

Aqueous Reactions

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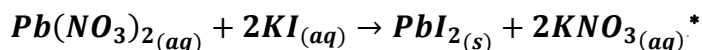
Defining Aqueous Reactions

Aqueous reactions are reactions that take place in water. To understand them, it is important to understand how compounds behave in water. Some compounds are **electrolytes**- they dissociate into separate ions in water. However, not all electrolytes behave the same way. Some are **strong electrolytes**, and dissociate completely, so no ions are left bonded together. Others are **weak electrolytes**- they only partly dissociate, and many of their ions are still bonded to each other. Other substances, **nonelectrolytes**, do not dissociate at all.

There are three main types of aqueous reactions: precipitation reactions, acid-base reactions, and oxidation-reduction (or redox) reactions.

Precipitation Reactions

Precipitation reactions produce an insoluble product- the precipitate. They contain two aqueous reactants, one aqueous product, and one solid product.

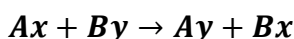


In this reaction, two soluble products, $\text{Pb}(\text{NO}_3)_2$ and KI , combine to form one soluble product, KNO_3 , and one insoluble product, PbI_2 . This is a precipitation reaction, and PbI_2 is the precipitate.

* In chemical equations, certain abbreviations are used to indicate the state of the substances involved. The abbreviations are as follows: s = solid; l = liquid; g = gaseous; aq = aqueous, or soluble in water.

Determining the Products of a Precipitation Reaction

To determine the products of a precipitation reaction, reverse the cation-anion pairs.** For example, at the beginning of the above reaction, lead is bonded to nitrate, and potassium is bonded to iodine. The products are these pairs reversed- lead with iodine, and potassium with nitrate. Precipitation reactions follow this formula:



The products are just the cation-anion pairs reversed, or the “outies” (A and Y joined) and the “innies” (B and X joined).

Determining Whether or Not a Reaction is a Precipitation Reaction

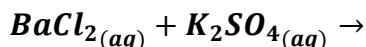
Once you know the products of a reaction, you can use the solubility rules to see if you have an insoluble product, and thus, a precipitation reaction.

Solubility Rules:

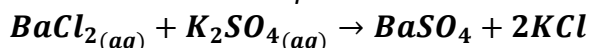
Soluble: All ionic compounds containing: 1. Alkali metals (group 1A) 2. Ammonium (NH_4^+) 3. Nitrate (NO_3^-) 4. Acetate ($\text{C}_2\text{H}_3\text{O}_2^-$) 5. Chloride ($\text{Cl}^-$) 6. Bromide ($\text{Br}^-$) 7. Iodide ($\text{I}^-$) 8. Sulfate ($\text{SO}_4^{2-}$)	Exceptions: 1. None 2. None 3. None 4. None 5. AgCl, PbCl_2, Hg_2Cl_2 6. AgBr, PbBr_2, Hg_2Br_2 7. AgI, PbI_2, Hg_2I_2 8. SrSO_4, BaSO_4, Hg_2SO_4, PbSO_4, CaSO_4
Insoluble: Compounds containing: 1. S^{2-} 2. CO_3^{2-} 3. PO_4^{3-} 4. OH^-	Exceptions: 1. When bonded to ammonium, alkali metals, Ca^{2+}, Sr^{2+}, or Ba^{2+} 2. When bonded to ammonium or alkalis 3. Same as above 4. When bonded to alkali metals, Ca^{2+}, Sr^{2+}, or Ba^{2+}

** If you need help determining the formulas for these new ionic compounds from the ions, look at the Academic Center for Excellence’s handout, “Naming Compounds.”

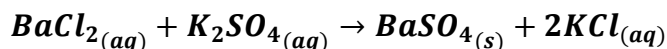
Example: Predict the products formed by the aqueous reaction below and determine whether or not the reaction is a precipitation reaction.



