

Week 5 • Oxidation-Reduction

TEACHING IDEAS AND RESOURCES

Overview:

As with Equilibrium, Oxidation-Reduction shows up at various times during the course. We first encounter it when we notice that elements can either be neutral or ions. The change from one to the other is oxidation or reduction. We end with exploiting the gain and loss of electrons to create electrochemical cells. Here is how to approach **Week 5**.

- Read through these **Teaching Ideas** about this topic.
- Read the **Week 5 Labs and Demonstrations** handout that discusses ways to illustrate redox ideas.
- Review the **various handouts** and **practice worksheets** that I use to teach this subject.
- Practice doing these problems. Ask questions and make comments in the **Discussion Board** about the Week 5 Resources and Week 5 AP Questions.

Basic Oxidation-Reduction Ideas:

Some Reactions Involve the Transfer of Electrons:

As soon as students see an example they can understand, I introduce the idea that an atom can gain or lose electrons. The easiest example is a reaction that changes copper metal (like a penny) into copper ion (a peacock blue solution). The copper has **lost electrons**. Something else in the reaction must have **gained electrons** because atoms do not just spray electrons off into the air. I tell student that this kind of reaction is so important that we have special names for them: **Oxidation** (losing electrons) and **Reduction** (gaining electrons). Burning a piece of magnesium is another easy demonstration that makes the same point.

Oxidation and Reduction are Linked... Redox:

If one chemical **gains** electrons (reduction), obviously something has to **lose** electrons (oxidation). This is an important consideration in Predicting Reactions. If a student has shown one type of change, they must include the other change as well.

Oxidizing Agents and Reducing Agents:

The chemical that is *oxidized* is called the “**reducing agent**” because it is the agent by which the other chemical is reduced. Since the reducing agent loses electrons, it *forces* or *allows* the other chemical to gain electrons. Similarly, the chemical that is *reduced* is the “**oxidizing agent**” which *forces* or *allows* the other chemical to be oxidized. In a reaction you can say this chemical is oxidized and that one is reduced or you can say that this chemical is oxidized and that chemical is the oxidizing agent. You can say that chemical is reduced and this chemical is the reducing agent. We keep track of certain chemicals that are good oxidizing or reducing agents to use in labs. See the “Stuff I Am Supposed To Know for the AP Test...” handout.

All Atoms Have Charges or Apparent Charges (Oxidation Numbers):

Gaining and losing electrons is pretty obvious when a neutral element changes to an ion or vice versa. More generally, however, even non-ionic species can transfer electrons. We assign numbers to the atoms called oxidation numbers. For example, when CH_4 burns, it changes from CH_4 to CO_2 . The oxidation number of C in CH_4 is -4 while the oxidation number in CO_2 is +4. The carbon atom is not an ion, but we say it gains 8 electrons during this change. At the same time the oxygen in O_2 changes from an oxidation number of 0 to 2-. Preview: The oxidation number is the “apparent charge”... it is the charge an atom would have if we assign all of its shared electrons to the more electronegative atom. We’ll look at this idea again later in these notes.

Electron-transfer reactions can be spontaneous or non-spontaneous:

When zinc metal contacts copper(II) sulfate, a redox reaction occurs. This is spontaneous. Zinc metal, zinc ions, copper metal, and copper ions can be put together to form an electrochemical cell. This is also a

spontaneous change. However, we can use electricity to coat a piece of copper with zinc by changing zinc ions into zinc metal even though the reverse reaction is the spontaneous reaction. This is electroplating or more generally electrolysis. All three of these cases show up in the course and are all redox. Corrosion and combustion are two other cases of spontaneous redox reactions.

All Reactions Are Either Oxidation-Reduction or Acid-Base:

Every reaction either involves a change in oxidation numbers (redox) or it does not. If oxidation numbers do not change, it is an acid-base reaction although you may need to use the Lewis acid-base definition to find the acid and the base.

Assigning Oxidation Numbers:

I am not trying to write a textbook here so your textbook will help you with assigning oxidation numbers. However, one idea that really helped me understand the rules for assigning oxidation numbers is the idea that the oxidation number for non-ionic species is the “**apparent charge**”. It is the charge an atom *would have* if you assigned all of its shared electrons (covalent bonds) to the *more electronegative atom*.

