Reduction Oxidation Reactions

Redox reactions is a shorthand way to write the words "reduction oxidation reactions." Redox reactions transfer electrons from one atom (or ion) to another atom (or ion). Redox occurs when metals react with nonmetals or during burning/combustion.

Example of a metal reacting with a nonmetal: $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$. Sodium metal reacts with chlorine gas to form sodium chloride.

Example of combustion: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$. Methane gas (natural gas) burns in oxygen gas to form carbon dioxide gas and water vapor. (You perform this reaction in lab when you light a Bunsen burner.)

Example of another redox reaction: $Cu^{2+}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{3+}(aq)$. Copper ion reacts with iron metal to form copper metal and iron ion. You perform this reaction in lab with the "iron nails" experiment, where the silvery iron nail changes to a copper color.

To tell if a redox reaction occurred, we need to keep track of the electrons that an atom has on both sides of a chemical reaction. We keep electrons by assigning oxidation numbers to individual atoms on both sides of the reaction arrow. Oxidation numbers have a similar meaning to the charge on an atom. There are seven rules to assign oxidation numbers. Follow them in order, and use the first one that applies to determine the oxidation number. The three examples above are used to explain each rule.

Seven Rules to Assign Oxidation Numbers
1. Elements always get an oxidation number of zero, 0. In the examples above, $Na(s)$, $Cl_2(g)$, $O_2(g)$, $Fe(s)$, and $Cu(s)$ all get assigned an oxidation number (or charge) of zero, because they are elements. The rule means that both the Cl atoms in $Cl_2(g)$ get an oxidation number of zero. This rule does not apply to $NaCl(s)$, $CH_4(g)$, $CO_2(g)$, and $H_2O(g)$ because these are not elements (they are compounds: $NaCl(s)$ is an ionic compound, while the rest are molecular compounds.) The rule also does not apply to $Cu^{2+}(aq)$ and Fe^{3+} because they are not elements (they are ions). (Note: if you cannot identify elements, you missed material from Lecture/Week 3.) Keep reading additional rules until reaching the first one that applies. After applying this rule, we assigned the following oxidation numbers to the examples above:

	$2Na(s)$	+	$Cl_2(g)$	$\rightarrow$	$2NaCl(s)$		
oxidation number	0		0				
	$CH_4(g)$	+	$2O_2(g)$	$\rightarrow$	$CO_2(g)$	+	$2H_2O(g)$
oxidation number			0				
	$3Cu^{2+}(aq)$	+	$2Fe(s)$	$\rightarrow$	$3Cu(s)$	+	$2Fe^{3+}(aq)$
oxidation number			0		0		

2. Monatomic ions get an oxidation number equal to the charge of the ion. In the examples above, $Cu^{2+}(aq)$ and $Fe^{3+}(aq)$ are ions. (Note: if you cannot identify ions, you missed material from Lecture/Week 6.) The compound $NaCl(s)$ is an ionic compound because it is made up of a metal and a nonmetal. Ionic compounds are made up of ions,

so we determine the oxidation numbers using this rule. To do this, use the charges given at the top of the respective columns of periodic table in which the elements appear. The charges at the top of the columns are the oxidation numbers. For Na, the charge at the top of the column (Group IA) is +1. For Cl in Group VIIA, the charge is −1. (Note: if you did not recognize ionic compounds, you missed material from Lecture/Week 12.) After applying this rule, we can assign the following oxidation numbers to the examples above:

	$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
oxidation number	0 0 +1 −1
	$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
oxidation number	0
	$3Cu^{2+}(aq) + 2Fe(s) \rightarrow 3Cu(s) + 2Fe^{3+}(aq)$
oxidation number	+2 0 0 +3

3. Hydrogen in compounds is +1. With this rule, we can assign the oxidation number for hydrogen as +1 in $CH_4(g)$ and $H_2O(g)$.

4. Oxygen in compounds is −2. With this rule, we can assign the oxidation number for oxygen as −2 in $CO_2(g)$ and $H_2O(g)$. After applying rules 3 and 4, we have the following:

	$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
oxidation number	0 0 +1 −1
	$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
oxidation number	+1 0 −2 +1−2
	$3Cu^{2+}(aq) + 2Fe(s) \rightarrow 3Cu(s) + 2Fe^{3+}(aq)$
oxidation number	+2 0 0 +3

5. In a molecular compound, assign the oxidation number of the more electronegative atom first. None of the above examples has a compound that fits this rule. As an example compound, use N_2F_4. The more electronegative atom is fluorine, so F gets assigned an oxidation number of −1 based on its column (Group VIIA) in the periodic table. Fluorine is the more electronegative element because it appears closest to the upper right corner of the periodic table. Increasing electronegativity follows the same trend as for increasing ionization energy, except that you must remember that the noble gases have zero electronegativity.

