| = phase boundary.

ANODE — CATHODE

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

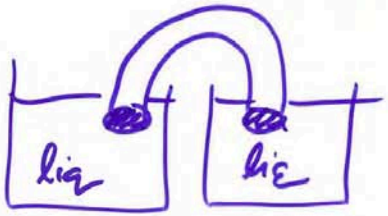
OX — RED

$Zn(s) \mid Zn^{2+}(aq)$: Zn-elec (s) / Zn^{2+}(aq)

$\parallel$: salt bridge — each end = phase boundary.

$Cu^{2+}(aq)$: liq

$Cu(s)$: solid



How can we predict a cell's voltage?

Standard Electrode Potentials

- Reduction potentials.

$$E^\circ_{cell} = E^\circ_{RHS} - E^\circ_{LHS} = E^\circ_{cathode} - E^\circ_{anode}$$

Table 19.1 lists std. elec. potentials.

good @ accepting e⁻s
good ox. agents

$3e^- + Au^{3+} \rightarrow$	Au	$E^\circ = +1.50\,V$	
$e^- + Ag^+ \rightarrow$	Ag	$E^\circ = +0.80\,V$	
$2e^- + Cu^{2+} \rightarrow$	Cu	$E^\circ = +0.34\,V$	
$2e^- + Pb^{2+} \rightarrow$	Pb	$E^\circ = -0.13\,V$	
$2e^- + Mg^{2+} \rightarrow$	Mg	$E^\circ = -2.37\,V$	
$e^- + Li^+ \rightarrow$	Li	$E^\circ = -3.05\,V$	

good @ donating e⁻s
good reducing agents.

$$E^\circ_{cell} = E^\circ_{RHS} - E^\circ_{LHS} \longrightarrow$$

More +ve E°, more favored rxn

More −ve E°, more favored reverse rxn

What's E°_{cell} for:

① $Mg(s) \,|\, Mg^{2+}(aq, 1M) \,||\, Ag^+(aq, 1M) \,|\, Ag(s)$

$E^\circ_{cell} = E^\circ_{RHS} - E^\circ_{LHS}$

$= +0.80V \ominus -2.37V$

$= +3.17V$

(+ve = spontaneous)

E°_{cell} = +ve $\leftrightarrow \Delta G^\circ$ = -ve

② $Cu(s) \,|\, Cu^{2+}(1M) \,||\, Au^{3+}(1M) \,|\, Au(s)$

$E^\circ_{cell} = E^\circ_{RHS} - E^\circ_{LHS}$

$= +1.50V \ominus +0.34V$

$= +1.16V$

Cell rxns? Chem eqns.

① ox: $Mg \rightarrow Mg^{2+} + 2e^-$

red: $(e^- + Ag^+ \rightarrow Ag) \times 2$

sum: $Mg(s) + 2Ag^+(aq) \xrightarrow{2e^-} Mg^{2+}(aq) + 2Ag(s)$

2nd cell:

ox: $(Cu \rightarrow Cu^{2+} + 2e^-) \times 3$

red: $(3e^- + Au^{3+} \rightarrow Au) \times 2$

$3Cu(s) + 2Au^{3+}(aq) \xrightarrow{6e^-} 3Cu^{2+}(aq) + 2Au(s)$

Thermodynamics

Physics: Electrical Energy (J) = Elec. Potential Difference (V) × Elec. charge (C)

Coulombs

Michael Faraday measured the |charge| on 1 mol e^-.

$1F = 96,500 C/mol$

↑ Faraday

can show: charge on n mole$^-$ × voltage.

$$\Delta G = -nF \times E_{cell}$$

$$\Rightarrow \boxed{\Delta G = -nFE_{cell}}$$

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

Effect of [ ] on E_{cell}

Nernst equation

redox rxn: $aA + bB \rightleftharpoons cC + dD$

$$Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ ← rxn quotient.

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

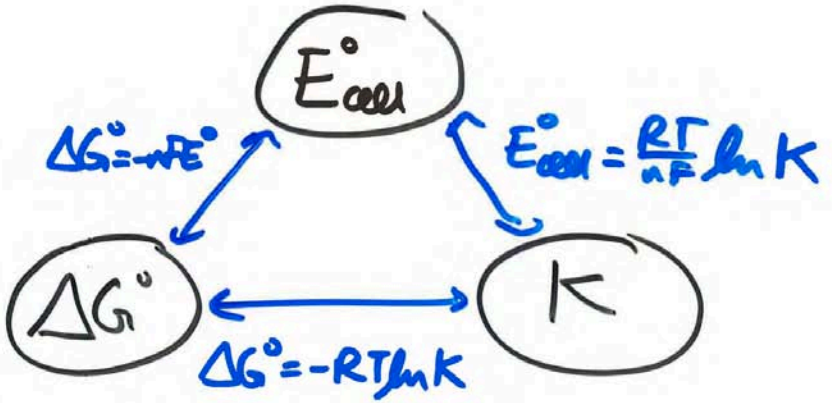
↑ actual conditions; ↑ STD conditions (1 atm, 1 M)

$$-nFE_{cell} = -nFE^\circ_{cell} + RT \ln Q \quad \div -nF$$

$$\boxed{E_{cell} = E^\circ_{cell} - \frac{RT}{nF} \ln Q}$$

Nernst eqn

cool



ΔG°	−ve	0	+ve
K	$\gg 1$	$= 1$	$\ll 1$
E°_{cell}	+ve	0	−ve

What's E°_{cell} and E_{cell} for:

$$Zn \,|\, Zn^{2+}(0.25M) \,\|\, Ag^+(0.0025M) \,|\, Ag$$

$$E^\circ_{cell} = E^\circ_{RHS} - E^\circ_{LHS}$$

$$= +0.80V - -0.76V$$

$$= +1.56V$$

rxn:

ox $\quad Zn \longrightarrow Zn^{2+} + 2e^-$

red $\quad (e^- + Ag^+ \longrightarrow Ag) \times 2$

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

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