

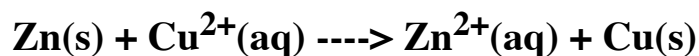
This is BCE#30.

I recommend you print out this page and bring it to class. [Click here](#) to show a set of five BCE30 student responses randomly selected from all of the student responses thus far in a new window.

John , here are [your responses](#) to the BCE and the [Expert's response](#).

For this BCE we will use this short version of the [Standard Reduction Potential Table](#). You might want to print the table out before beginning the BCE.

1. Calculate the standard cell potential for the reaction



$$E^{\circ} = 1.10 \text{ Volts}$$

75%

$$E^{\circ} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = 0.34 \text{ v} - (-0.76 \text{ v}) = +1.10 \text{ volts}$$

2. Write the equilibrium expression for the reaction in Question 1. (Remember how we handle substances in the solid phase.)

$$K = K = [\text{Zn}^{2+}]/[\text{Cu}^{2+}] \text{ (Remember this expression is also equal to Q) } 66\%$$

$K = [\text{Zn}^{2+}]/[\text{Cu}^{2+}]$ NOTE: since Cu and Zn are solids they do not appear in the equilibrium expression.

3. When we prepare an electrochemical cell using standard conditions the concentration of ions is 1 M at 298 K. When we measure the voltage the reaction is not at equilibrium. Calculate the value of Q for the reaction when the electrochemical half cells is first connected and the concentrations of all ions is 1.00 M.

$$Q = 1 \quad 75\%$$

$$Q = [\text{Zn}^{2+}]/[\text{Cu}^{2+}] = (1)/(1) = 1$$

4. Is this reaction thermodynamically favored? (Y/N)

Yes

75%

Yes, the reaction is spontaneous because E° is positive.

5. What is the value of K for the reaction in Question 1.

$$(\Delta G^{\circ} = -nFE^{\circ} \text{ where } F = 96,500 \text{ J volt}^{-1})$$

$$\text{and } \Delta G^{\circ} = -RT \ln K \text{ where } R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$$

$$\text{so } E^{\circ} = RT/nF \ln K \text{ or } E^{\circ} = 0.0257/n \ln K$$

n is the number of electrons transferred in the balanced chemical equation)

