

# AP CHEMISTRY

## TOPIC 11: ELECTROCHEMISTRY, PART B,

Day 126:

- Voltaic Cells (Galvanic Cells)

1. Explain why is necessary to use a salt bridge or a porous disk in a galvanic cell (voltaic cell) ?

**The salt bridge allows “counter ions” to flow into the two cell compartments to maintain electrical neutrality. Without a salt bridge (or porous disk), no sustained electron flow (or charged particles) can occur.**

2. Define the following:

a) Cathode

**The electrode at which reduction occurs.**

b) Anode

**The electrode at which oxidation occurs.**

c) oxidation half-reaction

**Oxidation half-reaction: The half-reaction in which electrons are products. In a galvanic cell (voltaic cell), the oxidation half-reaction always occurs at the anode.**

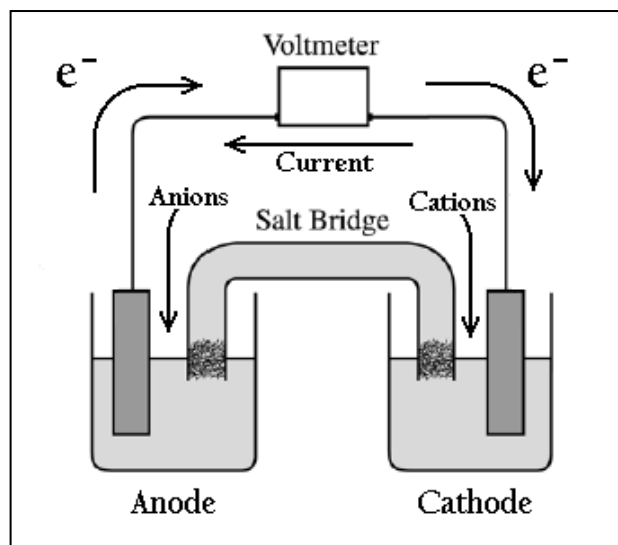
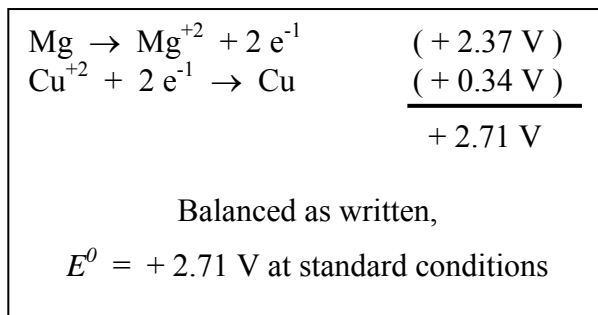
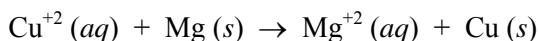
d) reduction half-reaction

**Reduction half-reaction: The half-reaction in which electrons are reactants. In a galvanic cell (voltaic cell), the reduction half-reaction always occurs at the cathode.**

3. Why isn't the value of the cell potential multiplied by the same integer used to multiply a half-reaction in balancing the equation for a cell reaction?

**Standard Reduction Potentials are an intensive property. As long as the concentrations of ions and gases are 1 M or 1 atm, standard reduction potentials and standard oxidation potentials are a constant and not dependent on the coefficients in the balanced equation.**

4. Sketch the galvanic cell based on the following overall reactions. Show the direction of electron flow, the direction of ion migration through the salt bridge, and identify the cathode and anode. Show the correct half-reactions and  $E^0$  for each half-reaction. Assume the conditions are standard.



5. Answer the following questions using the data from the Standard Reduction Potentials at 25°C.

a) Is  $\text{H}^+$  (*aq*) capable of oxidizing  $\text{Cu}$  (*s*) to  $\text{Cu}^{2+}$  (*aq*), explain why.

$2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{H}_2 \quad (0.00 \text{ V})$ $\text{Cu} \rightarrow \text{Cu}^{2+} + 2 \text{e}^- \quad (-0.34 \text{ V})$ <hr style="width: 50%; margin-left: auto; margin-right: 0;"/> $\text{NO !!!} \qquad \qquad \qquad -0.34 \text{ V}$	<p>No, <math>E^0</math> (for the reaction) has a negative value. When <math>E^0</math> is negative, <math>\Delta G</math> is positive (non-spontaneous)</p>
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b) Is  $\text{Fe}^{3+}$  (*aq*) capable of oxidizing  $\text{I}^-$  (*aq*), explain why.

$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+} \quad (+0.77 \text{ V})$ $2 \text{I}^- \rightarrow \text{I}_2 + 2 \text{e}^- \quad (-0.53 \text{ V})$ <hr style="width: 50%; margin-left: auto; margin-right: 0;"/> $\text{YES !!!} \qquad \qquad \qquad +0.24 \text{ V}$	<p>Yes, <math>E^0</math> (for the reaction) has a positive value. When <math>E^0</math> is positive, <math>\Delta G</math> is negative (spontaneous). Don't get "worked up" that we have a metal with a non-metal. This is not a single replacement reaction – 2 chambers!</p>
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c) Is  $\text{Fe}^{2+}$  (*aq*) capable of reducing  $\text{Cr}^{3+}$  (*aq*) to  $\text{Cr}^{2+}$  (*aq*), explain why.

$\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^- \quad (-0.77 \text{ V})$ $\text{Cr}^{3+} + \text{e}^- \rightarrow \text{Cr}^{2+} \quad (-0.41 \text{ V})$ <hr style="width: 50%; margin-left: auto; margin-right: 0;"/> $\text{NO !!!} \qquad \qquad \qquad -1.18 \text{ V}$	<p>No, <math>E^0</math> (for the reaction) has a negative value. When <math>E^0</math> is negative, <math>\Delta G</math> is positive (non-spontaneous). Read the question carefully – be able to distinguish what the ions are doing.</p>
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6. In a galvanic cell the concentration of  $\text{Pb}^{2+}$  is changed from 1.0 M to a 0.25 M, and the concentration of  $\text{Mg}^{2+}$  is changed from 1.0 M to a 0.10 M. Temperature is at 25°C.

a) First predict the oxidizing agent and the reducing agent for the spontaneous reaction in the cell.

$\text{Mg} \rightarrow \text{Mg}^{2+} + 2 \text{e}^- \quad (+2.37 \text{ V})$ $\text{Pb}^{2+} + 2 \text{e}^- \rightarrow \text{Pb} \quad (-0.13 \text{ V})$ <hr style="width: 50%; margin-left: auto; margin-right: 0;"/> $\text{"Best" Voltage} \qquad \qquad \qquad +2.24 \text{ V}$	<p>Oxidized, Reducing Agent Reduced, Oxidizing Agent</p>
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b) Calculate the cell potential for the galvanic cell at standard conditions (use the Nernst Equation – do it !!!)

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

$$E_{\text{cell}} = 2.24 \text{ V} - \frac{0.0592}{2} \log \left( \frac{(1.0)^1}{(1.0)^1} \right); \quad \log 1 = 0$$

$$E_{\text{cell}} = 2.24 \text{ V} - 0 = 2.24 \text{ V}$$

b) Calculate the cell potential for the galvanic cell under the new concentrations (changed concentrations).

$$E_{\text{cell}} = 2.24 \text{ V} - \frac{0.0592}{2} \log \left( \frac{(0.10)^1}{(0.25)^1} \right)$$

$$E_{\text{cell}} = 2.24 \text{ V} - (-0.011779) = 2.25 \text{ V}$$