

Oxidation-Reduction Reactions

What is an Oxidation-Reduction, or Redox, reaction?

Oxidation-reduction reactions, or redox reactions, are technically defined as any chemical reaction in which the oxidation number of the participating atom, ion, or molecule of a chemical compound changes. Some common redox reactions include fire, rusting of metals, browning of fruit, and photosynthesis.

In simpler terms, redox reactions involve the transfer of electrons from one substance to another. In a redox reaction, electrons can never be “lost”; if one substance loses electrons, another substance must gain an equal number of electrons. Therefore, oxidation and reduction always happen at the same time; they are a matched set.

What is Oxidation?

Oxidation is the loss of electrons by an atom, ion, or molecule. The atom, ion, or molecule that is oxidized will become more positively charged.

Oxidation may also involve the addition of oxygen *or* the loss of hydrogen.

What is Reduction?

Reduction is the gain of electrons by an atom, ion, or molecule. The atom, ion, or molecule that is reduced will become more negatively charged.

Reduction may also involve the loss of oxygen *or* the gain of hydrogen.

Tips for determining if an atom, ion, or compound has been Oxidized or Reduced:

- Remembering one of the two following mnemonics can be of assistance in determining the difference between oxidation and reduction:
 - 1) OIL RIG- **O**xidation **I**s **L**oss (of electrons), **R**eduction **I**s **G**ain (of electrons)
or
 - 2) LEO GER- **L**oss of **E**lectrons- **O**xidation, **G**ain of **E**lectrons- **R**eduction
- Atoms of metals tend to lose electrons to form positive ions (oxidation)
- Atoms of non-metals tend to gain electrons to form negative ions (reduction).
- Existing ions may also gain or lose electrons to become an ion with a different charge or an atom with a neutral charge.

- Always keep in mind when working with redox reactions that *electrons are negative!!*

What are “oxidation numbers” and how do I assign them?

Oxidation numbers are defined as the effective charge on an atom in a compound. A negative oxidation number, or “effective charge,” denotes the number of electrons an atom has available to donate as it participates in a chemical reaction to form a new substance. Similarly, a positive oxidation number indicates how many electrons an atom may acquire during a chemical reaction. By assigning each individual atom an oxidation number, we are able to keep track of electron transfer when new compounds are formed.

There are several general rules that are followed when assigning oxidation numbers to atoms:

- **For an atom in elemental form (an element standing alone with no charge) the oxidation number is always zero.**

Examples: S
N₂
K

S and K are atoms in the elemental form and therefore have no charge. N₂ is a molecule that is the elemental form of nitrogen, and its atoms have no charge. The oxidation number for these S, K, and N atoms will be zero.

- **For any monatomic ion (an ion consisting of only one element) the oxidation number equals the charge on the ion.**

Examples: O²⁻
H⁺

O²⁻ and H⁺ are both monatomic ions; thus, the oxidation number for O²⁻ will be -2 and the oxidation number for H⁺ will be +1.

- **Non-metals generally have negative oxidation numbers.**
- **Group I metals in a compound are always +1.**
- **Group II metals in a compound are always +2.**
- **Fluorine in a compound is always -1.**
- **Hydrogen in a compound with metals is always -1.**
- **Hydrogen in a compound with non-metals is always +1.**

- Oxygen in a compound is generally -2 (UNLESS in peroxides or with fluorine, in which case it is -1).
- Halogens (Group VII) in a compound are generally -1.
- The sum of the oxidation numbers of all atoms in a neutral compound is zero.

Example: Determine the oxidation number on S in Na_2SO_3 .

