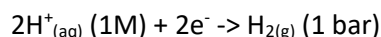


Electrochemistry Study Guide

1. Terms to know:

- Electrode Potential/Voltage: The amount of energy per unit of charge that a half-reaction either requires or releases (typically in units of Volts)
- Standard Hydrogen Electrode (SHE): A platinum electrode that provides a surface for the following reaction:



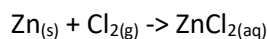
This electrode is arbitrarily assigned a potential of 0.

- Standard Electrode Potential/Voltage: The potential of an electrode in comparison with the SHE.
- Standard Cell potential/Voltage: The difference in standard potential between the cathode and the anode of an electrochemical cell:

$$E^\circ_{\text{cell}} = E^\circ_{\text{anode}} - E^\circ_{\text{cathode}}$$

2. How to calculate the Standard Cell Potential for a redox reaction:

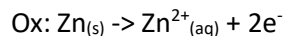
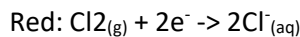
Take the following reaction:



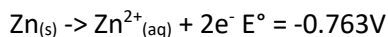
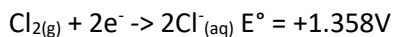
a. Assign oxidation numbers:

- $\text{Zn}_{(\text{s})} = 0$
- $\text{Cl}_{2(\text{g})} = 0$
- ZnCl_2 :
 - $\text{Cl} = -1$
 - $\text{Zn} = +2$

b. Write oxidation and reduction half-reactions:



c. Look up standard reduction potentials in a reference table:



d. Calculate Cell potential using the cell potential equation:

$$E^\circ_{\text{cell}} = E^\circ_{\text{anode}} - E^\circ_{\text{cathode}}$$

$$E^{\circ}_{\text{cell}} = 1.358 - (-0.763)$$

$$E^{\circ}_{\text{cell}} = 2.121\text{V}$$

Notice that the standard cell potentials in the reference table are all reduction potentials, even though you have one reduction half-reaction and one oxidation half-reaction. This is just convention. Generally, the more positive a reduction potential is, the more spontaneous the reduction reaction, so it makes sense that we get negative reduction potentials for your oxidation half-reactions.