STRENGTH OF OXIDIZING AND REDUCING AGENTS

Strength of oxidizing and reducing agents involved in a half-reaction can be determined from the pertinent reduction potential, $E_{\text{half-reaction}}$.

The greater $E_{\text{half-reaction}}$, the reactant gets reduced easier and is a better oxidizing agent (cathode).

E.g., at standard conditions:

- $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)} \quad E^o = 0.34$  
- $\text{Al}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Al(s)} \quad E^o = -1.66$

$\text{Cu}^{2+}$ is a stronger oxidizing agent than $\text{Al}$. Therefore, $\text{Cu}^{2+}$ can oxidize $\text{Al}$ but not vice versa.

The lower $E_{\text{half-reaction}}$, the product gets oxidized easier and is a better reducing agent (anode).

E.g., at standard conditions:

- $\text{Al}$ is a stronger reducing agent than $\text{Cu}$. Therefore, $\text{Al}$ can reduce $\text{Cu}^{2+}$ but not vice versa.

Check:

If the assumption that $\text{Fe}^{2+}$ can oxidize $\text{Sn}$ under standard condition is correct, $E^o$ of the cell in which $\text{Sn}$ is the anode must be positive (reaction is spontaneous).

<table>
<thead>
<tr>
<th>Assumption:</th>
<th>$E^o$, V</th>
<th>Correct:</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cathode</td>
<td>Fe$^{2+}$(aq) + 2e$^-$ $\rightarrow$ Fe(s)</td>
<td>-0.44</td>
</tr>
<tr>
<td>Anode</td>
<td>Sn$^{2+}$(aq) + 2e$^-$ $\rightarrow$ Sn(s)</td>
<td>-0.14</td>
</tr>
<tr>
<td>$E^o_{\text{cell}}$ = -0.30 V</td>
<td>$E^o_{\text{cell}} = 0.30$ V</td>
<td></td>
</tr>
</tbody>
</table>

Therefore the assumption is not correct but $\text{Sn}^{2+}$ can oxidize $\text{Fe}$ ($E^o_{\text{cell}} = 0.30$ V) at standard conditions.