Na_2SO_3 is neutral and has no overall charge. Therefore, the sum of the charges on each atom in the compound must equal zero. We will start by identifying the overall charge on oxygen since we know that it will generally always be -2; there are 3 of them so the overall charge for oxygen is -6 ($-2 \times 3 = -6$). Likewise, we know that sodium will always have a charge of +1 and there are 2 of them; thus, the overall charge on sodium is +2 ($+1 \times 2 = +2$). In order to figure out the unknown charge (and thus the oxidation number) on the sulfur ion, we must set up a mathematical equation:

$$-6 \text{ (-2 charge times 3 atoms of O)} + 2 \text{ (+1 charge times 2 ions of Na)} + X \text{ (x charge times 1 atom of S)} = 0 \text{ (Overall charge on Na}_2\text{SO}_3\text{)}$$

$$-4 \text{ (Sum of charges on O atoms and Na ions)} + X \text{ (Charge of S atom)} = 0 \text{ (Overall charge on Na}_2\text{SO}_3\text{)}$$

In solving for X, we determine that the oxidation number on sulfur is +4.

Example: Determine the oxidation number on C in COCl_2 .

COCl_2 is neutral and has no overall charge. As with the last example, the sum of the overall charges on each atom must equal zero. Oxygen generally has a charge of -2; there is only 1 of them so the overall charge on oxygen is -2 ($-2 \times 1 = -2$). Chlorine (a halogen) generally has a charge of -1; since there are 2 of them, the overall charge on chlorine is -2 ($-1 \times 2 = -2$). At this point, we are ready to set up our mathematical equation to solve for the unknown charge (oxidation number) on carbon.

$$-2 \text{ (-2 charge times 1 atom of O)} + -2 \text{ (-1 charge times 2 atoms of Cl)} + X \text{ (x charge times 1 atom of C)} = 0 \text{ (Overall charge on COCl}_2\text{)}$$

$$-4 \text{ (Sum of charges on O and Cl atoms)} + X \text{ (Charge on C atom)} = 0 \text{ (Overall charge on COCl}_2\text{)}$$

In solving for X, we can determine that the oxidation number on carbon to be +4.

- When a monatomic ion is attached to a polyatomic ion, the known charge of the polyatomic ion can be used to determine the oxidation number of the monatomic ion.

Example: Determine the oxidation number on Cu in CuSO_4 .

SO_4^{2-} has a charge of -2. We know that the compound CuSO_4 has an overall charge of 0 (neutral). Therefore, we can set up a mathematical equation to solve for the overall charge (oxidation number) on Cu.

$$-2 \text{ (Charge on } \text{SO}_4 \text{ ion)} + X \text{ (Charge on Cu)} = 0 \text{ (Overall charge on } \text{CuSO}_4\text{)}$$

In solving for X, we determine that the oxidation number on copper is +2.

- The sum of the oxidation numbers in a polyatomic ion equals the charge on the ions.

Example: Determine the oxidation number on S in SO_4^{2-} .

SO_4^{2-} has an overall charge of -2. This means that the sum of the overall charges on each atom in the ion must equal -2. Again, we will begin with oxygen. We know that generally oxygen carries a charge of -2; there are 4 oxygen atoms in this ion meaning that the overall charge on oxygen is -8 (-2 x 4 = -8). At this point, we can set up a mathematical equation to solve for the unknown charge (oxidation number) on S.

$$-8 \text{ (-2 charge times 4 O atoms)} + X \text{ (Charge on S)} = -2 \text{ (Overall charge on } \text{SO}_4^{2-}\text{)}$$

In solving for X, we determine that the oxidation number on sulfur is +6.

Example: Determine the oxidation number on Mn in MnO_4^- .

In MnO_4^- , the overall charge on the molecule is -1. This tells us that the sum of the overall charges on each atom in the ion must equal -1. Beginning with oxygen, we can determine the overall charge on oxygen to be -8 (-2 x 4 = -8). We can now set up our mathematical equation in order to solve for the unknown charge (oxidation number) on manganese.

$$-8 \text{ (-2 charge times 4 O atoms)} + X \text{ (Charge on Mn)} = -1 \text{ (Overall charge on } \text{MnO}_4^- \text{)}$$

In solving for X, we determine that the oxidation number on manganese is +7.

