2 \rightarrow 2H^+ + 2e^-\right)$ $\quad 0$ $\quad$ SHE

$4e^- + 4H^+ + O_2 \rightarrow H_2O \quad +1.229$

$2H_2 + O_2 \rightarrow 2H_2O \quad +1.229 = E^\circ$

$$E_{cell} = E^\circ_{cell} - \left[\frac{RT}{nF} \ln Q\right]$$

$$= 1.229 - \left[\frac{0.0257}{4} \ln Q\right]$$

$$= 1.229 - \left[\frac{0.0257}{4} \ln 0.0889\right]$$

$$= \boxed{1.245 \text{ Volts}}$$

$$Q = \frac{[PRO]}{[REACT]}$$

$$Q = \frac{1}{(P_{H_2})^2 (P_{O_2})^1}$$

$$= \frac{1}{(3.00)^2 (1.25)}$$

$$Q = 0.0889$$

(EX) Nernst
¿Given standard reduction potentials, H+/H2(g) = 0 V and O2(g),H+/H2O(l) = 1.229 V, calculate E(cell) when pressure of H2 is 3.00 atm, and O2 is 1.25 atm?

$$2\left(H_2 \longrightarrow 2H^+ + 2e^-\right) \qquad \not{0}$$

$$4e^- + 4H^+ + O_2 \longrightarrow 2H_2O \qquad 1.229$$

$$2H_2 + O_2 \longrightarrow 2H_2O \qquad +1.229 = E^\circ$$

arrangement picked to give (+) voltage

$$Q = \frac{1}{(P_{H_2})^2 (P_{O_2})} = \frac{1}{(3.00)^2 (1.25)} = 0.089$$

$$E = 1.229 - \frac{0.0257}{4} \ln 0.0889 = 1.2446 = \boxed{1.245\ V}$$

(EX) Nernst: calculate concentration

What is the [Ag+] in a half cell if the reduction potential of the Ag+/Ag couple is 0.40 V?

$Ag^+ + 1e^- \rightarrow Ag \rightarrow Q = \frac{1}{[Ag^+]}$

given: 0.40 V

from Table: +0.7994 V

$\mathcal{E} = \mathcal{E}^\circ - \frac{0.0257}{n} \ln \frac{1}{[Ag^+]}$ ← x

$n = 1$

$\ln \frac{1}{x} = \frac{0.7994 - 0.40}{0.0257} = 15.541$

$e^{15.541}$

$\frac{1}{x} = e^{15.541} = \frac{5.615 \text{ E6}}{1}$

$x = 1.8 \times 10^{-7}$

(EX) Nernst: calculate concentration
¿What is the [Ag+] in a half cell if the reduction potential of the Ag+/Ag couple is 0.40 V?