The first step is to predict the products, which we do by reversing the pairs, giving us BaSO_4 , and KCl . *Remember to balance the equation.*

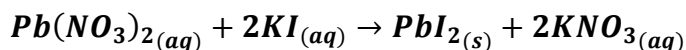


Next, we use the solubility rules to determine if this is a precipitation reaction. The table tells us that compounds containing alkali metals, such as potassium, are soluble—thus KCl is soluble. We also see that sulfate is soluble *except when bonded to barium!* Thus, BaSO_4 is insoluble, and this is a precipitation reaction. The whole balanced equation is:

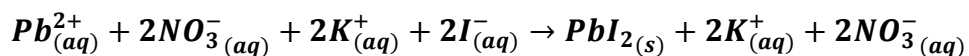


Ionic Equations

Something that is useful when dealing with precipitation reactions is the ability to write ionic equations, which show the compounds as individual ions. Until now, you have been writing chemical equations in this form:



Equations written this way are known as **molecular equations**. They have a variation known as a **complete ionic equation**, in which all soluble strong electrolytes are written as individual ions. Thus, the above reaction becomes:



Now, each soluble strong electrolyte is written as separate ions. The equation is still balanced (for example, there are two nitrate ions in the compound $\text{Pb}(\text{NO}_3)_2$, so NO_3^{-} has a coefficient of 2), and everything is in the same state (aqueous or solid).

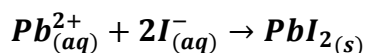
When writing a complete ionic equation, remember that only soluble strong electrolytes are written as individual ions. You already have the guidelines for determining if something is soluble; below is a table which can be used to determine if a substance is a strong, weak, or nonelectrolyte.

Electrolytic Behavior of Soluble Compounds:

	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic Compound	All	None	None
Molecular Compound	Strong acids (coming later!)	Weak acids and bases (coming later!)	All other compounds

Remember, only *soluble* strong electrolytes are written as individual ions. Thus, in the above equation, although PbI_2 is an ionic compound and thus a strong electrolyte, it is *not* written as separate ions because it is insoluble.

A shorter ionic equation is the **net ionic equation**. In the complete ionic equation above, the potassium and nitrate ions appear in identical forms on both sides of the equation. The lead and iodine ions undergo a change from individual ions to an insoluble compound, but the potassium and nitrate ions do not. Ions which appear in identical forms on both sides are called spectator ions, and do not actively participate in the reaction. If we eliminate them, the net ionic equation is left. The net ionic equation of the above reaction looks like this:



We simply took out the spectator ions- the potassium and nitrate- and ended with the net ionic equation.

To summarize the series of steps to get from one form of an equation to another:

1. Write a balanced molecular equation, just like you've been doing.
2. Rewrite the equation, showing all soluble, strong electrolytes as individual ions, to get the complete ionic equation. Keep it balanced.
3. Eliminate all spectator ions to get the net ionic equation.

Acid-Base Reactions

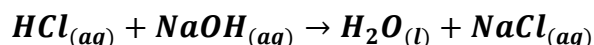
Acids are substances that release H^+ ions in water. Bases accept these H^+ ions, and produce OH^- in water (occasionally a base such as ammonia, NH_3 , won't contain OH^- . Most bases, though, contain hydroxide). Like other electrolytes, there are both strong and weak acids and bases. It is important to know the strong acids and bases from the weak:

Strong Acids	Common Strong Bases
1. Hydrochloric acid (HCl)	1. LiOH
2. Hydrobromic acid (HBr)	2. NaOH
3. Hydroiodic acid (HI)	3. KOH
4. Chloric acid ($HClO_3$)	4. RbOH
5. Perchloric acid ($HClO_4$)	5. CsOH
6. Nitric acid (HNO_3)	6. $Ca(OH)_2$
7. Sulfuric acid (H_2SO_4)	7. $Sr(OH)_2$
	8. $Ba(OH)_2$

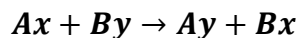
All other acids and almost all other bases you will encounter are weak.

Acid-Base Reactions: Neutralization Reactions

When acids and bases react, a **neutralization reaction** occurs. In this reaction, the acid donates an H^+ ion. This joins with the hydroxide ion from the base to form water, while the anion from the acid and the cation from the base join to form an ionic compound. Here is a typical acid base reaction:

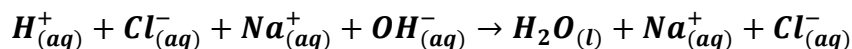


In this reaction, hydrochloric acid joins with sodium hydroxide, a base. The H^+ from the acid, and the OH^- from the base join to form water, while the Cl^- and Na^+ ions join to form sodium chloride. In fact, these neutralization reactions have the same form as the precipitation reactions we looked at earlier:

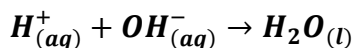


In general, any neutralization reaction between an acid and base produces water and a **salt** (the term salt here refers to any ionic compound whose cation comes from a base, and whose anion comes from an acid- for example, Na^+ from NaOH, and Cl^- from HCl in the above reaction).

