Slide 1

RedOx Chemistry

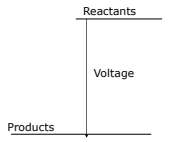
I balanced the equation and calculated the frigging voltage, now what?

Slide 2

Keep in mind what the voltage is:

It's the hill the electrons fall down!

If you catch 'em you can make them do work!



Slide 3

How much work can an electron do if an electron could do work?

Well, it's all a question of energy.

Joules is energy, volts is electrical potential.

Gravitation equivalent:

Joules is energy, height gravitational potential

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Joule = Coulomb * Volt

The charge * the potential = the energy change in the reaction

If 2 Coulombs falls through a 3 V electric field, 6 Joules of energy are liberated!

E = q * V

E = #e^-*1.602x10^{-19} C/e^- * V

q = charge

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How do I figure out the charge?

The particle that carries charge is the electron. The total charge is really just the number of electrons multiplied by the charge on each electron.

1 electron has a charge of 1.602x10^{-19} C

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Clicker question:

How much energy is available if a mole of electrons pass through a 9 V battery?

A. 9 Joules
B. 1.4x10^{-18} J
C. 8.7x10^{5} J
D. 5.4x10^{24} J
E. Loving you…

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$E = q * V$

$$1 \text{ mol } e^- * \frac{6.022 \times 10^{23} \text{ } e^-}{1 \text{ mol } e^-} * \frac{1.602 \times 10^{-19} \text{ C}}{1 \text{ } e^-}$$

$$= 9.6485 \times 10^4 \text{ C}$$

$E = 96{,}485 \text{ C} * 9 \text{ V} = 8.68 \times 10^5 \text{ J}$

(Remember a Joule = 1 CV)

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96485 C/mol e^-

That number is soooo important it has a name: Faraday's constant (F)

We'll see this a lot in this chapter.

It's the conversion factor from moles to Coulombs.

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If an electron falls in a forest…

Consider the first example from yesterday:

$2 Fe^{3+} + Sn^{2+} \rightarrow 2 Fe^{2+} + Sn^{4+}$

$E^0_{cell} = +0.62$ V

How much energy is liberated if 5.0 g of $FeCl_3$ undergoes the above reaction?

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If an electron falls in a forest…

$2\ Fe^{3+} + Sn^{2+} \rightarrow 2\ Fe^{2+} + Sn^{4+}$ $\quad E^0_{cell} = +0.62\ V$

Remember the ½ reaction:

$2\ Fe^{3+} + 2\ e^- \rightarrow 2\ Fe^{2+}$

This tells you how many electrons actually get moved per ion of iron!

How much energy is liberated if 5.0 g of $FeCl_3$ undergoes the above reaction?

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If an electron falls in a forest…

$2\ Fe^{3+} + Sn^{2+} \rightarrow 2\ Fe^{2+} + Sn^{4+}$

$E^0_{cell} = +0.62\ V$

$2\ Fe^{3+} + 2\ e^- \rightarrow 2\ Fe^{2+}$

It's all just UNITS! UNITS! UNITS!

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If an electron falls in a forest…

$2\ Fe^{3+} + 2\ e^- \rightarrow 2\ Fe^{2+}$

$$5.0\ g\ FeCl_3 * \left(\frac{1\ mol\ FeCl_3}{162.2\ g\ FeCl_3}\right) * \left(\frac{1\ mol\ Fe^{3+}}{1\ mol\ FeCl_3}\right) =$$

$$= 0.0308\ mol\ Fe^{3+} * \frac{2\ mol\ e^-}{2\ mol\ Fe^{3+}} * \frac{6.022 \times 10^{23}\ e^-}{1\ mol\ e^-}$$

$$= 1.855 \times 10^{22}\ e^- * \frac{1.602 \times 10^{-19}\ C}{1\ e^-} = 2972\ C$$

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2 Important Amounts

Energy:

2972 C * 0.62 V = 1842 CV = 1842 J

Current:
electrons = 1.855×10^{22} electrons
= 0.0308 moles electrons

Current = "amperes" = Coulombs per second

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A couple points:

Where does that 1842 J of energy go?