6. For neutral compounds, the sum of the oxidation numbers must be zero. Using this rule, we can assign the oxidation numbers for carbon in CH_4 and CO_2 and nitrogen in N_2F_4. CH_4 has four hydrogen atoms at +1 (Rule 3), giving a total of +4. When added to the oxidation number of carbon, the sum must be zero. Mathematically, it is $1C + 4(+1) = 0$, or $C=-4$. (The 1 coefficient of carbon comes from the 1 carbon atom in CH_4, while the 4 coefficient comes from the 4 hydrogen atoms in CH_4.) CO_2 has 2 oxygen atoms at −2 (Rule 4). The oxidation number of carbon comes from solving: $1C + 2(-2) = 0$. (The 1 coefficient of carbon comes from the 1 carbon atom in CO_2, while the 2 coefficient comes from the 2 oxygen atoms in CO_2.) Solving gives $C=+4$ in CO_2. For N_2F_4, the oxidation number for nitrogen comes from solving $2N + 4(-1) = 0$, giving $N=+2$.

7. For polyatomic ions, the sum of the oxidation numbers equals the charge on the polyatomic ion. All oxidation numbers have been assigned in the previous examples, so we need a new example. Take the polyatomic ion phosphate, PO_4^{3-}, from the page of formulas used on exams. Oxygen gets assigned an oxidation number of −2 from Rule 4. The oxidation number of phosphorous comes from solving $1P + 4(-2) = -3$, giving P=+5. Try $Cr_2O_7^{2-}$. Assign oxygen −2 (Rule 4). Calculate chromium from $2Cr + 7(-2) = -2$, giving Cr=+6.

As practice, try working these: Ca(s), N^{3-}, CaF_2, $H_4C_2O_2$, CF_4, and ClO_3^-.

Here are the answers. Ca(s) is an element, so the oxidation number is 0 by Rule 1. N^{3-} is an ion, so the oxidation number is −3 by Rule 2. CaF_2 is an ionic compound, so oxidation numbers get assigned using the periodic table. Ca is in Group IIA, so it is +2. F is in Group VIIA, so it is −1. For $H_4C_2O_2$, H is +1 by Rule 3, and O is −2 by Rule 4. The oxidation number for C comes applying Rule 6 and solving $4(+1) + 2(C) + 2(-2) = 0$, giving C=0. For CF_4, apply Rule 5 to assign F=−1, then Rule 6 to give $1C + 4(-1) = 0$. Solving gives C=+4. For ClO_3^-, assign O=−2 by Rule 4, then apply Rule 7 giving $1Cl + 3(-2) = -1$. Solving gives Cl=+5.

Now that you can assign oxidation numbers, you can tell the atoms that undergo redox (by getting oxidized and reduced). An **oxidation** (or oxidized) means an increase in oxidation number of an atom between reactants and products. A **reduction** (or reduced) means a decrease in oxidation number of an atom between reactants and products. Using the first example above, we see that Na goes from 0 (reactant) to +1 (product). This is an increase in oxidation number, so the Na atom is oxidized (or is the oxidation). Cl goes from 0 (reactant) to −1 (product). This is a decrease in oxidation number, so the Cl atom is reduced. Looking at the entire reactants (rather than just the atoms), we say that sodium metal, Na(s), was oxidized, chlorine gas, $Cl_2(g)$, was reduced. Chemists usually discuss redox reactions using this last terminology.

	$2Na(s) +$	$Cl_2(g) \rightarrow$	$2NaCl(s)$
oxidation number	0	0	+1 −1

In addition, chemists often refer to the **oxidizing agent** and the **reducing agent**. By definition, the oxidizing agent is the reactant that gets reduced (it gets reduced as it oxidizes the other reactant). The reducing agent is the reactant that gets oxidized (it gets oxidized as it reduces the other reactant). Using this terminology, chlorine gas, $Cl_2(g)$, is the oxidizing agent, and sodium metal, Na(s), is the reducing agent.

