Nernst Equation

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The *Nernst Equation* enables one to determine electromotive forces (emf) of many processes, for instance the resting potential of cell membranes. We can then deduce the*biological standard potentials* which are important in studying biological process such as action potential during a spike of a neuron in response to a stimulus.

Introduction

The *Nernst Equation* is derived from the emf and the Gibbs energy **under non-standard conditions**.

$$E_o = E_{o reduction} - E_{o oxidation}$$

When E_o is positive, the reaction is spontaneous. When E_o is negative, the reaction is not spontaneous. Since the change inGibbs free energy, ΔG, is also related to spontaneity of a reaction, therefore, ΔG and E are related. Specifically,

$$\Delta G = -nFE$$

where, n is # of electrons transferred in the reaction, F is the Faraday constant (96500 C/mol) and E is potential difference. Under standard conditions, this equation is then

$$\Delta G_o = -nFE_o.$$

Since,

$$\Delta G = \Delta G_o + RT \ln Q \quad (1)$$

Substituting $\Delta G = -nFE$ and $\Delta G_o = -nFE_o$ into equation (1), we have:

$$-nFE = -nFE_o + RT \ln Q$$

Divide both sides of the equation above by $-nF$, we have

$$E=E_o - RTnF\ln Q(2)$$

Equation (2) can be rewritten in the form of log base 10:

$$E=E_o - 2.303RTnF\log Q(3)$$

At standard temperature T = 298K, the $2.303RTF$ equals .0592 V, so equation (3) turns into:

$$E = E_o - .0592Vn\log Q$$

The equation above indicates that the electrical potential of a cell depends upon the reaction quotient Q of the reaction. As the redox reaction proceeds, reactants are consumed, thus concentration of reactants decreases. Conversely, the products concentration increases due to the increased in products formation. As this happens, cell potential gradually decreases until the reaction is atequilibrium, at which $\Delta G = 0$.

Example 1

For the Zn-Cu redox reaction:

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)} E^o_{cell} = +1.10\ V$$

SOLUTION

Initially, $[Cu^{2+}] = [Zn^{2+}] = 1.0$ M at standard T = 298K As the reaction proceeds, $[Cu^{2+}]$ decreases as $[Zn^{2+}]$ increases. Lets say after one minute, $[Cu^{2+}] = 0.05$ M while $[Zn^{2+}] = 5.0$ M. According to Nernst, cell potential after 1 minute is:

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

$$E = 1.10V - \frac{0.0592V}{2}\log\frac{5.0\ M}{.05\ M}$$

$$E = 1.04\ V$$

As you can see, the initial cell potential is $E = 1.10\ V$, after 1 minute, the potential drops to 1.04 V. As the reaction continues to progress, more Cu^{2+} will be consumed and more Zn^{2+} will be generated. As a result, the cell potential continues to decrease and when the cell potential drops down to 0, the concentration of reactants and products stops changing. This is when the reaction is at equilibrium.

At equilibrium, the reaction quotient $Q=K_{eq}$. Also, at equilibrium, $\Delta G = 0$ and $\Delta G = -nFE$, so $E = 0$.

Therefore, substituting $Q=K_{eq}$ and $E = 0$ into the Nernst equation, we have:

$$0=E_o - RTnF\ln K_{eq}$$

At standard conditions, the equation above simplifies into:

$$0=E_o - 0.0592n\log K_{eq}$$

This equation can be rearranged into:

$$\log K_{eq}=nE_o0.0592$$

The equation above indicates that the equilibrium constant **K_{eq} is proportional to the standard potential** of the reaction. Specifically, when:

- $K>1, E_o>0$, reaction favors products formation.
- $K<1, E_o<0$, reaction favors reactants formation.

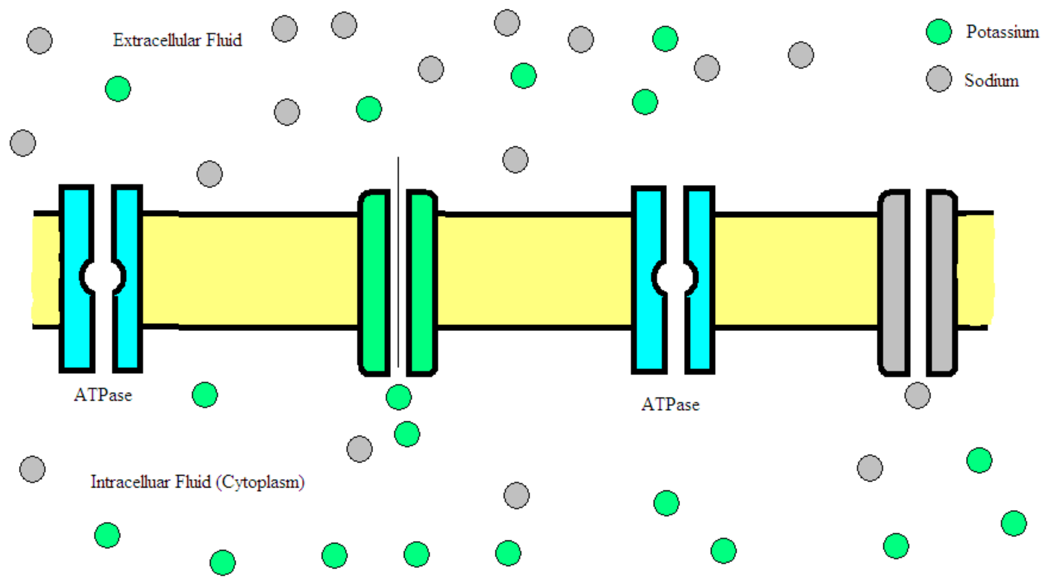
This result fits Le Châtlier's Principle,which states that when a system at equilibrium experiences a change, the system will minimize that change by shifting the equilibrium in the opposite direction.

Biological Application of the Nerst Equation

The emf can be determined measuring the current in a galvanic cell. This method is important when determining the standard potential of oxidizing agents. But our interest is the *biological standard potentials* such as neurons firing.

Action Potential of a Neuron Cell

Perhaps the most fascinating system in our body is the *nervous system*. The neuron is the basic operating unit of the *nervous system*and its mechanism is still under a lot of research. Electrolytes *Sodium* and *Potassium* are the most prevalent electrolytes in our body. Intracellular concentrations of *Potassium* is higher inside the cell in comparison to extracellular. For *Sodium* it's the opposite more extracellular and less intracellular. This difference in concentration is the mechanics of how a neuron and other cells have resting potentials. ATPase pumps and selective membrane channels allow for the concentration gradient difference. With the *Nernst Equation*we can deduce the membrane potential of a neuron for our discussion. The concentration of potassium inside is 150 mM, and 15 mM outside. Plugging these values in the equation above out comes the neuron potential which can be depolarize in response to ...



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Outside Links

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