Consider the HCl molecule to the right. While H^+ ions and Cl^- ions can exist, the molecule, HCl, is not ionic. (Like all acids, HCl only forms ions in aqueous solution.) We assign oxidation numbers to the two atoms. Chlorine has a larger electronegativity (3.0) than hydrogen (2.1). Therefore, we assign the shared pair of electrons to chlorine. Chlorine has 7 valence electrons, but in HCl we have assigned it 8, so its apparent charge is -1. Hydrogen has 1 valence electron, but in HCl we have assigned it 0, so its apparent charge is +1. Easy.



This probably doesn't seem to make anything easier. However, it helps when we look at the rules for assigning oxidation numbers.

Rule	Reason
F, in its compounds, is always -1.	F is the most electronegative atom and so we will always assign shared electrons to it. It will always have an octet and an apparent charge of -1.
O, in its compounds, is usually -2.	O is the second most electronegative atom and usually gets the shared electrons assigned to it. However, what about OF_2 ? Since F is more electronegative, F gets the shared electrons assigned to it and so O has only 4 electrons and an apparent charge of +2. What about H_2O_2 ? When you draw the Lewis dot structure, you have the problem of assigning the pair of electrons shared by the two oxygen atoms. We give each O atom one of the shared electrons so each O ends up with 7 electrons. The apparent charge on O is -1.
Family I, in its compounds, is always +1. Family II, in its compounds, is always +2. Al, in its compounds, is always +3.	The alkali metals, alkaline earth metals, and aluminum don't often form covalent bonds. When they do, however, they always have the lower electronegativity value and end up with none of the shared electrons.
H, in its compounds, is usually +1.	Covalent bonds generally form between nonmetals. H usually has a lower electronegativity (2.1) than the atoms with which it shares electrons so it has its one electron assigned away and has an apparent charge of +1. However, in compounds such as LiH, MgH_2 , or AlH_3 , it is the atom with the larger electronegativity and so gets an apparent charge of -1.

Atoms such as chlorine can have several different oxidation states (consider Cl in ClO_4^- , ClO_3^- , ClO_2^- , ClO^- , and Cl^-) depending on the compound. Its largest oxidation state of any atom, however, will always be its group number because of the situation where you assign away all of its valence electrons. If you come across a tricky molecule that the above rules don't cover, just look up the electronegativity values. Here is an interesting movie: http://cwx.prenhall.com/petrucci/medialib/media_portfolio/text_images/045_CommonOxiSta.MOV

Three Classes of Redox Reactions:

There are going to be three situations where we see oxidation-reduction reactions.

Spontaneous Redox	Electrochemical Cells	Electrolysis/Electroplating
When substances come in contact, one may gain electrons and one may lose electrons. $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$ $3\text{Cu}^{2+} + 2\text{Al} \rightarrow 3\text{Cu} + 2\text{Al}^{3+}$ combustion and corrosion are other good examples	Any spontaneous redox may be set up in such a way that the flow of electrons occurs in an external circuit and causes electricity. Batteries and fuels cells are examples.	Electricity can <i>cause</i> a chemical change. $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ silver plating, gold plating, or chrome plating are examples.

Introducing Oxidation-Reduction:

I officially introduce redox with the copper(II) chloride and aluminum foil labette. (Note: you cannot use copper(II) sulfate unless you add some sodium chloride. I think the chloride has something to do with getting through the tenacious oxide coating on the aluminum metal.) We see the brown (looks like rust) solid forming on the aluminum metal surface. The aluminum metal disintegrates as the blue color of the solution disappears. There is enough visual evidence here to help students build the model that the copper ion is becoming copper metal (gaining electrons) and that the aluminum metal is becoming aluminum ions (losing electrons). We name these two processes.

Predicting Reactions:

When we are predicting reactions, if we see a **neutral element** as a reactant or product, we know the reaction must be oxidation-reduction. We look for the oxidation and the reduction. An unfamiliar example is lithium metal reacting in air (forming a black coating). Neutral Li *must* be forming the Li^+ ion because it is the only other form of Li that we know. This is oxidation, so there must also be a reduction. Air contains O_2 and N_2 . It is reasonable to think that O in O_2 forms O^{2-} (with which we are familiar) or N in N_2 forms ??? Here is where the periodic table comes in handy. Even though N^{3-} is not a common ion, we can see that it forms just like sulfide, oxide, chloride, fluoride, as these atoms fill their electron shell. We should tuck away in our “bag of tricks” the fact that Li reacts with nitrogen gas. The product is $\text{Li}_3\text{N}(\text{s})$. [Note: Students often mistakenly show the products as separate ions because alkali metals compounds are always soluble, but again, we need to remind students to *check to see if there is a solution* before we think about solubility.] $6\text{Li}(\text{s}) + \text{N}_2(\text{g}) \rightarrow 2\text{Li}_3\text{N}(\text{s})$

