CHE 1302

Basic Principles of Modern Chemistry II

Week 14

Hello and welcome to the weekly resources for Chemistry 1302! This resource covers topics typically taught by professors during the 14th week of classes.

On our website, https://baylor.edu/tutoring, you'll find the following links:

"Online Study Guide Resources" – If you don't see the topics you're learning right now, click here to find the weekly resources for the rest of the semester!

"How to Participate in Group Tutoring" - See if there is a Chemistry 1302 group tutoring session being hosted this semester – these are weekly question/answer sessions taught by our master tutors!

You can also view tutoring times for your course or schedule a private 30-minute appointment! Check out the website to learn more. You can also give us a call at (254)710-4135, or drop in. Our hours are Monday-Thursday 9 am – 8 pm on class days.

KEY WORDS: Oxidation, Reduction, Balancing Redox Reactions, Galvanic Cells

TOPIC OF THE WEEK: Balancing Oxidation-Reaction Reactions

This last bit of Chem II is part of electrochemistry: it focuses on the movement of electrons. In an oxidation-reduction reaction, one item gets oxidized and the other reduced. Let's start by looking at the definitions of oxidation and reduction.

Oxidation is the loss of electrons.

Reduction is the gain of electrons. These electrons have to come from somewhere, right? In an oxidation-reduction reaction, they come from something that has been oxidized.

An oxidation-reduction reaction, then, can be said to be made of two "half-reactions:"

The first is an oxidation. Something loses an electron or electrons, resulting in a positive charge or charges for that item.

$$A \to A^+ + e^-$$

The second is a reduction . Something else gains the electron(s) lost by *A.*

$$B^+ + e^- \to B$$

The sum of these two reactions (the electrons cancel out, because the same electrons are "produced" on one side of the equation and absorbed on the other.

$$A + B^+ \rightarrow A^+ + B$$

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

Note: oxidizing agent – does the oxidizing, gets reduced; reducing agent – does the reducing, gets oxidized

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Highlight 1: Oxidation Number Rules

Oxidation number is a measure of how oxidized an atom is.

Atoms are oxidized if (1) they lose electrons completely or (2) they are covalently bonded to atoms more electronegative than they are (that is, the electronegative atoms hog the electrons in the covalently bond, leaving the other electron-deficient).

These rules can be used to determine oxidation number—annotated from https://www.thoughtco.com/rules-for-assigning-oxidation-numbers-607567#:~:text=Rules%20for%20Assigning%20Oxidation%20Numbers%201%20The%20convention,in%20compounds%20is%20usually%20-2.%20...%20More%20items

1. The oxidation number of a free element (by itself & not an ion, or just bonded to other atoms of the same element) is **zero**.
    a. No electrons are being pulled away from the atom. If it is bonded to another atom of the same element, this other atom has the exact same electronegativity.
2. The oxidation number of a monoatomic ion equals the **charge** of the ion.
    a. An ion has lost/gained electrons.
3. The usual oxidation number of a hydrogen is **+1**. If it is bonded to compounds less electronegative than hydrogen, its oxidation number will be **-1**.
    a. Hydrogen wants to lose or gain one electron so it will not have a half-filled valence electron shell
4. The oxidation number of a Group IA element in a compound is **+1**. IIA is **+2**. Etc.
    a. These compounds want to lose 1 or 2 electrons (respectively) to have full valence shells
5. The oxidation number of a Group VIIA element is **-1**, **except** when that element is combined with one having a higher electronegativity.
    a. In summary, to figure out an atom's oxidation state, count the number of more electronegative atoms that it is bound to.
6. The sum of the oxidation numbers of all the atoms in a neutral compound is **0**.
7. The sum of the oxidation numbers in a polyatomic ion is equal to the **charge** of the ion.

Highlight 2: Balancing Redox Reactions

Reactions' being balanced have two aspects:

Mass: We've covered this topic before now. Both sides of an equation must have the same number of atoms of elements.

Charge: The change in oxidation number from the oxidation half-reaction must equal the change in oxidation number from the reduction half-reaction.