For practice, identify what gets oxidized, reduced, as well as the oxidizing agent and reducing agent for the remaining chemical reactions (the work to determine oxidation numbers appears above):

	$CH_4(g) +$	$2O_2(g) \rightarrow$	$CO_2(g) +$	$2H_2O(g)$
oxidation number	−4 +1	0	+4 −2	+1−2
	$3Cu^{2+}(aq) +$	$2Fe(s) \rightarrow$	$3Cu(s) +$	$2Fe^{3+}(aq)$
oxidation number	+2	0	0	+3

Here are the answers. For the first reaction, $CH_4(g)$ gets oxidized (C increases oxidation number from −4 in the reactant to +4 in the product), while $O_2(g)$ gets reduced (O decreases in oxidation number from 0 in the reactant to −2 in the product). As a result, $CH_4(g)$ is the reducing agent (it does the reducing) and $O_2(g)$ is the oxidizing agent (it does the oxidizing). For the second example, $Cu^{2+}(aq)$ is reduced (+2 to 0), while Fe(s) is oxidized (0 to +3). $Cu^{2+}(aq)$ is the oxidizing agent, while Fe(s) is the reducing agent.

As a final point for redox reactions, as one substance or species gets oxidized in a chemical reaction, another substance or species must get reduced. Oxidation and reduction in chemical reactions always come in pairs. Oxidation and reduction, you can't have one without the other.

Nuclear Chemistry

Nuclear chemistry deals with changes to the nucleus (protons and neutrons) of an atom. The rest of chemistry deals only with changes to the electron cloud of an atom. In nuclear chemistry for this course, the nucleus decays and emits one of three types of natural radiation, alpha particles, beta particles, or gamma rays.

Alpha particles look like helium nuclei, 2 protons, 2 neutrons, 0 electrons, and a +2 charge. Alpha particles fly out of the nucleus of a decaying atom. Because of the +2 charge, alpha particles can be deflected by a magnetic field. Because of the relatively high mass (2 protons and 2 neutrons giving 4 amu), alpha particles are the slowest moving of the three types of natural radiation and have the least penetrating power. Alpha particles can be stopped by paper, clothing, or less desirably, skin.

Beta particles look just like electrons, having nearly 0 mass and a −1 charge. Beta particles differ from electrons because beta particles fly out of the nucleus of a decaying atom at high speed. Because of the −1 charge, beta particles can be deflected by a magnetic field. Because of the relatively low mass (slightly more than 0 amu), beta particles are faster than and have more penetrating power than alpha particles. Beta particles can be stopped by 30cm of wood, aluminum foil, or less desirably, 1cm of flesh.

Gamma rays are high energy light, having exactly no mass and no charge. Gamma rays come out of the nucleus of a decaying atom at the speed of light. Because gamma rays have no charge, they cannot be deflected by a magnetic field. Because they have no mass, they have the greatest velocity of the three particles and the most penetrating power. Gamma rays can be stopped by 10cm of lead or 30cm of concrete. Because of their high energy, gamma rays pass directly through the body. Occasionally gamma rays interact with the body, often causing considerable chemical damage and mutations.

All three types of natural radiation can lead to death by causing enough chemical damage to an organism so that the organism cannot repair itself.

Only certain isotopes of atoms are radioactive. The term **nuclide** refers to the nucleus of a specific isotope. The nuclide carbon-14, ${}^{14}_{6}C$, has an atomic number (Z) of 6 (meaning 6 protons), and a mass number (protons + neutrons) equal to 14. Because the number of protons is known to be 6 for carbon from the atomic number, the number of neutrons must be 8 (6+8=14).

The half life of a radioactive nucleus is the time it takes for radiation level to drop by half of the initial value. The half life is different for each nuclide. Given this definition, you should be able to calculate the mass or amount of material remaining after a given number of half lives.

For example, start with 120mg of carbon-14, ${}^{14}_{6}C$. After 1 half life, 60mg remains (half of the original mass). After 2 half lives, 30mg remains (half of the amount at 1 half life or one-fourth of the original amount). After 3 half lives, 15mg remains (half of the amount at 2 half lives or one-eighth of the original amount).

What mass of a 512mg of carbon-14, ${}^{14}_{6}C$ remains after 6 half lives? Half of the sample disappears at each half life, so ½ * ½ * ½ * ½ * ½ * ½ = 1/64th of the original sample remains, giving 512mg/64 = 8mg.