Q: How do changing conditions affect the voltage of a cell?

The Nernst Equation has been removed from the AP Exam as of 2014. The actual calculation of ΔE at alternate (non-standard) conditions has been minimized. However being able to generalize Voltage as conditions change is extremely important.

Nernst Equation

$E = E° - (RT/nF)\ln Q$ $Q = [P]/[R]$ or $E = E° - (0.0592/n)\log Q$

@ equilibrium: $\log K = nE°/.0592$

Objective:

- How does changing concentrations affect voltage of reaction?
- How does changing temperature affect voltage?
- How does changing the posts affect the voltage?

1. (Brady882) A cell employs the following half reactions

   $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ $E = -.25V$
   $Cr^{3+}(aq) + 3e^- \rightarrow Cr_{(s)}$ $E = -.74 V$ ← Flip

   $3 Ni^{2+} + 2Cr \rightarrow 3Ni + 2Cr^{+3}$

   a) Calculate E°. $\left.\begin{matrix}-.25\\ +.74\end{matrix}\right\}$ 0.49V

   b) As the reaction proceeds, what will happen to the magnitude of E? Goes to zero

   c) What will be the value of E when the reaction goes to equilibrium? zero

   d) What happens to the concentration of each reactant as the reaction proceeds?

   $[Ni^{2+}]$ = ? ↓
   $[Cr^{3+}]$ = ? ↑

   e) How can you adjust the concentrations to increase voltage above E°? yes, Reverse

   STP = 1M, so increase $[Ni^{2+}]$ ↑ 1M or ↓ $[Cr^{3+}]$ Below 1M or Both

   f) Using the Nernst equation calculate the new voltage.
   Calculate the cell potential when $[Ni^{2+}] = 1.0E\text{-}4$ M $[Cr^{3+}] = 2.0E\text{-}3$M.

   $Q = \frac{[Cr^{+3}]^2}{[Ni^{2+}]^3}$

   $4.0E6 = \frac{[2.0E\text{-}3]^2}{[1.0E\text{-}4]^3}$

   $0.42V = 0.49V - \underbrace{\frac{0.0592}{6}\log(4E6)}_{.065}$

   total e^- → 6

2. A reaction is set up with solid (Cu and 1M = Cu^{2+}) and solid (zinc and 1M = Zn^{2+})

   a) Write out a balanced reaction for a spontaneous reaction between these reactants and determine the voltage.

   $Cu^{+2} + Zn \rightarrow Cu^0 + Zn^{2+}$ V = 1.1V

   b) Calculate Q if we are at standard conditions.

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

   c) Calculate Q if we add water to both half cells to lower the concentrations to .5M.

   $\frac{.5}{.5} = 1$

   d) Would the voltage change?

   NO,

STP Q = 1

$\frac{1}{1^2} = 1$

3. A reaction is set up with $2Ag^{+1} + Zn \Rightarrow 2Ag + Zn^{2+}$ V = 1.56V

a) Calculate Q if we add water to both half cells to lower the concentrations to 0.5M.

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

b) Would or the voltage change, if so how? Down

c) ~~How can you alter conditions to reverse the condition in~~ "a"?

Start Q = 1 Now Q = 2

K = very large

$\frac{.5}{(.5)^2} = \frac{.5}{.25}$ = 2

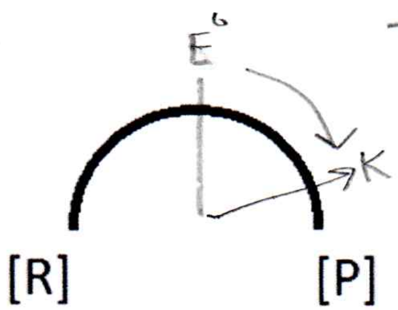
voltage = 0

4. A reaction is set up with $Cu^{2+} + Mg \Rightarrow Cu + Mg^{2+}$ V = 2.71V

a) What is the difference between E° and E? Some other Condions so V is diffent

STP V = 2.71

b) What is the relationship between and E° and K? Label the picture below.



+E° = K ↑ than 1

larger E° means more product favored to larger K.

5) What is the relationship between E° and E and K

$\log K = \frac{nE°}{.0592}$

a) If a reaction has a standard voltage of 0.5V, what is the K?

$\frac{2(0.5)}{0.0592} = \log K$

K = ____

use inverse log.

try it.

b) If a reaction has a +E° which direction will the reaction proceed?

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c) If a reaction has a +E which direction will the reaction proceed?

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