Hello everyone! This week’s resource will cover topics in Chapter 17 of the approved textbook, Chemistry: An Atoms First Approach.

Our Group Tutoring sessions will be every Thursday from 5:15-6:15PM at Sid Richardson Room 74. We will spend time solving problems and answering individual questions. I hope to see you there!

Keywords: Balancing Oxidation-Reduction Reactions, Galvanic Cells, Standard Reduction Potentials

**TOPIC OF THE WEEK**

**Balancing Oxidation-Reduction Reactions**

Any redox process involves the transfer of electrons. Oxidation and reduction both must take place – we cannot have one without the other. As a rule of thumb, remember the acronym “OIL RIG”: Oxidation Is Loss, Reduction Is Gain

The schematic below shows steps required to balance a redox equation in acidic (excluding steps colored green) or basic mediums.

Here is a sample calculation video on how to balance a redox equation: https://www.youtube.com/watch?v=fdbrhQAM9Gw&ab_channel=TheOrganicChemistryTutor

**PRACTICE PROBLEM 1:**

Use the half-reaction method to balance the equation below:
MnO$_2$(s) + HCl(aq) $\rightarrow$ MnCl$_2$(aq) + Cl$_2$(g) + H$_2$O(l)

**HIGHLIGHT #1: Galvanic Cells**

A galvanic (or voltaic) cell is a device that uses a spontaneous redox reaction to produce electric current. **Oxidation takes place at the anode** while **reduction takes place at the cathode**. Usually, a salt bridge is used to allow ions migrate to neutralize charge without mixing solutions. A typical galvanic cell is shown below:

![Galvanic Cell Diagram](https://www.chem.ubc.ca/~lind/113/chem113-S12/101-S12/GalvanicCell9.png)

**HIGHLIGHT #2: Standard Reduction Potentials**

This measures the tendency of reduction to occur at an electrode relative to the Standard Hydrogen Electrode (SHE). The SHE is used as a reference and assigned a value of 0 volt. **All half-reactions are written as reductions** (E$_o$).

A positive (+) E$_o$ value means the half-cell is easier to reduce than hydrogen while a negative (-) E$_o$ value means the half-cell is more difficult to reduce than hydrogen. An electrochemical cell can be described by a line notation that shows the anode on the left and the cathode on the right.

For example, **Cu(s)|Cu$^{2+}$(aq)||Ag$^+$(aq)|Ag(s)** is the line notation for the galvanic cell shown in the figure above. “|” stands for phase boundary and “||” represents salt bridge or porous barrier separating two half-cells. The **overall standard cell potential** (E$_{\text{cell}}$) is given as:

\[
E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}
\]

E$_{\text{cell}} > 0$ = spontaneous reaction; E$_{\text{cell}} < 0$ = non-spontaneous reaction

**PRACTICE PROBLEM 2:**

E$^o$ for Al$^{3+}$/Al is -1.676 V; E$^o$ for Fe$^{2+}$/Fe is -0.440 V. Determine the cell voltage and write the equation for the spontaneous reaction that occurs with these two half-cells.
THINGS YOU MAY STRUGGLE WITH

1. Not properly understanding how to find the oxidation state of an atom in a molecule to figure out which atom is being reduced or oxidized in a redox reaction. A good review from Chapter 6 is necessary in this case.
2. Some students easily forget to consider if a solution is acidic or basic before proceeding to balance a given redox reaction.
3. How to figure out the species being reduced, reducing agent, oxidizing agent and the species being oxidized from given cell potentials.
4. In calculating the overall cell potential, students sometimes subtract the cathode from the anode instead of doing the reverse.
5. Not remembering the proper equations needed to find different parameters in electrochemistry.
6. Forgetting to balance a redox equation before getting the value of n (number of moles of electrons) needed to substitute into the Nernst equation.

That’s all this week! Please reach out if you have any questions and don’t forget to visit the Tutoring Center website for further information at www.baylor.edu/tutoring.

PRACTICE PROBLEM ANSWERS

1. $\text{MnO}_2(\text{s}) + 4\text{HCl}(\text{aq}) \rightarrow \text{MnCl}_2(\text{aq}) + \text{Cl}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

2. $E^\circ_{\text{cell}} = +1.236 \text{ V}; 2\text{Al}(\text{s}) + 3\text{Fe}^{2+}(\text{aq}) \rightarrow 2\text{Al}^{3+}(\text{aq}) + 3\text{Fe}(\text{s})$