Oxidation-Reduction Chemistry ("Redox" Chemistry)

I. What is "redox" chemistry?

a. A redox reaction takes place when electrons are transferred from one atom or ion to another atom or ion.

b. "Redox" is an abbreviation for "oxidation-reduction" chemistry. I guess it's abbreviated as "redox' because it's faster than saying "oxidation-reduction" and easier than saying "ox-red". ☺

c. Redox chemistry is not new to you. We have seen redox reactions all year long. For instance, all single replacement reactions are redox reactions. However, we will be talking about these reactions in a new way in this chapter.

d. Here is an example of a redox reaction that you have done in the laboratory this year:

\+

→

Solid zinc metal, which has a silvery appearance

Copper (II) nitrate solution. Remember that copper *solutions* have a blue appearance.

Solid copper metal is produced. It is brownish orange, just like a penny. It is forming on the outside of the zinc metal at the bottom of the test tube. As the amount of solid copper that is formed increases, the amount of copper metal ions that are dissolved in the solution decreases. Eventually, the solution will become less blue, because the blue color of the solution is due to the presence of dissolved copper. The solid zinc metal, on the other hand, disappears as it dissolves.

Here is the chemical equation:

$$\mathbf{Zn\ (s) + Cu(NO_3)_2\ (aq) \rightarrow Cu\ (s) + Zn(NO_3)_2\ (aq)}$$

Here is the net ionic equation:

$$\mathbf{Zn + Cu^{2+} \rightarrow Cu + Zn^{2+}}$$

Zn Before (reactant)

Cu^{2+} Before (reactant)

Cu + Zn^{2+} Afterward (products)

Zinc loses electrons, and its charge goes up from "0" (before) to "+2" (afterward)

The copper (II) ion gains electrons and its charge goes down from "+2" (before) to "0" (afterward)

↑ ↑ (Look at their charges afterward!)

II. Who cares about redox chemistry?

   a. Redox chemistry occurs when electrons travel from one place to another in a chemical reaction. Why does this occur? Well, that's like figuring out why some things fall down to the ground while other things don't. Imagine a large object in the sky, like a parachutist dropped from an airplane. The person will fall down to the ground because that is a spontaneous process. On the other hand, a jet airplane weighing several tons can leave the ground and go into the sky when its engines are turned on because *that* is a spontaneous process.
   b. Water falls down a waterfall spontaneously, and the motion of the water can be used to turn a turbine, which is a device that turns the motion of the water into electricity. Similarly, electrons move from one chemical to another in a redox reaction; we can make the electrons do work for us as they move from one place to another. A hydroelectric dam works by forcing the falling water to move through a turbine; a battery in your iPod works by forcing the electrons in a chemical reaction to move through wires as one chemical loses electrons to another chemical that "wants" the electrons more. (How do we know which way the electrons will flow – which chemical will lose electrons and which one will want to gain the electrons? We will answer this question later in the section on **electrochemistry**.)
   c. Therefore, one of the reasons that we care about redox chemistry is that understanding redox chemistry is the key to understanding electrochemistry, which is the conversion of chemical energy into electrical energy. Electrical energy is the energy due to the flow of electrons from one place to another. If you can understand that moving water (over a hydroelectric dam) can produce useful power for us, then you should realize that electrochemistry is just making use of moving electrons instead of moving water. In fact, the speed at which the electrons flow through a wire is described as **current**, just like a river that has lots of water moving through it is said to have a very high current.

III. Assigning oxidation numbers

   a. In order to identify a reaction as being a redox reaction, you have to see if anything is oxidized or reduced. The easiest way to do this is to look at the changes in the **oxidation numbers** of the elements in the reactions.
      i. The change in the oxidation number lets us know which substance is losing electrons, and which substance is gaining the electrons.
      ii. The oxidation number is basically the same thing as the "charge" of the element. (In case you care, it's not quite the same thing. Oxidation numbers can have non-whole number values. Oxidation numbers can be assigned to the elements

in a molecular compound, which by definition has no ions and therefore has no charges. But hey – don't worry about this until AP Chemistry! ☺)

b. Your textbook has a list of rules about how to assign oxidation numbers. Here is a quick summary:

   i. Any element that is not in a compound (i.e., it is "all by itself") has an oxidation number of 0. Examples:

Element	**Oxidation number**
Cl_2	0
O_2 (the oxygen that you need to live)	0
O_3 (this is ozone)	0
Al	0
Ca	0
Fe	0

   ii. Any monatomic ion has an oxidation number that is equal to its charge. Examples:

Ion	**Oxidation number**
Cl^{-1}	-1
O^{2-}	-2
Al^{3+}	+3
Ca^{2+}	+2
Fe^{2+}	+2
Fe^{3+}	+3

   iii. Oxygen is *usually* -2
   iv. Fluorine is -1
   v. H is *usually* +1
   vi. All oxidation numbers in a compound have to add up to 0 Examples:

Compound	**Oxidation numbers**
HCl	H = +1, Cl = -1
H_2O	H = +1, O = -2
Al_2O_3	Al =+3, O =-2
CaF_2	Ca =+2, F = -1
FeO	Fe =+2, O = -2
Fe_2O_3	Fe =+3, O = -2
CH_4	C = -4, H = +1

CO_2	C = +4, O = -2

vii. All oxidation numbers in a polyatomic ion have to add up to the charge on the polyatomic ion. Examples:

Polyatomic Ion	**Oxidation numbers**
CO_3^{2-}	C = +2, O = -2
SO_4^{2-}	S = +6, O = -2
SO_3^{2-}	S =+4, O =-2
O_2^{2-}	O = -1
Hg_2^{2+}	Hg = +1
ClO_4^{-1}	O = -2, Cl = +7

IV. Identifying redox reactions

a. The easiest way to identify redox reactions is to see how the oxidation numbers change as the reactants turn into products.

b. **Oxidation** occurs when an element's oxidation number <u>goes up</u>.

c. **Reduction** occurs when an element's oxidation number <u>goes down</u>.

d. Example: Is the reaction below a redox reaction? If so, identify the substances that are oxidized and reduced.

$Zn + S \rightarrow ZnS$

Solution:

$Zn^{(0)} + S^{(0)} \rightarrow Zn^{(2+)}S^{(2-)}$

ox # going up (Zn) ; ox # going down (S)

Zn: goes from 0 → +2; therefore
Zn is oxidized

S: goes from -2 → 0; therefore
S is reduced

Yes, this is a redox reaction

e. Example: Is the reaction below a redox reaction? If so, identify the substances that are oxidized and reduced.

$ZnS \rightarrow Zn + S$

Solution!

$\overset{(2+)}{Zn}\overset{(2-)}{S} \rightarrow \overset{(0)}{Zn} + \overset{(0)}{S}$

Yes, this is a redox reaction.

Zn^{2+} is reduced (it went down from (+2) → (0))

S^{2-} is oxidized (it went up from (-2) → (0))

f. Example: Is the reaction below a redox reaction? If so, identify the substances that are oxidized and reduced.

$Zn + Cu(NO_3)_2 \rightarrow Zn(NO_3)_2 + Cu$

Solution!

$\overset{(0)}{Zn} + \overset{(+2)}{Cu}(\overset{(+5)(-2)}{NO_3})_2 \rightarrow \overset{(+2)}{Zn}(\overset{(+5)(-2)}{NO_3})_2 + \overset{(0)}{Cu}$

$Cu \mid (NO_3^-)_2$

(+2) | $-1 \times 2 = -2$ → NO_3^{-1} (+5)(-6)

NO_3^- is the nitrate ion

Zn: goes from $0 \rightarrow +2$ oxidized

Cu^{+2}: goes from $+2 \rightarrow 0$ reduced

Yes, this is a redox reaction

g. Example: Is the reaction below a redox reaction? If so, identify the substances that are oxidized and reduced.

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Solution!

$\overset{(-4)(+1)}{CH_4} + 2\overset{(0)}{O_2} \rightarrow \overset{(+4)(-2)}{CO_2} + 2\overset{(+1)(-2)}{H_2O}$

Carbon is oxidized $(-4 \rightarrow +4)$

oxygen is reduced $(0 \rightarrow -2)$

Yes, this is a redox reaction.

h. Example: Is the reaction below a redox reaction? If so, identify the substances that are oxidized and reduced.

$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

Solution:

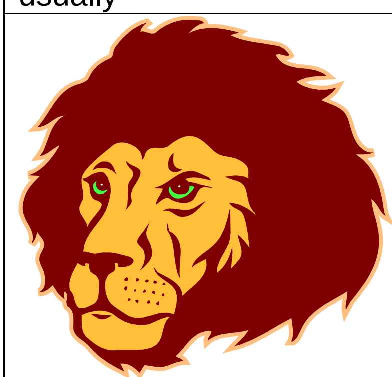
$\overset{(+1)}{Ag}\,\overset{(+5)(-2)}{NO_3}{}^{-1} + \overset{(+1)}{Na}\overset{(-1)}{Cl} \rightarrow \overset{(+1)}{Ag}\overset{(-1)}{Cl} + \overset{(+1)}{Na}\overset{(+5)}{N}\overset{(-2)}{O_3}$

solved this in earlier problem

No, this is not a redox reaction. No species go up in oxidation #; no species decrease in oxidation #.

