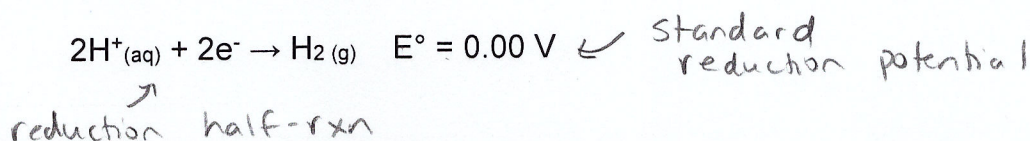


Cell Potential

- The **electric potential difference (E°)** is a measure of the tendency of electrons to flow from then anode to the cathode. The higher the electric potential difference, the greater the tendency for electrons to flow.
 - Electric potential difference is measured in units of volts (V) and is also referred to as **voltage** or just electric potential or potential difference
 - The degree symbol (E°) simply indicates the electric potential being measured at standard conditions
- Experiments were carried out to find the electric potential difference for different half-reactions. The potential difference for individual half-reactions written as a reduction half-reaction, are called **standard reduction potentials**
 - A list of standard reduction potentials are listed in pg. 7 of data book
 - Since a voltaic cell cannot work with only one half-reaction, the standard reduction potentials for individual half-reactions were recorded with hydrogen acting as the other half-reaction/ half-cell:



- The hydrogen half-cell has been assigned a voltage of 0.00V. Therefore the hydrogen half-reaction is considered the to be the standard/reference half-cell or half-reaction because all other standard reduction potentials are referenced to hydrogen
- The standard reduction potentials for individual half-reactions are useful for calculating the overall cell potential that will be produced by a voltaic cell

- The **standard cell potential** for any voltaic cell can be calculated using the following formula:

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

* never change these values!
do not change sign or multiply by a coefficient!

- To use this formula, it is crucial to first identify the oxidation half-reaction (occurs at anode) and the reduction half-reaction (occurs at cathode)

↙

- Strongest reducing agent
- oxidation half-rxn
- most negative reduction potential

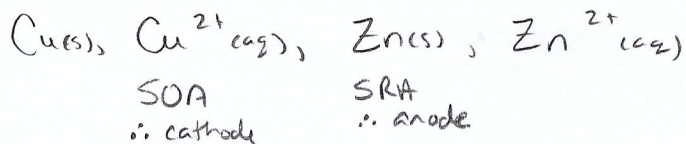
↘

- strongest oxidizing agent
- reduction half-rxn
- most positive reduction potential

Cu(s) - Zn(s) cell

EXAMPLES:

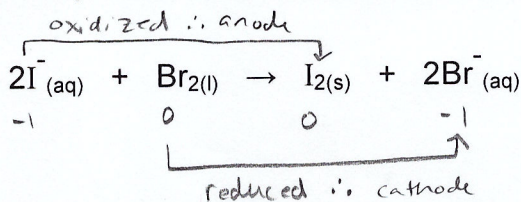
1. Calculate the standard electric potential for a voltaic cell that has Cu_(s) and Zn_(s) as electrodes.



$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 0.34\text{V} - (-0.76\text{V})$$

$$E^\circ_{\text{cell}} = 1.10\text{V}$$

2. Calculate the standard cell potential for the voltaic cell in which the following reaction occurs:

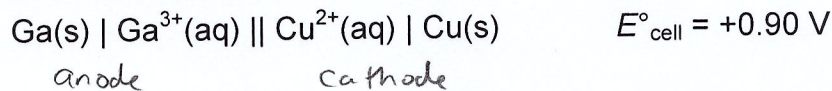


$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$E^\circ_{\text{cell}} = 1.07\text{V} - (0.54\text{V})$$

$$E^\circ_{\text{cell}} = 0.53\text{V}$$

3. Use the standard cell described below to determine the standard reduction potential of the gallium half-cell.



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

$$E^\circ_{\text{anode}} = E^\circ_{\text{cathode}} - E^\circ_{\text{cell}} = 0.34\text{V} - (0.90\text{V})$$

$$E^\circ_{\text{anode}} = -0.56\text{V}$$

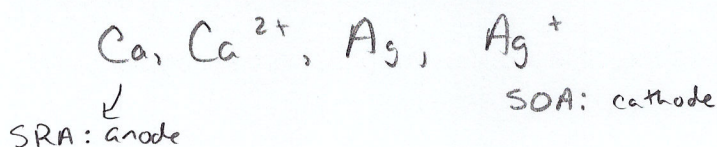
4. Assume that the reference half-cell is changed to cadmium. $\rightarrow E^\circ = -0.40V$
 a. What would be the standard reduction potential for calcium? \therefore set to 0.0V by adding 0.40V

$$-2.87V + 0.40V = \boxed{-2.47V}$$

- b. What would be the standard reduction potential for silver?

