

# The Nernst Equation

Not all systems are at equilibrium when measured. Slow kinetics, adding or removing of substrate, and other factors can keep the system from attaining equilibrium - for a while.

Values of  $E^\circ$  are applicable *only* to equilibrium conditions. To calculate cell potentials for non-equilibrium systems, we use the **Nernst Equation**:

$$E = E^\circ - (RT/nF) \ln Q$$

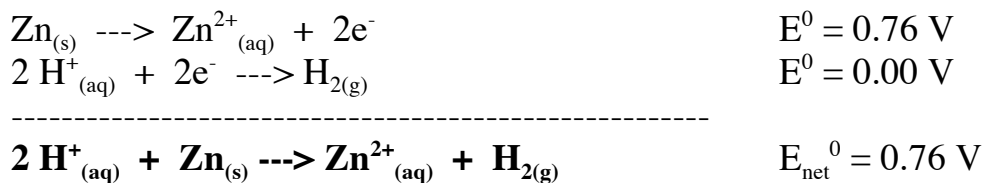
where

- $E$  = cell potential (V) at nonstandard conditions
- $E^\circ$  = cell potential (V) at standard equilibrium conditions
- $R$  = Gas constant (8.314 J/K mol)
- $T$  = Temperature (K)
- $F$  = Faraday constant (96,485 C/mol  $e^-$ )
- $n$  = the number of moles of electrons transferred
- $Q$  = reaction quotient; for the reaction:  $a A_{(aq)} + b B_{(aq)} \rightleftharpoons c C_{(aq)} + d D_{(aq)}$ ,

$$\text{then } Q = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

*Example:* Find  $E$  when  $[Zn^{2+}] = 0.0010 \text{ M}$ ,  $P(H_2) = 0.10 \text{ atm}$  and  $pH = 0$  at  $298 \text{ K}$ .

$$[H^+] = 10^{-pH} \text{ M} = 10^{-0} \text{ M} = \mathbf{1.0 \text{ M}}$$
; note that  $n = 2$



Use the Nernst equation to find  $E$ :

$$E = E^\circ - (RT/nF) \ln Q$$

$$Q = \frac{[Zn^{2+}] P_{H_2}}{[H^+]^2} = [0.0010 * 0.10 / (1.0)^2] = 0.00010$$

$$E = 0.76 \text{ V} - (8.314 * 298 \text{ K} / 2 * 96500) \ln 0.00010$$

$$E = 0.76 \text{ V} + 0.12 \text{ V} = \mathbf{0.88 \text{ V}}$$