Watch for **good oxidizers (oxidizing agents)** and **good reducers (reducing agents)** in the predicting reactions section of the exam. When you see “an acidified solution of potassium permanganate” students need to recognize that this is a strong oxidizing agent... and that one of the reactants is going to get oxidized. Students (and teachers) need to *memorize* what acidified potassium permanganate turns into.

The unbalanced half-reaction is: $\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + \text{H}_2\text{O}$ (notice that K^+ is a spectator ion.)

There are several good oxidizers and good reducers (see the “Stuff I Should Know...” handout.) Notice that the Mn in permanganate is in a highly oxidized state. This is one reason why it is a good oxidizer.... it is going to change to some lower oxidation state (get reduced) and cause another chemical to be oxidized. Also notice that the H^+ ions play an important role in forming H_2O . The Mn needs to get away from the O’s so it can be reduced. In a neutral or basic solution, the products are different.

Balancing Redox Equations:

I leave this to you and your textbook to figure out. The only trick I can offer is to always balance equations as though they are in acidic solutions. If the solution is actually basic (alkaline), do one final step. Add OH^- ions to both sides of the equation... as many as you have H^+ ions. On one side you will form water molecules. Cancel out any H_2O molecules common to both sides and the equation is now balanced for a basic solution. I am going to have my students practice this skill a little more in the future because the predicting reactions section will likely contain a redox reaction every year.

Reduction Potentials:

Key Idea: Each metal has a different “pull” for electrons. We can write this as a **reduction** reaction.

For example: $\text{Cu}^{2+} + 2\text{e}^{-} \rightarrow \text{Cu}^{\circ}$ or $\text{Na}^{+} + \text{e}^{-} \rightarrow \text{Na}^{\circ}$.

Even though these two reactions *look* very similar, the *extent to which they occur* is very different. Copper(II) ions grab onto electrons very well. This is why we can make money, pipes, and gutters out of copper metal. Sodium metal is another story completely. This difference in tendency is quantified by E° , the reduction potential. Books do a great job of explaining E° , so I won't do it here, but students need to understand that since each metal has a different pull for electrons, a transfer of electrons can occur when different metals come in contact in a moist environment. The strange feeling that people with metal fillings have when they chew on aluminum foil is due to this difference in pull for electrons and the transfer of electrons that results.

Since a good chart of Reduction Potentials is supplied on the AP chemistry exam, it is good to help students *understand and use this chart* to predict products and calculate voltages. Some of the handouts and labettes practice this idea. One of the labettes also helps students see where these voltages come from. **Be warned**, however, that you will never measure the values they have on those charts. In the Activity of Metals labette, at the end, you will see that the voltages are in the correct order, but they will not match the published values. It is more of an art than science to get those voltages. Besides the fact that the labette is not run under standard conditions (0.100 M solutions instead of 1.00 M solutions, etc.) to measure the published values the electrodes are sometimes coated in mercury and other tricks of the trade.

Electrolysis:

Electricity can be used to cause a chemical change. This is the idea behind electrolysis and electroplating. I do an activity (see handouts) in which students do and analyze a few examples of electrolysis. It is important for students to realize that during electrolysis, either the dissolved ion will change or water will change. I begin the labette by finding the reactions of water as it is oxidized and reduced.

Reduction: (think of H_2O as H^{+} and OH^{-}) $2\text{H}_2\text{O} + 2\text{e}^{-} \rightarrow \text{H}_2 + 2\text{OH}^{-}$ solution becomes basic

Oxidation: (think of H_2O as 2H^{+} and O^{2-}) $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^{+} + 4\text{e}^{-}$ solution becomes acidic

If there are ions in solution, they might be oxidized or reduced instead of the water. You can tell by using the reduction potential chart... for the positive ion, the greater the E° value, the more likely to be reduced. For negative ions, the smaller the E° (more negative) the more likely to be oxidized. I alert my students to think... if the positive ion is Na^{+} or K^{+} , there is no way that will be reduced to reactive sodium or potassium metal under water. In the same way, if the negative ion is SO_4^{2-} , sulfur is in its highest oxidation state and cannot be further oxidized.