How do I tell what is oxidized and what is reduced in a reaction?

There are several steps involved in determining what is oxidized and what is reduced in a reaction:



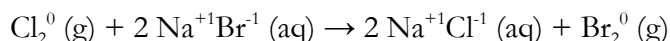
- Using the rules for determining oxidation numbers above, write out the charge (oxidation number) held by each atom or molecule in a reaction.

I can determine that $\text{Cl}_2(\text{g})$ and $\text{Br}_2(\text{g})$ will both have a charge of zero (both are in elemental form). We know that both $\text{NaBr}(\text{aq})$ and $\text{NaCl}(\text{aq})$ are neutral compounds; therefore, the sum of the charges of the atoms in each of these compounds will be zero.

NaBr : We know that the charge on Na is +1 and that the charge on Br is -1. The sum of these is equal to zero (+1 + -1 = 0).

NaCl : The charges on Na and Cl can be determined in the same manner as NaBr above: Na has a charge of +1 and Cl has a charge of -1. This brings their sum to zero (+1 + -1 = 0).

I can now write out the reaction again, this time including the charges on each atom.



- I now determine the change in charge on each atom from products to reactants and whether electrons were gained or lost.

$\text{Cl}^0 \rightarrow \text{Cl}^{-1}$ Cl has gone from a charge of 0 to -1; Cl has **GAINED** one electron (***Remember electrons are negative!!***)

$\text{Na}^{+1} \rightarrow \text{Na}^{+1}$ Na has maintained the same charge, meaning it has neither gained nor lost any electrons.

$\text{Br}^{-1} \rightarrow \text{Br}^0$ Br has gone from a charge of -1 to 0; Br has **LOST** one electron.

- Based on whether electrons were gained or lost (***Remember LEO GER or OIL RIG!!***), I can tell which atom is oxidized and which is reduced.

Cl has gained one electron; therefore, Cl is reduced.

Na has neither gained nor lost so it is neither oxidized nor reduced.

Br has lost one electron; therefore, Br is oxidized.

What is the “Activity Series” and how do I use it?

Metals vary based on how easily they are oxidized (or lose electrons). The activity series is a list of metals in aqueous solutions that shows the relative ease with which metals are oxidized. It allows one to be able to predict the outcome of reactions between metals, metal salts, and acids. Metals at the top of the list are more easily oxidized (lose electrons more easily) than metal at the bottom of the list.

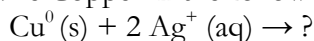
The activity series is as follows:

Metal	Oxidation Reaction				
Lithium	Li	\rightleftharpoons	Li ⁺	+	e ⁻
Rubidium	Rb	\rightleftharpoons	Rb ⁺	+	e ⁻
Potassium	K	\rightleftharpoons	K ⁺	+	e ⁻
Barium	Ba	\rightleftharpoons	Ba ²⁺	+	2e ⁻
Calcium	Ca	\rightleftharpoons	Ca ²⁺	+	2e ⁻
Sodium	Na	\rightleftharpoons	Na ⁺	+	e ⁻
Magnesium	Mg	\rightleftharpoons	Mg ²⁺	+	2e ⁻
Aluminum	Al	\rightleftharpoons	Al ³⁺	+	3e ⁻
Manganese	Mn	\rightleftharpoons	Mn ²⁺	+	2e ⁻
Zinc	Zn	\rightleftharpoons	Zn ²⁺	+	2e ⁻
Chromium	Cr	\rightleftharpoons	Cr ³⁺	+	3e ⁻
Iron	Fe	\rightleftharpoons	Fe ²⁺	+	2e ⁻
Cobalt	Co	\rightleftharpoons	Co ²⁺	+	2e ⁻
Nickel	Ni	\rightleftharpoons	Ni ²⁺	+	2e ⁻
Tin	Sn	\rightleftharpoons	Sn ²⁺	+	2e ⁻
Lead	Pb	\rightleftharpoons	Pb ²⁺	+	2e ⁻
Hydrogen	H ₂	\rightleftharpoons	2 H ⁺	+	2e ⁻
Copper	Cu	\rightleftharpoons	Cu ²⁺	+	2e ⁻
Silver	Ag	\rightleftharpoons	Ag ⁺	+	e ⁻
Mercury	Hg	\rightleftharpoons	Hg ²⁺	+	2e ⁻
Platinum	Pt	\rightleftharpoons	Pt ²⁺	+	2e ⁻
Gold	Au	\rightleftharpoons	Au ³⁺	+	3e ⁻