$$K = 1.50e37 \quad 66\%$$

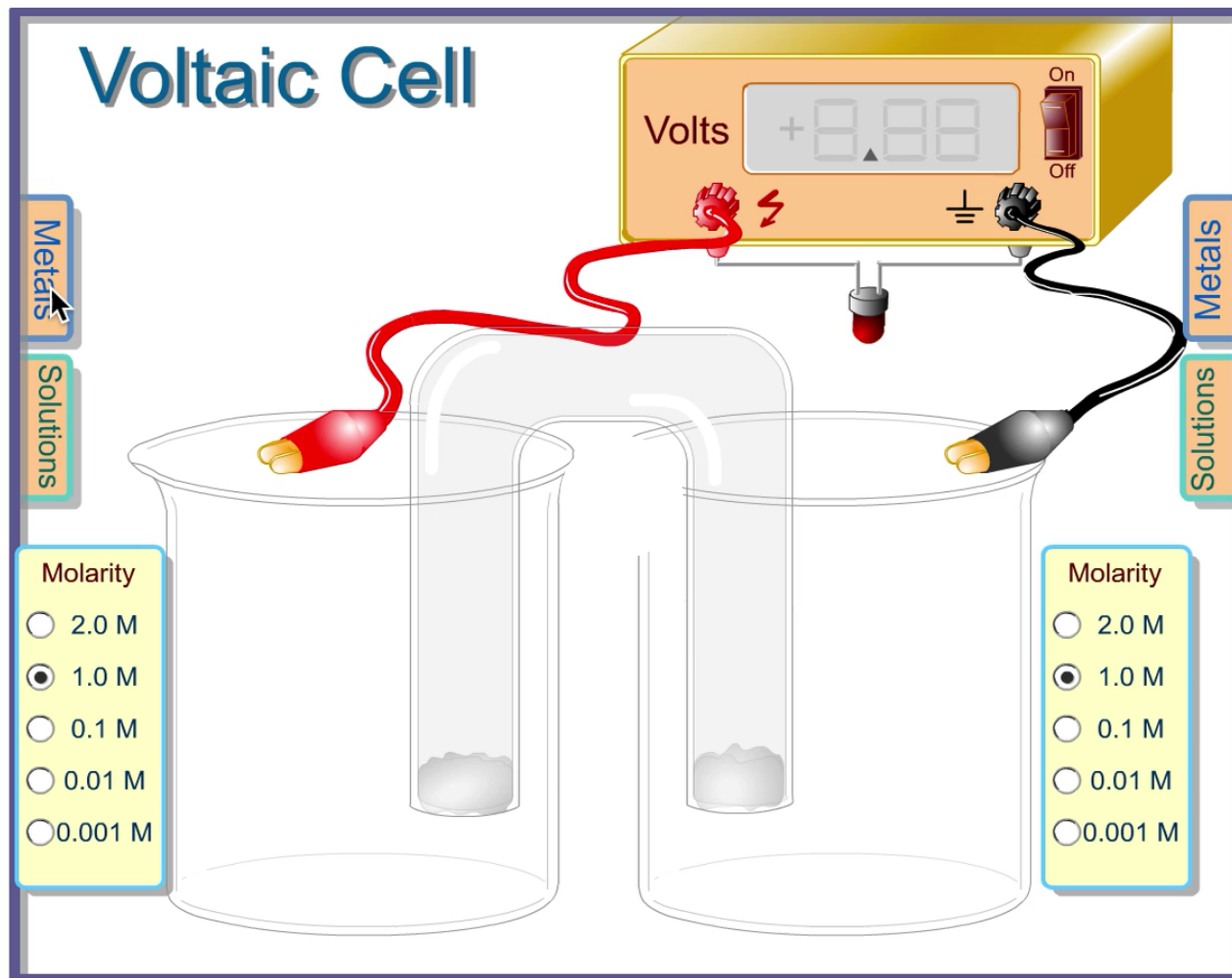
$\Delta G^{\circ} = -RT \ln K$ and $\Delta G^{\circ} = -nFE^{\circ}$ therefore $-nFE^{\circ} = -RT \ln K$ and $E^{\circ} = (RT/nF) \ln K$.
Substituting we have,

$$+1.1 \text{ v} = -(8.314 \text{ J/mol K} \cdot 298 \text{ K} / 2 \text{ mol e}^{-} \cdot 96500 \text{ J/V}) \ln K$$

$$\ln K = +1.1 \text{ v} / (0.0257/2 \text{ mol e}^{-}) = 85.69$$

$$K = 1.6 \times 10^{37}$$

Look at the QuickTime movie below. Play it, and answer the questions that follow based on the movie.



6. Suppose we prepared an electrochemical cell based on the same reaction, but instead of 1 M Zn^{2+} we made the $[\text{Zn}^{2+}] = 0.1 \text{ M}$ and we'll leave the $[\text{Cu}^{2+}] = 1.0 \text{ M}$. Watch the movie above to see this electrochemical cell and enter the voltage produced.

$E = 1.13 \text{ volts}$ 92%

$E = 1.13 \text{ volts}$

7. Use the equation

$$E_{\text{cell}} = E^{\circ} - (0.0257/n) \ln [\text{Zn}^{2+}]/[\text{Cu}^{2+}]$$

(NOTE: $n = \#$ of electrons transferred)

calculate E_{cell} for the reaction. How does your calculated value compare to the measured value in Question 6.

Ecell = 1.13 volt is what I calculated. It agrees with the movie. Interesting that when the concentration of a product is decreased the cell potential increases.

$$E = 1.10 \text{ volts} - (0.0257/2 \text{ mol } e^-) \ln [Zn^{2+}]/[Cu^{2+}]$$

$$E = 1.10 \text{ volts} - (0.0257/2 \text{ mol } e^-) \ln (0.1/1.0)$$

$$E = 1.10 \text{ volts} - 0.0128 \ln 0.1$$

$$E = 1.10 \text{ volts} - 0.0128 (-2.303)$$

$$E = 1.10 \text{ volts} - (-0.03) = 1.13 \text{ volts}$$

different equations?

The calculated value and the measured value are the same. I did the calculation right!!

8. Is there anything about the questions that you feel you do not understand? List your concerns/questions.

nothing

9. If there is one question you would like to have answered in lecture, what would that question be?

nothing