Unit 8, Lesson 01: Introduction to Electrochemistry and Oxidation/Reduction (Redox) Reactions

These youtube videos by Tyler DeWitt are excellent. Before reading the notes, please watch:
1. Introduction to Oxidation Reduction (Redox) Reactions
2. Oxidizing Agents and Reducing Agents

Notes (some review, some background and some material covered in the videos):
Recall that **thermochemistry** is the study of the changes in heat energy (Q) that accompany chemical or physical changes. In Unit 3, we focused on the inter-conversion of chemical potential energy (including enthalpy, H) and thermal kinetic energy transferred as heat (Q).

Electrochemistry is the study of the changes in electrical energy that accompany chemical reactions which involve electrons moving from one chemical species to another. The movement of electrons results in chemical potential energy (including enthalpy, H) being converted to electrical energy.

That is, electrical energy results when electrons are transferred from atoms or ions with a lower attraction for electrons (usually lower electronegativity) to atoms or ions with a higher attraction for electrons (usually higher EN).

Last year when we studied single displacement reactions, we created an Activity Series for Metals by placing a pure metal (made of neutral metal atoms) into a solution which contained the ions of a different metal.

- If a reaction occurred, we knew that the pure metal atoms were giving electrons to the metal ions in solution. That is, the metal ions in solution had a stronger attraction for electrons than the pure metal, so they pulled electrons away from the pure metal.
- If a reaction did not occur, we knew that the pure metal atoms were not giving electrons to the metal ions in solution. That is, the metal ions in solution had a weaker attraction for electrons than the pure metal, so they were unable to pull electrons away from the pure metal.

Example 1: Copper metal reacts with a solution of silver nitrate.
The overall reaction is: $Cu\,(s) + 2\,AgNO_3\,(aq) \rightarrow 2\,Ag(s) + Cu(NO_3)_2\,(aq)$

Write the **FULL** ionic equation by dissociating the ions of all soluble (aq) species. Remember to keep the equation balanced:

$$Cu\,(s) + 2\,Ag^{1+}\,(aq) + 2\,NO_3^{1-}\,(aq) \rightarrow 2\,Ag(s) + Cu^{2+}\,(aq) + 2\,NO_3^{1-}\,(aq)$$

Write the **NET** ionic equation by 'cancelling out' any ions that appear on both sides of the full ionic equation. These ions are called 'spectator ions' because they do not participate in the reaction.

$$Cu\,(s) + 2\,Ag^{1+}\,(aq) \rightarrow 2\,Ag(s) + Cu^{2+}\,(aq)$$

Now we can break the net ionic equation into two "half reactions":

$$Cu\,(s) \rightarrow Cu^{2+}\,(aq)$$
$$2\,Ag^{1+}\,(aq) \rightarrow 2\,Ag(s)$$

Balance the charge of each half reaction by adding electron(s) to the appropriate side of each equation:

$$Cu\,(s) \rightarrow Cu^{2+}\,(aq) + 2\,e^-$$
$$2\,Ag^{1+}\,(aq) + 2\,e^- \rightarrow 2\,Ag(s)$$

The half reactions allow us to see that one neutral copper atom is losing two electrons, creating one copper (II) ion, while two silver ions are each gaining one electron to create two neutral silver atoms. The number of electrons lost by the pure metal is equal to the number of electrons gained by the metal ions.

Example 2: Magnesium metal reacts in a solution of lead (II) nitrate.