1. Into playing music on your iPod…
2. Heat…

Conservation of energy still applies:
1842 J = 300 J of iPod+250 J of wrist watch + 500 J of alarm clock + 792 J of waste heat…

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V, E, and I

Voltage is the potential the electron can fall through.

Energy (CV=J) tells you how much work you can do. It's what you get when an electron falls through the potential.

Current (I) Amperes tells you how many electrons it takes to do it and how long it takes to do it! (It's kind of the electrical equivalent of MOLES! MOLES! MOLES!)

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Clicker Question

Suppose I have 2 half-cells that I want to hook up to make a battery, what would be the voltage of the battery?

$I_{2\,(aq)} + 2\ e^- \rightarrow 2\ I^-_{(aq)}$ $E^0_{red} = 0.54$ V

$2\ H_2O_{(l)} + 2\ e^- \rightarrow H_{2\,(g)} + 2\ OH^-_{(aq)}$ $E^0_{red} = -0.83$ V

A. -0.29 V
B. +0.29 V
C. 1.37 V
D. – 1.37 V

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Clicker Question

Suppose I have 2 half-cells that I want to hook up to make a battery, what would be the voltage of the battery?

$I_{2\,(s)} + 2\ e^- \rightarrow 2\ I^-_{(aq)}$ $E^0_{red} = 0.54$ V

$2\ H_2O_{(l)} + 2\ e^- \rightarrow H_{2\,(g)} + 2\ OH^-_{(aq)}$ $E^0_{red} = -0.83$ V

$I_{2\,(s)} + 2\ e^- \rightarrow 2\ I^-_{(aq)}$ $E^0_{red} = 0.54$ V

$H_{2\,(g)} + 2\ OH^-_{(aq)} \rightarrow 2\ H_2O_{(l)} + 2\ e^-$ $E^0_{ox} = +0.83$ V

Every redox reaction MUST have both an oxidation and a reduction! E_{cell} must be + for it to be a spontaneous reaction.

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Clicker 2

Suppose I'm going to use my brand new 1.37 V battery to do work (assume the V is constant). If I need to do 1000 J of work, how much I_2 will I need?

$I_{2\,(s)} + H_{2\,(g)} + 2\ OH^-_{(aq)} \rightarrow 2\ I^-_{(aq)} + 2\ H_2O_{(l)}$

$E^0_{cell} = 1.37$ V

A. 4.63×10^4 g
B. 1.92 g
C. 9.26×10^4 g
D. 0.96 g
E. 8×10^1 g

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$E = q*V$

$1000\ J = q*V = q*1.37\ V$

$q = 7.299 \times 10^{2}\ C$

$7.299 \times 10^{2}\ C * \frac{1\ e^-}{1.602 \times 10^{-19}\ C} = 4.56 \times 10^{21}\ e^-$

$4.56 \times 10^{21}\ e^- * \frac{1\ mol}{6.022 \times 10^{23}\ e} = 7.566 \times 10^{-3}\ mol\ e^-$

$7.566 \times 10^{-3}\ mol * \frac{1\ mol\ I_2}{2\ mol\ e^-} = 3.78 \times 10^{-3}\ mol\ I_2$

$3.78 \times 10^{-3}\ mol\ I_2 * \frac{253.8\ g}{mol\ I_2} = 0.959\ g\ I_2$

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So, to do 1000 J of work requires about 1 g of I_2 to be consumed…suppose I don't have 1 g of I_2 in my battery?

Yes, that's right, your battery dies!!!

But even before it dies, it starts to decay!!!

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How long will my 1.37 V battery last?

Not very!

$E^0_{cell} = 1.37\ V$

What's the "naught" for?
1 atm, 298 K, and…

1 M concentration!!!
As soon as I start using up the I_2, it's not E^0 anymore!!!

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Suppose I'm not at standard conditions!

Turns out that ΔG depends on E_{cell} (and vice versa):

$\Delta G = - nFE_{cell}$

F = Faraday's constant = 96,485 C/mol

n = # of moles of electrons transferred (from balanced equation)

Correcting E can be done by correcting ΔG!!!

