Today

Review for our Quiz!

Thermo and Electrochemistry

What happens when the conditions are not standard

Nernst Equation

What is the oxidation number of N in KNO_3?

A. 0

B. +1

C. -1

D. +3

E. +5

K is +1, O is -2
molecule is no charge
1(+1) + 3(-2) = -5
N must be +5

Balance this half reaction

$$HNO_2 \longrightarrow NO_3^-$$

How many protons are in the balance 1/2 reaction?

(keeping the coeficient for NO_3^- as 1)

A. 0

B. 1 on the left

C. 1 on the right

D. 2 on the left

E. 3 on the right

$$HNO_2 \longrightarrow NO_3^-$$
$$HNO_2 + H_2O \longrightarrow NO_3^-$$
$$HNO_2 + H_2O \longrightarrow NO_3^- + 3H^+$$
$$HNO_2 + H_2O \longrightarrow NO_3^- + 3H^+ + 2e^-$$

Given following standard reduction potential,
which do you think would make the best reducing agent?

$$Cl_2(g) + 2e^- \longrightarrow 2Cl^- \qquad E° = 1.36\ V$$
$$I_2(s) + 2e^- \longrightarrow 2I^- \qquad E° = 0.54\ V$$
$$Br_2(l) + 2e^- \longrightarrow 2Br^- \qquad E° = 1.06\ V$$

A. Cl^-

B. Cl_2

C. I_2

D. I^-

E. Br_2

reducing agents are oxidized
hardest to reduce is easiest to oxidize
Lowest potential

Need to pick the reduced species
(it can be oxidized)

You reduce H^+ to H_2 in an electrochemical cell.
Your cell has a current of 1 Amp for 10 minutes
What is the total charge that is passed through the cell?

A. 1 C

B. 10 C

C. 600 C

D. 6000 C

$1\ A \times (10\ min) \times (60\ s\ min^{-1}) = 600\ C$

You reduce H^+ to H_2 in an electrochemical cell.
Your cell has a current of 1 Amp for 10 minutes
How many moles of electrons pass through the cell?

A. 600 C / F

B. 600 C x F

C. 1 A x F

F is $C\ mol^{-}$
Therefore the number of moles of electrons is q/F

You reduce H^+ to H_2 in an electrochemical cell.
The number of moles of electrons that pass through the cell is 6.2×10^{-3}. How many moles of H_2 are formed?

$$2H^+ + 2e^- \dashrightarrow H_2(g)$$

A. 6.2×10^{-3}

B. 3.1×10^{-3}

C. 1.2×10^{-2}

For every mole of H_2 you need two moles of electrons

You reduce H^+ to H_2 in an electrochemical cell.
Your cell has a current of 1 Amp for 10 minutes.
How many moles of H_2 are formed?

$$2H^+ + 2e^- \dashrightarrow H_2(g)$$

A. 6.2×10^{-3}

B. 3.1×10^{-3}

C. 1.2×10^{-2}

For every mole of H_2 you need two moles of electrons

In the following standard cell, what is E°_{cell} ?

$Zn(s) \mid Zn^{2+} \parallel H^+ \mid H_2\,(g)$

$Zn^{2+} + 2e^- \longrightarrow Zn(s) \quad E^\circ = -0.76\,V$
$2H^+ + 2e^- \longrightarrow H_2 \quad E^\circ = 0.0\,V$

A. 0.0 V

B. +0.76 V

C. -0.76 V

anode on the left
$Zn \mid Zn^{2+}$ anode $H^+ \mid H_2$ cathode

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 0 - (-0.76) = +0.76\,V$

Voltaic Cell

In the following standard Ecell, what is the sign of the cathode?

$Zn(s) \mid Zn^{2+} \parallel H^+ \mid H_2\,(g)$

$Zn^{2+} + 2e^- \longrightarrow Zn(s) \quad E^\circ = -0.76\,V$
$2H^+ + 2e^- \longrightarrow H_2 \quad E^\circ = 0.0\,V$

A. +

B. -

C. neither

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 0 - (-0.76) = +0.76\,V$

Voltaic Cell therefore cathode +

$E^\circ cell = 0$

For a battery which of the following is correct?

A. $E > 0$, $\Delta G > 0$, $K > 1$

B. $E > 0$, $\Delta G < 0$, $K > 1$

C. $E > 0$, $\Delta G < 0$, $K < 1$

D. $E < 0$, $\Delta G > 0$, $K > 1$

E. $E < 0$, $\Delta G < 0$, $K > 1$

F. $E < 0$, $\Delta G < 0$, $K < 1$

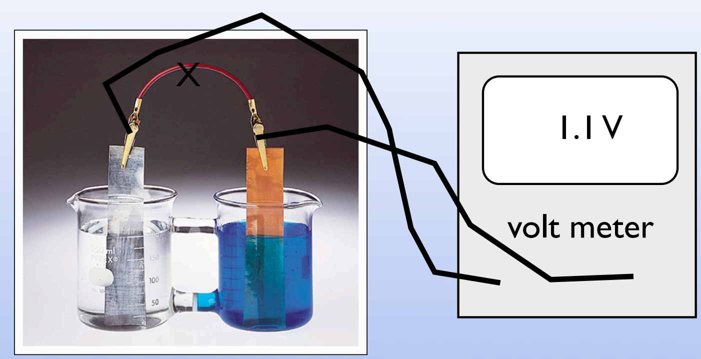
Battery = voltaic
Spontaneous

$E > 0$

$\Delta G < 0$
$K > 1$

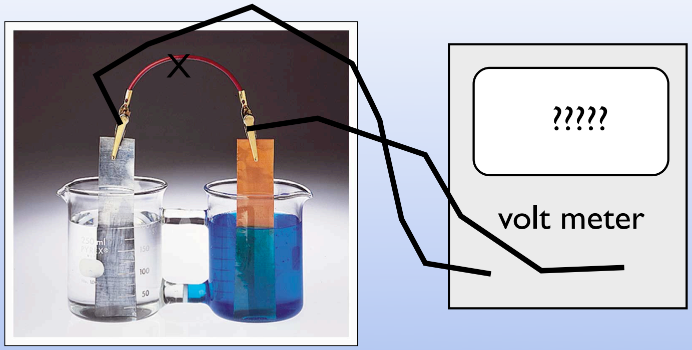
hint its not F.

