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:

$$2H^+_{(aq)} + 2e^- \rightarrow H_{2\,(g)} \quad E° = 0.00\ V$$

← standard reduction potential

↑ reduction half-rxn

  - 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:

$$E°_{cell} = E°_{cathode} - E°_{anode}$$

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

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

Anode:
- strongest reducing agent
- oxidation half-rxn
- most negative reduction potential

Cathode:
- 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.

$Cu_{(s)}$, $Cu^{2+}_{(aq)}$, $Zn_{(s)}$, $Zn^{2+}_{(aq)}$

SOA ∴ cathode   SRA ∴ anode

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

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

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

oxidized ∴ anode

$$2I^-_{(aq)} + Br_{2(l)} \rightarrow I_{2(s)} + 2Br^-_{(aq)}$$

-1   0   0   -1

reduced ∴ cathode

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

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

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

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

$Ga(s) | Ga^{3+}(aq) || Cu^{2+}(aq) | Cu(s)$ $E^\circ_{cell} = +0.90$ V

anode   cathode

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

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

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

$\nearrow E^\circ = -0.40\,V$ ∴ set to 0.0V by adding 0.40V

4. Assume that the reference half-cell is changed to cadmium.
   a. What would be the standard reduction potential for calcium?

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

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

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

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

$Ca,\ Ca^{2+},\ Ag,\ Ag^{+}$

SRA: anode (Ca)

SOA: cathode (Ag^{+})

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 1.20\,V - (-2.47\,V)$

$\boxed{E^\circ_{cell} = 3.67\,V}$

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

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

$E^\circ_{cell} = 0.80\,V - (-2.87\,V$

$\boxed{E^\circ_{cell} = 3.67\,V}$

* 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 $E^o_{cell} = E^o_{cathode} - E^o_{anode}$, 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)** $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
   $= -1.66\text{ V} - (-2.37\text{ V}) = +0.71\text{V}$

   (b) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
   $= +2.87\text{ V} - (-2.93\text{ V}) = +5.80\text{ V}$

   (c) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
   $= +1.23\text{ V} - (0.80\text{ V}) = +0.43\text{ V}$
4. **(a)** $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
   $= +0.34\text{ V} - (-0.26\text{ V}) = +0.60\text{V}$

   (b) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
   $= +1.50\text{ V} - (-0.40\text{ V}) = +1.90\text{ V}$

   (c) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
   $= -0.04\text{ V} - (-0.46\text{ V}) = +0.50\text{ V}$
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 $F_2(g)$ and the strongest reducing agent Li(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. $E^o_{cell} = E^o_{cathode} - E^o_{anode}$
   $= \text{"X" V} - (-0.76\text{ V}) = +1.75\text{ V}$

   X = 0.99 V