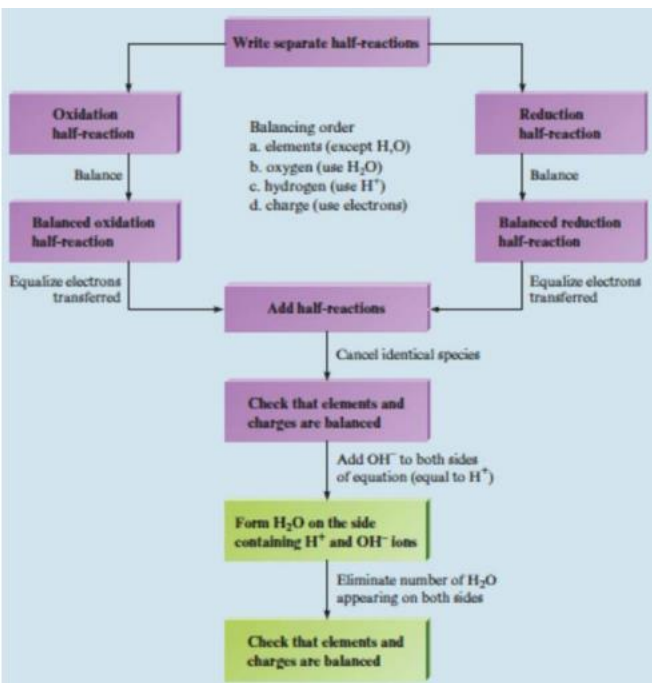
How to balance a half reaction in acidic conditions:

1. Balance by mass. If there are, for example, **oxygens** on one side of the equation, put **water** (assume the oxygen source is water) on the other to balance. The hydrogens from the water now must be balanced; put **protons** (H^+) on the other side of the equation.
2. Add electrons as necessary to balance charge.
3. Check: Number of electrons should be $O.N._{initial} - O.N._{final}$ where O.N.=oxidation number

How to balance a half reaction under basic conditions:

1. Perform steps listed above
2. Add OH^- to the balanced half reaction assuming acid conditions
    a. Necessary to get rid of the H^+ because it is not reasonable to say that there is a lot of it in a basic solution – review pH

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

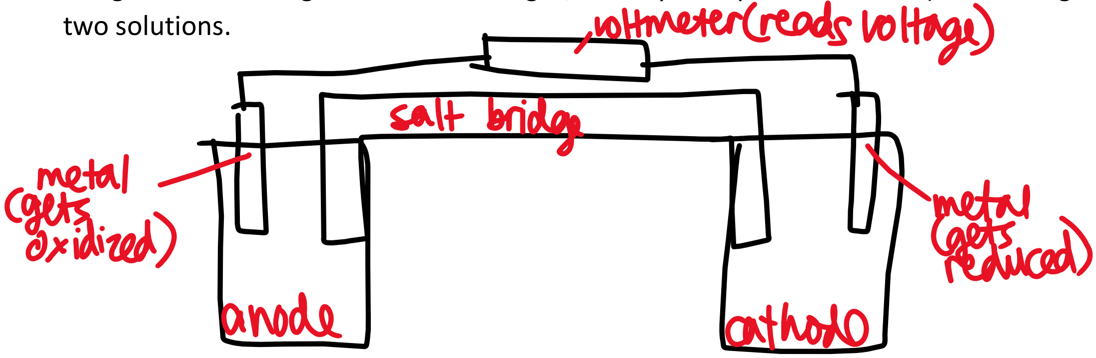


Here is a sample calculation:

https://www.youtube.com/watch?v=fdbrhQAM9Gw&ab_channel=TheOrganicChemistryTutor

Highlight 3: Electrochemical Cells

Redox reactions take place in electrochemical cells. A cell consists of two solutions connected by a salt bridge. The salt bridge has a lot of charges; it easily transports electrons (but nothing else) between the two solutions.



In one solution, substance is being oxidized, generating electrons. This side of the cell is called the "anode;" negative charge originates here. The other side of the cell receives the electrons, uses them for reduction reactions, and is called the "cathode."

Electrons flow from the anode to the cathode; the **current** is measured in Amperes, or Coulombs (a unit of charge) per second. 1 A = 1 C/s

Galvanic cells: electrochemical cells in which the redox reactions drive the current.

Electrolytic cells: electrochemical cells in which the current drives redox reactions.

See Week 15 for more info on voltage and measuring cell potential.

Check Your Learning

1. Does copper get oxidized or reduced if it goes from $CuSO_4$(aq) to Cu(s)?
2. Use the half-reaction method to balance the equation below: MnO_2(s) + HCl(aq) → $MnCl_2$(aq) + Cl_2(g) + H_2O(l) .
3. In an electrochemical cell, electrons flow from the (cathode/anode) to the (cathode/anode).

Things You May Struggle With

1. Understand oxidation number to make sure that you correctly identify the oxidizing and reducing agents in a redox reaction.
2. Don't forget to consider whether a solution is acidic or basic before proceeding to balance a given redox reaction.
3. When balancing acidic half reactions, add water first, to balance out the oxygens. Add protons after this, and electrons last! Then, if the solution is basic, add hydroxide.

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. Answers to the Check Your Learning section are below.

1. Reduced
2. $MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + Cl_2(g) + 2H_2O(l)$
3. Anode, cathode