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Change in Temperature:

$\Delta G = \Delta H^0 - T \Delta S^0$

We did this when we were talking about K because, you might recall:

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

Compare with:
$\Delta G^0 = - nF E^0$

Don't you love it when the world is symmetrical

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The most common correction is for concentration

As you run an electrochemical cell, you are constantly oxidizing (reducing) the chemical species.

That means that you are only E^0 for 1 nanosecond because then the concentrations are no longer molar!

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Nernst already did the algebra

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

$\Delta G^0 = - nF E^0$

Remember the ΔG^0 correction for concentration?

$\Delta G = \Delta G^0 + RT \ln Q$
(Q is "K" when you're not at equilibrium)

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Nernst already did the algebra

$\Delta G^0 = - nF E^0$

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

$E = E^0 - \frac{RT}{nF} \ln Q$ (Nernst Equation)

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Different versions, same equation:

R = 8.314 J/mol K
T = 298 K
F = 96485 C/mol e^-

$E_{cell} = E^0_{cell} - \frac{0.0257}{n} \ln Q$
You might also recall that:
ln Q = 2.303*log Q

$E_{cell} = E^0_{cell} - \frac{0.0592}{n} \log Q$

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Clicker Question

Suppose I have 2 half-cells that I want to hook up to make a battery, what would be the voltage of the battery if the concentration of I_2, I^-, Br^-, and Br_2 are each 1 M?

$I_{2\,(s)} + 2\ e^- \rightarrow 2\ I^-_{(aq)}$ $\quad E^0_{red} = 0.54$ V

$Br_{2\,(aq)} + 2\ e^- \rightarrow 2\ Br^-_{(aq)}$ $\quad E^0_{red} = 1.09$ V

A. 0.68 V
B. 1.52 V
C. 1.63 V
D. 0.55 V

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Clicker Question

Suppose I have 2 half-cells that I want to hook up to make a battery, what would be the voltage of the battery if the concentration of I_2, I^-, Br^-, and Br_2 are each 1 M?

$2\ I^-_{(aq)} \rightarrow I_{2\,(s)} + 2\ e^-$ $\quad E^0_{ox} = -0.54$ V

$Br_{2\,(aq)} + 2\ e^- \rightarrow 2\ Br^-_{(aq)}$ $\quad E^0_{red} = 1.09$ V

$E^0_{cell} = 1.09 + (-0.54) = 0.55$ V

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Clicker Question

Suppose I have 2 half-cells that I want to hook up to make a battery, what would be the voltage of the battery if the concentration of I_2 and I^- and Br_2 and Br^- are each 0.1 M?

$I_{2\,(s)} + 2\ e^- \rightarrow 2\ I^-_{(aq)}$ $\quad E^0_{red} = 0.54$ V

$Br_{2\,(aq)} + 2\ e^- \rightarrow 2\ Br^-_{(aq)}$ $\quad E^0_{red} = 1.09$ V

A. 0.60 V
B. 0.52 V
C. 0.58 V
D. 0.55 V
E. 0

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$2\ I^-_{(aq)} \rightarrow I_{2\ (s)} + 2\ e^-$ $\quad E^0_{ox} = -0.54\ V$

$Br_{2\ (aq)} + 2\ e^- \rightarrow 2\ Br^-_{(aq)}$ $\quad E^0_{red.} = 1.09\ V$

$2\ I^-\ (aq) + Br_2\ (aq) \rightarrow I_2(s) + 2\ Br^-\ (aq)$

$Q = \frac{[Br-]^2}{[I-]^2\,[Br_2]} = \frac{[0.1]^2}{[0.1]^2\,[0.1]} = 10$

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$E_{cell} = E^0_{cell} - \frac{0.0592}{n} \log Q$

$E_{cell} = 0.55\ V - \frac{0.0592}{2} \log 10$

$E_{cell} = 0.55\ V - 0.0296 = 0.52\ V$

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A Nernst Example

What is the standard cell potential of:

$2I^-_{(aq)} + Br_{2\ (aq)} \rightarrow I_{2\ (s)} + 2\ Br^-_{\ (aq)}$

We already solved this!