Electrolysis Calculations:

Now we are beginning to mix a little physics (electricity) with chemistry. We need to have a little understanding of physics units. It is common to use a water analogy here.

Physics Unit	Meaning	Water Analogy
Coulombs	amount of charge	amount of water
Amperes	flow of electricity	flow of water
Voltage	electrical potential	pressure of water

A coulomb is the amount of charge that will flow through a 1 amp circuit in 1 second. $1\text{ C} = 1\text{ amp}\cdot\text{sec}$
The Faraday constant (F) is the connection between physics and chemistry. 1 mole of e^- 's = 96,500 Coulombs
These two ideas can be turned into conversion factors. Say you wanted to know how many grams of Cu^0 is plated out when a 0.250 amp circuit is run for 15.0 minutes through a solution of copper(II) chloride. Setup:

$$.250\text{ amp} \times 15.0\text{ min} \times \frac{60\text{ sec}}{1\text{ min}} \times \frac{1\text{ C}}{1\text{ amp}\cdot\text{sec}} \times \frac{1\text{ mole } e^-}{96,500\text{ C}} \times \frac{1\text{ mole Cu}}{2\text{ mole } e^-} \times \frac{63.5\text{ g Cu}}{1\text{ mole Cu}} =$$

So, the important points are that the equation begins with two values (current and time), $1\text{ C} = 1\text{ amp}\cdot\text{sec}$ is an important conversion, and the Faraday constant is another important conversion. The Faraday constant is given on the AP exam equation table.

Spontaneous Oxidation-Reduction Reactions:

When two chemicals are placed together, they often can go through a spontaneous reaction. We have seen this in the **Intro to Redox** reaction in which aluminum metal reacts with copper(II) chloride. This occurs in the first part of the **Activity of Metals** lab in which four metals and four solutions of their ions react to show which metal is most active. The **Growing Crystals in Gels** lab also illustrates this kind of spontaneous redox reaction.

I also like to preview a little oxidation and reduction with an ornament lab that we do just before the Winter Break. I scanned in my handout (<http://www.chemmybear.com/blackboard/ornament.jpg>) but this needs to be re-typed. It is a cute little lab and the chemistry is good. We have changed the procedure now that there is that great blue painters tape. Old brown masking tape had to be allowed to dry overnight or else you could not remove it without leaving behind glue. The blue tape can be removed instantly.

These are all reactions in which the electron transfer occurs right away. You can set up the reactions, however, so that the transfer of electrons occurs through an outside circuit. These are electrochemical cells. Several cells can be put together in series to make a battery.

Electrochemical Cells:

I will not say much about this area. You and your students simply need to slog through this material until it makes sense. I love the CHEMStudy video: [Electrochemical Cells](#), but it is difficult stuff and needs to be viewed more than once. I'm not sure the students can stay awake through it.

The **online tutorial** I listed is also an excellent presentation of the ideas. I have a worksheet I use to practice this concept with my students and the second part of the Activity of Metals lab is a chance for students to build small electrochemical cells.



I do like to begin this section by showing my students my two-potato clock. When I first saw this in the store, I was put off because I simply thought the concept was that the potato was a conductor and there was a hidden battery. I was wrong. Into each potato you stick two different pieces of metal (zinc and copper) and the potato plays the role of the salt bridge. There are two "potato cells" in series to provide the electricity needed for the digital clock. It lasts as long as there is moisture in the two potatoes. I've seen various forms of this clock.

Just before Winter Break I also demonstrate the silver Christmas Tree. I do mine in a Petri dish with 0.100 M AgNO_3 . I write the reaction on the board and discuss which change is oxidation and which is reduction... which atom is gaining electrons and which is losing electrons.

At this point I sing my carol: *Oh Chemist's Tree*

Oh Chemist's Tree, Oh Chemist's Tree,
Thou copper limbs are silvery.

When silver gains electrons free
No longer ions will they be.

The copper's metal, sweet and nice,
But silver steals it's electrons... twice.

Oh Chemist's Tree, Oh Chemist's Tree,
Form copper ion, oxidatively.



My version:

<http://www.chemmybear.com/groves/tree.html>