We'll look at standard concentrations



1 M Zn^{2+} (aq) and 1 M Cu^{2+} (aq)
(note this is ridiculously concentrated)

What about other concentrations?



10^{-3} M Zn^{2+} (aq) and 10^{-1} M Cu^{2+} (aq) ???

Relationship between E and ΔG

ΔG is energy
E is electrical potential

Electric work (energy) is -charge x potential

work = $-q \times E$

$\Delta G = work_{max}$
$\Delta G = - q \times E_{max}$

From now on well now the Potential we calculate are the theoretical maximum
Real world never actually that good

Relationship between E and ΔG

$\Delta G = - q \times E$

What is the charge q?

$q = n \times F$

n is number of moles of electrons
F is the charge of one mole of electrons
F = 96,485 C (Faraday's Constant)

$$\Delta G = - nFE$$

Other concentrations and equilibrium

Let's remember equlibrium!

$\Delta G = \Delta G° + RT\ln Q$

at equilibrium $\Delta G = 0$
so $\Delta G° = -RT\ln K$

$-nFE = -nFE° + RT\ln Q$

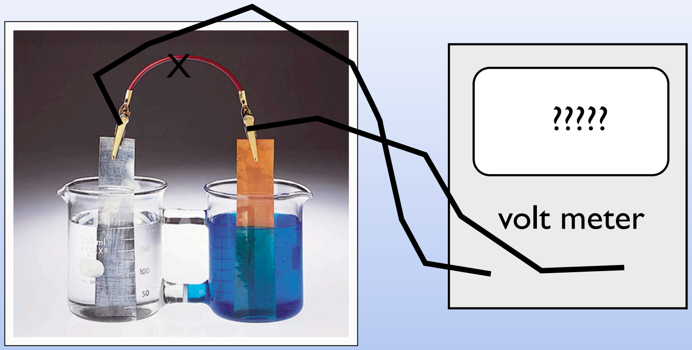
$$E = E° - \frac{RT}{nF}\ln Q$$

assume 25°C

$$E = E° - \frac{0.0591}{n}\log Q$$

log!

What about other concentrations?



10^{-3} M Zn^{2+} (aq) and 10^{-1} M Cu^{2+} (aq) ???

1 M Zn^{2+} (aq) and 1 M Cu^{2+} (aq) standard

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

$$Q = \frac{[Zn^{2+}]}{[Cu^{2+}]} = \frac{1}{1} = 1$$

$$E = E^\circ - \frac{0.0591}{n} \log Q$$

$$E = 1.10\text{V} - \frac{0.0591}{2} \log(1) = 1.10\text{V}$$

10^{-3} M Zn^{2+} (aq) and 10^{-1} M Cu^{2+} (aq) ???

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

$$Q = \frac{[Zn^{2+}]}{[Cu^{2+}]} = \frac{(10^{-3})}{(10^{-1})} = 10^{-2}$$

$$E = E^\circ - \frac{0.0591}{n} \log Q$$

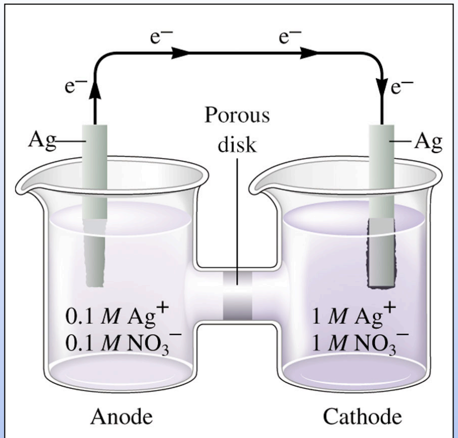
$$E = 1.10\text{V} - \frac{0.0591}{2} \log(10^{-2}) = 1.16\text{V}$$

$$E = E^\circ - \frac{0.0591}{n} \log Q$$

Current will flow until E = 0
Equilibrium

$$E^\circ = + \frac{0.0591}{n} \log K$$

$$\log K = \frac{nE^\circ}{0.0591}$$



Concentration Differences will lead to potential difference

.1 M Ag^+ (aq) and 1 M Ag^+ (aq)

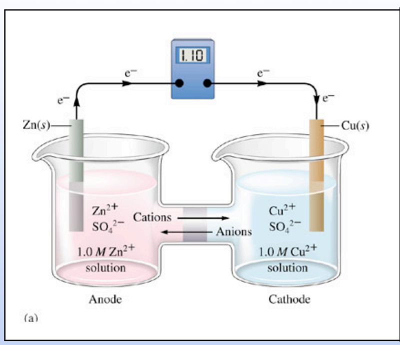
Same reaction! E° = 0 V

$$Q = \frac{[Ag^+]_{anode}}{[Ag^+]_{cathode}} = \frac{.1}{1} = .1$$

$$E = E° - \frac{0.0591}{n}\log Q$$

$$E = 0\,V - \frac{0.0591}{1}\log(.1) = 0.0591\,V$$

each factor of ten will be another 0.0591 V



If E < 0, then the reaction can be force in the non-spontaneous direction by applying a potential greater than E to the cell

F = 96,485 C

q = I x t