$$0.80V + 0.40V = \boxed{1.20V}$$

- c. Calculate the standard net cell potential for the calcium-silver cell if cadmium is the reference half-cell.



$$E_{cell}^\circ = E_{cathode}^\circ - E_{anode}^\circ = 1.20V - (-2.47V)$$

$$\boxed{E_{cell}^\circ = 3.67V}$$

- d. Calculate the standard net cell potential for the calcium-silver cell using the hydrogen half-cell as the reference.

$$E_{cell}^\circ = E_{cathode}^\circ - E_{anode}^\circ$$

$$E_{cell}^\circ = 0.80V - (-2.87V)$$

$$\boxed{E_{cell}^\circ = 3.67V}$$

* E_{cell}° do not change when reference half-cell changes!

Now try pg. 487 #1- 3, 4a,b & pg. 490 #2- 4, 6, 7 & Practice Problems

Cell Potential Practice Problem

1. Assume that the reference half-cell is changed to a standard mercury–mercury(II) half-cell.
 - (a) What would be the reduction potential of a standard chlorine half-cell? **[+0.51 V]**
 - (b) What would be the reduction potential of a standard nickel half-cell? **[- 1.11 V]**
 - (c) What would be the net cell potential of a standard chlorine–nickel cell if mercury was the half-cell? **[+1.62 V]**
 - (d) Calculate the net cell potential when hydrogen is the reference half-cell? Why is the answer to part c the same as the answer obtained using the standard hydrogen half-cell as the reference? **[All standard reduction potential are just shifting by a constant value depending on the assigned reference potential. However, the difference between standard reduction potentials will not change.]**

Section 13.1 Review Answers

Student Textbook page 490

1. It is important to keep each half-cell separate so that an instantaneous reaction does not occur, "short-circuiting" the voltmeter.
2. The standard half-cell potentials are measured using the hydrogen half-cell as a reference. In the equation $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$, the hydrogen half-cell is given a value of zero. Any measured voltage would therefore belong to the non-hydrogen half of the cell.
3. (a) $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$
 $= -1.66 \text{ V} - (-2.37 \text{ V}) = +0.71 \text{ V}$
(b) $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$
 $= +2.87 \text{ V} - (-2.93 \text{ V}) = +5.80 \text{ V}$
(c) $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$
 $= +1.23 \text{ V} - (0.80 \text{ V}) = +0.43 \text{ V}$
4. (a) $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$
 $= +0.34 \text{ V} - (-0.26 \text{ V}) = +0.60 \text{ V}$
(b) $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$
 $= +1.50 \text{ V} - (-0.40 \text{ V}) = +1.90 \text{ V}$
(c) $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$
 $= -0.04 \text{ V} - (-0.46 \text{ V}) = +0.50 \text{ V}$
5. No, a student using the Alberta Chemistry Data Booklet could not build a voltaic cell with a standard cell potential of 7.0 V. The strongest oxidizing agent $\text{F}_2(\text{g})$ and the strongest reducing agent $\text{Li}(\text{s})$ would only yield 5.91 V. By building a gold-aluminium cell in series with two cells, a total of 7.32 V could be achieved.
6. The cell potential describes the potential difference between two electrodes of a cell, or the amount of energy on a charge as it moves between two electrodes. The cell potential is dependent on both the anode and the cathode used. The standard reduction potential is a measure of the amount of energy for only the reduction half of cell. Since reduction cannot happen without oxidation, reduction potentials are measured against a standard reference, the hydrogen half-cell, which is set at a reduction potential of 0.00V.
7. $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$
 $= \text{"X"} \text{ V} - (-0.76 \text{ V}) = +1.75 \text{ V}$
 $X = 0.99 \text{ V}$