Overall reaction: $Mg\ (s) + Pb(NO_3)_2\ (aq) \rightarrow Pb\ (s) + Mg(NO_3)_2\ (aq)$

Full Ionic: $Mg\ (s) + Pb^{2+}\ (aq) + 2\ NO_3^{1-}\ (aq) \rightarrow Pb(s) + Mg^{2+}\ (aq) + 2\ NO_3^{1-}\ (aq)$

Net Ionic: $Mg\ (s) + Pb^{2+}\ (aq) \rightarrow Pb(s) + Mg^{2+}\ (aq)$

Half reactions with electrons: $Mg\ (s) \rightarrow Mg^{2+}\ (aq) + 2\ e^-$

$Pb^{2+}\ (aq) + 2\ e^- \rightarrow Pb(s)$

Summary:

- Mg metal lost two electrons which were gained by the lead ions
- Mg loses electrons more easily than lead; that is, magnesium is more reactive than lead
- reaction is spontaneous at room temperature...do you remember ΔG? It is negative at SATP

Example 3: Copper metal does not react in a solution of zinc nitrate.

Overall reaction: $Cu\ (s) + Zn(NO_3)_2\ (aq) \rightarrow$ NR

Summary:

- Cu metal did not lose electrons to the zinc ions (Zn^{2+})
- Zn metal must lose electrons more easily than copper, so zinc is more reactive than copper
- this reaction is non-spontaneous at room temperature...ΔG is positive at SATP

Definitions:

When an atom or ion gains electrons, it is ‘reduced’ because its charge goes down (becomes more negative)

- the **G**ain of **E**lectrons is called '**R**eduction' (GER)
- the atom or ion that is reduced is called the 'oxidizing agent' because it takes electrons from another atom or ion and causes that atom or ion to be oxidized

When an atom or ion loses electrons, it is 'oxidized'

- the **L**oss of **E**lectrons is called **O**xidation (LEO)
- the atom or ion that is oxidized is called the 'reducing agent' because it gives electrons to another atom or ion and causes that atom or ion to be reduced

Oxidation and reduction occur simultaneously, so they are called 'redox' (**red**uction-**ox**idation) reactions. The electrons that are lost by one atom or ion must be gained by another atom or ion.

We can predict which atom or ion will be oxidized or reduced using an 'activity series' which arranges the elements from the most reactive at the top (they lose electrons most easily so they will be oxidized) to the least reactive at the bottom (they gain electrons most easily so they will be reduced).

A neutral metal atom will lose electrons to (be oxidized by) any metal ion below it on the activity series.

Unit 8, Lesson 01: Activity Series for Metals

Based on experimental evidence, metals can be arranged in an activity series, from most reactive (lose electrons most easily) to least reactive (gain electrons most easily).

A neutral metal ***atom*** will be oxidized by (lose electrons to) any metal ***ion*** below it on the activity series.

If a neutral metal ***atom*** is below the metal ***ion*** on the activity series, no reaction will occur.

Use the Activity Series to predict if a reaction will occur between the following metal atoms and ions. If a reaction will occur, complete the equations. If no reaction will occur, write "NR" (no reaction):

Most reactive metal (loses electrons to any ion below it and is oxidized)

Lease reactive ion (gains electrons from any metal above it and is reduced)

Activity Series

lithium
potassium
barium
calcium
sodium
magnesium
aluminum
zinc
chromium
iron
cadmium
cobalt
nickel
tin
lead
copper
mercury
silver
platinum
gold

Practice Questions:

1. Calcium metal with nickel (II) nitrate solution:

a) Balanced chemical equation:

____

b) Full ionic equation:

____

c) Net ionic equation: ____

d) Oxidation half reaction: ____

e) Reduction half reaction: ____

f) Oxidizing agent: ____

g) Reducing agent: ____

2. Tin metal placed in a solution of zinc sulfate:

a) Balanced chemical equation: ____

b) Full ionic equation: ____

c) Net ionic equation: ____

d) Oxidation half reaction: ____

e) Reduction half reaction: ____

f) Oxidizing agent: ____

g) Reducing agent: ____

3. Lead metal placed in a solution of silver nitrate:

a) Balanced chemical equation: ____

b) Full ionic equation: ____

c) Net ionic equation: ____

d) Oxidation half reaction: ____

e) Reduction half reaction: ____

f) Oxidizing agent: ____

g) Reducing agent: ____