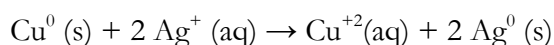
<http://www.grandinetti.org/Teaching/Chem121/Lectures/SolutionReactions/assets/ActivitySeries.png>

- Any metal on the list can be oxidized by (lose electrons to) the ions of the elements below it (ion will be reduced or gain electrons).
- Metals that are oxidized go from a solid form into solution and ions of the elements that do the oxidizing come out of solution to form solids.

Example: Will Silver oxidize Copper in the following equation:



Silver is below copper in the activity series; therefore, copper will be oxidized by silver and the equation will be as follows:

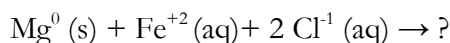


Copper loses electrons. It goes from a charge of zero to a charge of +2; therefore, copper is oxidized.

Silver gains electrons. It goes from a charge of +1 to a charge of 0 (remember electrons are negative!!); therefore, silver is reduced.

Example: Will an aqueous solution of iron (II) chloride oxidize magnesium metal? If so, write the balanced molecular equation.

The first step in solving this equation is to set up the reactants side of the equation, including the oxidation number of each:

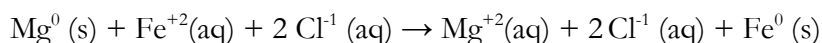


The next step in solving this equation is to look at the activity series; if iron is below magnesium in the activity series, then iron will oxidize magnesium.

Magnesium	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$
Aluminum	$\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-$
Manganese	$\text{Mn} \rightarrow \text{Mn}^{2+} + 2\text{e}^-$
Zinc	$\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$
Chromium	$\text{Cr} \rightarrow \text{Cr}^{3+} + 3\text{e}^-$
Iron	$\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$

Iron is indeed below magnesium, meaning that the iron ion will oxidize magnesium, or cause it to lose electrons.

Because we now know that this reaction will occur, we write out the products and oxidation numbers on each.



Magnesium has gone from a solid form into solution and has been oxidized (lost electrons; went from an oxidation number of 0 to +2).

Iron has been pulled out of solution to form a solid and has been reduced (gained electrons; went from an oxidation number of +2 to 0).

Sample Problems

Determine the oxidation number on each of the following:

1. Na (s)
2. F^{-1} (aq)
3. Li^{+1} (aq)

Determine the oxidation number for the indicated element in each of the following substances:

1. S in SO_2 (g)
2. S in Na_2SO_3 (aq)
3. Cr in $Cr_2O_7^{2-}$ (aq)

Determine which element is being oxidized and which is being reduced in the following reactions:

1. N_2 (g) + $3H_2$ (g) \rightarrow $2NH_3$ (g)
2. Mg (s) + $CoSO_4$ (aq) \rightarrow $MgSO_4$ (aq) + Co (s)
3. $2Al$ (s) + $6HBr$ (aq) \rightarrow $2AlBr_3$ (aq) + $3H_2$ (g)

Use the activity series to determine the following. If a reaction will occur, write a balanced molecular equation.