i. Notice that synthesis, decomposition, single replacement, and combustion must *always* be redox! Just look that the lone (uncombined) elements: they always have an oxidation number of 0 when they are alone, but are always something *other than zero* when they are in a compound. Thus their oxidation numbers change and the reaction must be redox.

j. A helpful chart:

Oxidation	**Reduction**
An increase in oxidation number	A decrease in oxidation number
A loss of electrons	A gain of electrons
A gain of oxygen, usually	A loss of oxygen, usually
A loss of hydrogen, usually	A gain of hydrogen, usually
 **Figure 1. This is Leo the lion, and he goes "ger!" when he gets angry.**	"**LEO** goes **GER**!" **L**ose **E**lectrons **O**xidation **G**ain **E**lectrons **R**eduction

V. Identifying the oxidizing agent and the reducing agent

a. The substance that gets oxidized is the reducing agent.

b. The substance that gets reduced is the oxidizing agent.

c. In the example problems in the last section, identify the oxidizing agent and reducing agent.

Example IV.d) Zn is the reducing agent, S is the oxidizing agent
Example IV.e) Zn^{2+} is the oxidizing agent, S^{2-} is the reducing agent
Example IV.f) Zn is the reducing agent, Cu^{2+} is the oxidizing agent
Example IV.g) C is the reducing agent, O is the oxidizing agent
Example IV.h) There is no redox reaction, so there are no oxidizing or reducing agents.

VI. Oxidation and reduction always occur at the same time.
   a. If one substance loses electrons, then another substance must always gain electrons.
   b. Therefore, oxidation can never happen without reduction, and vice-versa.

VII. Balancing reactions using the half-reaction method
   a. Balancing redox equations in acidic solution

Balancing redox reactions using the half-reaction method

Balance the following reaction, which takes place in acidic solution:

$$MnO_4^-{}_{(aq)} + Fe^{2+}{}_{(aq)} \rightarrow Fe^{3+}{}_{(aq)} + Mn^{2+}{}_{(aq)}$$

reduction half reaction $MnO_4^- \rightarrow Mn^{2+}$

oxidation half reaction $Fe^{2+} \rightarrow Fe^{3+}$

$$5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$$

$8H^+ + MnO_4^-$: +7; with $5e^-$: +2; $Mn^{2+} + 4H_2O$: +2

A) Mn already balanced
B) balance O w/ H_2O
c) balance H w/ H^+
d) balance the charge by adding e^-'s

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

A, B, C already done

Now, # of e^-'s going must equal # of e^-'s going out

$$5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$$
$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

$$5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$$
$$5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$$

$$8H^+_{(aq)} + 5Fe^{2+}_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)} + 5Fe^{3+}_{(aq)}$$

b. Balancing redox reactions in basic solution

Balancing the follow reaction which takes place in basic solution.

$$Cr(OH)_{3\,(s)} + ClO_3^-{}_{(aq)} \rightarrow CrO_4^{2-}{}_{(aq)} + Cl^-{}_{(aq)}$$

oxidation
reduction

$$\begin{cases} Cr(OH)_3 \rightarrow CrO_4^{2-} \\ ClO_3^- \rightarrow Cl^- \end{cases}$$

$$\begin{cases} H_2O + Cr(OH)_3 \rightarrow CrO_4^{2-} + 5H^+ + 3e^- \\ 6e^- + 6H^+ + ClO_3^- \rightarrow Cl^- + 3H_2O \end{cases}$$

$$\begin{cases} \cancel{2H_2O} + 2Cr(OH)_3 \rightarrow 2CrO_4^{2-} + \cancel{10H^+}\ (4H^+) + \cancel{6e^-} \\ \cancel{6e^-} + \cancel{6H^+} + ClO_3^- \rightarrow Cl^- + \cancel{3H_2O}\ (H_2O) \end{cases}$$

solve as in acidic solution, then add OH⁻'s to each side to "cancel out" H⁺ ions.

$$2Cr(OH)_3 + ClO_3^- + 4OH^- \rightarrow 2CrO_4^{2-} + Cl^- + 4H^+ + H_2O + 4OH^-$$

$$4OH^-_{(aq)} + 2Cr(OH)_{3\,(s)} + ClO_3^-{}_{(aq)} \rightarrow 2CrO_4^{2-}{}_{(aq)} + Cl^-_{(aq)} + 5H_2O_{(l)}$$