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Split into ½ reactions

$2I^-_{(aq)} \rightarrow I_{2\,(s)} + 2\,e^-$ $E^0_{ox} = -E^0_{red} = -0.54$ V

$Br_{2\,(aq)} + 2\,e^- \rightarrow 2\,Br^-_{(aq)}$ $E^0_{red} = 1.09$ V

$E^0_{cell} = 1.09\text{ V} - 0.54\text{ V} = +0.55$ V

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A Nernst Example

What is the cell potential of when the concentrations are all 0.100 M:

$2I^-_{(aq)} + Br_{2\,(aq)} \rightarrow I_{2\,(s)} + 2\,Br^-_{(aq)}$

$$E = E^0 - \frac{0.0592}{n} \log Q$$

What's "n"?

Look at the ½ reactions!

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Split into ½ reactions

$2I^-_{(aq)} \rightarrow I_{2\,(s)} + 2\,e^-$ $E^0_{ox} = -E^0_{red} = -0.54$ V

$Br_{2\,(aq)} + 2\,e^- \rightarrow 2\,Br^-_{(aq)}$ $E^0_{red} = 1.09$ V

$E^0_{cell} = 1.09\text{ V} - 0.54\text{ V} = +0.55$ V

2 electrons are being exchanged, so n=2!

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A Nernst Example

What is the cell potential of when the concentrations of I_2 and I-, Br_2 and Br^- are each 0.1 M?

$2I^-_{(aq)} + Br_{2\,(aq)} \rightarrow I_{2\,(s)} + 2\,Br^-_{(aq)}$

$E = E^0 - \frac{0.0592}{n} \log Q$

What's Q?

$Q = \frac{[Br^-]^2}{[I^-]^2[Br_2]}$

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A Nernst Example

What is the cell potential of when the concentrations of I_2 and I-, Br_2 and Br^- are each 0.10 M:

$2I^-_{(aq)} + Br_{2\,(aq)} \rightarrow I_{2\,(s)} + 2\,Br^-_{(g)}$

$E = E^0 - \frac{0.0592}{n} \log \frac{[Br^-]^2}{[I^-]^2[Br_2]}$

$E = 0.55\ V - \frac{0.0592}{2} \log \frac{0.10]^2}{[0.100]^2[0.10]}$

$E = 0.55\ V - \frac{0.0592}{2} \log 10 = 0.55\ V - 0.0296\ V$

$E = 0.52\ V$ (for a split second!)

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A Nernst Example

If the aqueous concentrations of I_2 and I- are 0.1 M while Br_2 and Br^- are each 0.1 M with a volume of 2 L in each half-cell, how much water would need to be added to the iodine half-cell to create a cell potential of 0.50 V:

$2I^-_{(aq)} + Br_{2\,(aq)} \rightarrow I_{2\,(s)} + 2\,Br^-_{(aq)}$

$E = E^0 - \frac{0.0592}{n} \log \frac{[Br^-]^2}{[I^-]^2[Br_2]}$

$0.50\ V = 0.55\ V - \frac{0.0592}{2} \log \frac{[0.1]^2}{[x]^2[0.100]}$

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A Nernst Example

$0.50\ V = 0.55\ V - \frac{0.0592}{2} \log \frac{[0.1]^2}{[x]^2[0.100]}$

$-0.05\ V = -0.0296 \log 0.1/x^2$

$1.689 = \log (0.1/x^2)$

$10^{1.689} = 10^{\log(0.1/x2)}$

$48.89 = 0.1/x^2$

$x^2 = 0.00204$

$X = 0.0452\ M$

Is that the answer?

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Course not!

I start with 2 L of 0.100 M I- and I_2. I need to get the concentration down to 0.0204 M. How much water do I add?

It's just a dilution problem:

$M_1V_1=M_2V_2$

$(0.100\ M)\ (2\ L) = (0.0452\ M)*(X\ L)$

$X= 4.7\ L$

So, I would need to add 2.7