1. Mn (s) + $NiCl_2$ (aq) \rightarrow
2. Fe (s) + $CuSO_4$ (aq) \rightarrow
3. Cu (s) + $Cr(CH_3COO)_3$ (aq) \rightarrow

Answers to Sample Problems

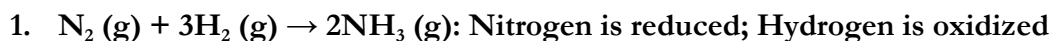
Determine the oxidation number on each of the following:

1. **Na (s): 0** Sodium is in elemental form; the oxidation number is therefore zero.
2. **F⁻¹ (aq): -1** The fluorine ion is monatomic; the oxidation number equals the charge on the ion.
3. **Li⁺¹ (aq): +1** The lithium ion is monatomic; the oxidation number equals the charge on the ion.

Determine the oxidation number for the indicated element in each of the following substances:

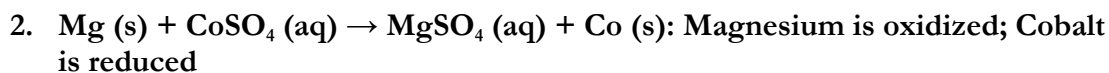
1. **S in SO₂ (g): +4** Total charge on oxygens: $(2 \times -2) = -4$. Since the molecule is neutral, positive and negative charges must cancel out:
$$-4 \text{ (charges on oxygens)} + X \text{ (charge on sulfur)} = 0 \text{ (overall charge)}$$
Therefore, the charge on sulfur must be +4.
2. **S in Na₂SO₃ (aq): +4** Total charge on oxygens: $(3 \times -2) = -6$. Total charge on sodiums: $(2 \times +2) = +4$. Since the molecule is neutral, positive and negative charges must cancel out:
$$-6 \text{ (charges on oxygens)} + 4 \text{ (charges on sodiums)} + X \text{ (charge on sulfur)} = 0 \text{ (overall charge)}$$
Therefore the charge on sulfur must be +4.
3. **Cr in Cr₂O₇²⁻ (aq): +6** Total charge on oxygens: $(7 \times -2) = -14$. The ion has an overall charge of -2; when added together, the charges of chromium and oxygen must equal -2:
$$-14 \text{ (charges on oxygens)} + 2X \text{ (charge on each chromium)} = -2 \text{ (overall charge)}$$
Therefore, the charge on chromium will be +6.

Determine which element is being oxidized and which is being reduced in the following reactions:



Nitrogen goes from a charge of 0 to -3; it is gaining electrons.

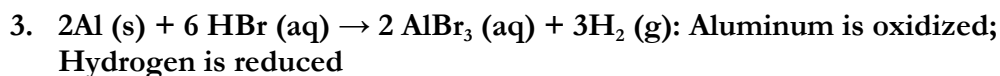
Each hydrogen goes from a charge of 0 to +1, for a total of +3 for all three; they are losing electrons.



Magnesium goes from a charge of 0 to +2; it is losing electrons.

Cobalt is going from a charge of +2 to 0; it is gaining electrons.

The charge on SO_4^{2-} doesn't change; therefore, it is neither oxidized nor reduced. Also the atoms within the polyatomic ion are neither oxidized nor reduced.

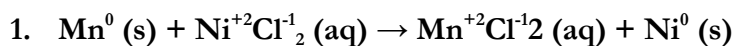


Aluminum goes from a charge of 0 to +3; it is losing electrons.

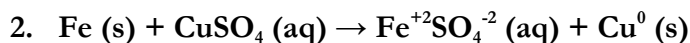
Each hydrogen goes from a charge of +1 to 0; it is gaining electrons.

The charge on bromide ion doesn't change; therefore it is neither oxidized nor reduced.

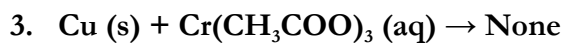
Use the activity series to determine if a reaction will occur; if so, write a balanced molecular equation.



Nickel will oxidize manganese because nickel is below manganese on the activity series.



Copper will oxidize iron because copper is below iron on the activity series.



Chromium will not oxidize copper because chromium is above copper on the activity series.