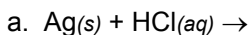


**Unit 6: Redox Reactions and Electrochemistry**

## Practice Set 5: Electrolytic Cells

1. Using either an activity series (Table 11.2) or Table of Standard Reduction Potentials, predict whether a reaction takes place, and if so, give a balanced reaction equation.

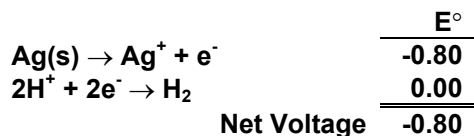


**Answer:**

**Using Activity Series - no reaction. Ag is lower on the list than H**

**Using Standard Reduction Potentials**

**Write out half-reactions in a way that will produce the equation in question (that is with solid Ag and aqueous  $\text{H}^+$  ions as reactants ( $\text{Cl}^-$  will be a spectator ion). Remember to reverse the sign of  $E^\circ$  if you reverse an equation:**



**Conclusion – since voltage ( $E^\circ$ ) is negative, the reaction WILL NOT BE spontaneous. No reaction will occur.**

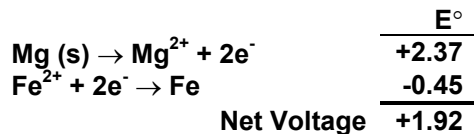


**Answer:**

**Using Activity Series - a reaction will occur. Mg is higher on the list than Fe**

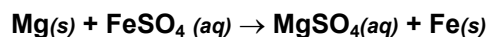
**Using Standard Reduction Potentials**

**$\text{SO}_4^{2-}$  will be a spectator ion.**



**Conclusion – since a positive voltage results, the reaction WILL BE spontaneous.**

**The balanced equation for the reaction will be:**



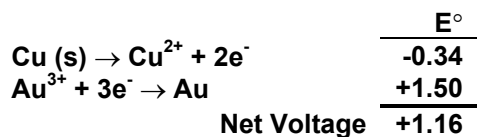


**Answer:**

**Using Activity Series - a reaction will occur. Cu is higher on the list than Au**

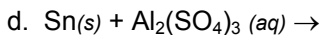
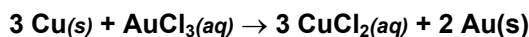
**Using Standard Reduction Potentials**

**$\text{Cl}^-$  will be a spectator ion.**



**Conclusion – since a positive voltage results, the reaction WILL BE spontaneous.**

**The balanced equation for the reaction will be:**

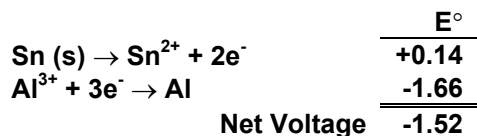


**Answer:**

**Using Activity Series - no reaction. Sn is lower on the list than Al**

**Using Standard Reduction Potentials**

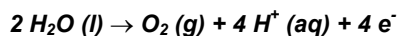
**$\text{SO}_4^{2-}$  will be a spectator ion.**



**Conclusion – since the voltage is negative, this reaction WILL NOT BE spontaneous.**

2 Consider the electrolysis of water. Describe the events at the anode in terms of:

a. the reaction ***Oxidation always occurs at the anode. Thus the anode reaction is:***



b. pH  ***$\text{H}^+$  are being produced. As  $[\text{H}^+]$  increases, pH decreases (the solution becomes more acidic)***

c. gas produced  ***$\text{O}_2$  gas is produced at the anode***