WORKSHEET: REDOX REACTIONS - SOLUTIONS

Oxidation-Reduction

1. Oxidation is the loss of electrons. Reduction is the gain of electrons. For an oxidation reaction, the electrons are products.

2. a) The Ag⁺ is being reduced. It is gaining electrons: \( 2 \text{Ag}^+ + 2 e^- \rightarrow 2 \text{Ag}^0 \).
   b) The Ag⁺ is also the oxidizing agent. It causes Cu⁺ to be oxidized:
      \( \text{Cu}^+ \rightarrow \text{Cu}^{2+} + 2 e^- \).
   c) The Cu⁺ is being oxidized.
   d) The Cu⁺ is also the reducing agent. It provides the electrons to reduce Ag⁺.

3. a) \( \begin{array}{c}
F_2 + 2e^- \rightarrow 2F^- \\
\text{Al}^{3+} + 3e^- \rightarrow \text{Al} \\
\text{Li}^+ + e^- \rightarrow \text{Li}
\end{array} \) 2.87 V
   -1.66 V
   -3.05 V
   c) The strongest oxidizing agent is F₂. The strongest reducing agent is Li.
   d) i \( \text{Al}^{3+} + 3 \text{Li} \rightarrow \text{Al} + 3 \text{Li}^+ \quad \xi^o = (-1.66 + 3.05) \quad V = +1.39 \text{ V spontaneous} \)
   ii \( 2 \text{Al}^{3+} + 6 \text{F}^- \rightarrow 2 \text{Al} + 3 \text{F}_2 \quad \xi^o = (-1.66 - 2.87) \quad V = -4.53 \text{ V non spontaneous} \)
   iii \( 3 \text{F}_2 + 2 \text{Al} \rightarrow 2 \text{AlF}^+ + 6 \text{F}^- \quad \xi^o = (2.87 + 1.66) \quad V = +4.53 \text{ V spontaneous} \)

 Standard Reduction Potentials

4. a) \( \begin{array}{c}
2 \text{Br}^- \rightarrow \text{Br}_2(g) + 2 e^- \\
3 e^- + \text{MnO}_4^- + 4\text{H}^+ \rightarrow \text{MnO}_2(s) + 2\text{H}_2\text{O}
\end{array} \)
   b) bromine half-reaction = oxidation (lose e⁻)
   manganese half-reaction = reduction (gain e⁻)
   c) \( 3(2 \text{Br}^- \rightarrow \text{Br}_2(g) + 2 e^-) \\
2(3 e^- + \text{MnO}_4^- + 4\text{H}^+ \rightarrow \text{MnO}_2(s) + 2\text{H}_2\text{O})
\)
   It is a six electron process.
   d) \( 6 \text{Br}^- + 2 \text{MnO}_4^- + 8\text{H}^+ \rightarrow 3 \text{Br}_2(g) + 2 \text{MnO}_2(s) + 4 \text{H}_2\text{O} \)
   e) Br⁻ is being oxidized (losing e⁻). It is the reducing agent.
   f) MnO₄⁻ is being reduced (gaining e⁻). It is the oxidizing agent.
   g) \( \text{MnO}_4^- \): each O has an oxidation state of 2⁻.
   \( \text{Mn}^{7+} + 4(2-) = 1- \)
   Mn = 7+
   \( \text{MnO}_2 \): each O has an oxidation state of 2⁻.
   \( \text{Mn}^{7+} + 2(2-) = 0 \)
   Mn = 4+
   so manganese goes from 7⁺ to 4⁺.

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