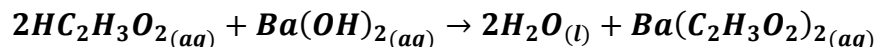
Now, write the complete ionic form of the above reaction. You should get:



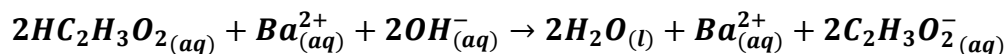
Next, write the net ionic equation. It should look like this:



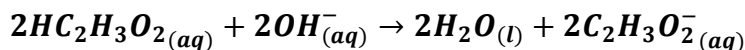
This equation summarizes the essential feature of the reaction between any strong acid and strong base- a hydrogen ion and a hydroxide ion forming water. Any reaction between a strong acid and a strong base can be brought back to this. Only strong acid-strong base reactions can be reduced this way, however. With weak acids or bases, it's a bit different, as we see in this reaction involving acetic acid:



When writing the complete ionic equation for this reaction, remember that acetic acid is a weak electrolyte. Thus, the complete ionic equation is:



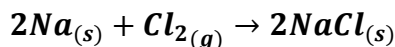
If we write the net ionic equation, it comes to:



Because this reaction involves a weak acid, we cannot reduce it to the same short equation that we could in our strong acid-strong base reaction.

Oxidation-Reduction Reactions

Oxidation-reduction, or **redox reactions**, take place between metals and non-metals. Below is a good example of a redox reaction:



At the beginning of the reaction, the sodium metal is composed of atoms—all with a net charge of zero (the positive charges of sodium's 11 protons are balanced by the negative charges of its 11 electrons). The same is true of the chlorine gas (each chlorine atom has 17 protons and 17 electrons). However, the product contains sodium in the form Na^+ and chlorine in the form Cl^- . Somewhere along the line, these neutral atoms became ions. This is the result of a transfer of electrons. In the above reaction, an electron is transferred from a sodium atom to a chlorine atom. The sodium atoms now have ten electrons, instead of eleven. This means that the sodium has one more proton than electron and has a charge of +1. The chlorine now has one more electron than proton and has a charge of -1. The positively charged sodium ion and the negatively charged chloride ion are attracted to each other and bond—forming sodium chloride. A reaction like this, that involves the transfer of electrons, is a redox reaction. For further information on redox reactions, see the [ACE handout *Oxidation-Reduction Reactions*](#).

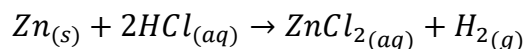
Oxidation and Reduction

During a redox reaction, two processes take place—oxidation and reduction. When something loses electrons and becomes more positively charged, it has been **oxidized**. When something gains electrons, becoming more negatively charged, it has been **reduced** (think of the *gain* of electrons as *reducing* the positive charge). We also say that the substance that is reduced is the **oxidizing agent** (causes the oxidation of the other), while the substance that is oxidized is the **reducing agent** (causes the reduction of the other). In the above reaction, sodium, which loses an electron, is oxidized while chlorine, which gains the electron, is reduced. Thus, sodium is the reducing agent, and chlorine is the oxidizing agent. In order to determine which part of a reaction is being oxidized and which is being reduced, we use **oxidation numbers**. These are numbers used to keep track of the electrons in a reaction. Each atom is assigned one oxidation number on the reactant side of the equation and another on the product side. The change in oxidation numbers from one side of the reaction to the other lets us know if something has been oxidized or reduced. The rules for oxidation numbers are as follows:

1. **The oxidation number for an atom in its elemental form is always zero.**
 - i. Each atom in the H_2 molecule has an oxidation number of zero, as does each atom in the P_4 molecule.
 - ii. If you are not sure if something is in its elemental form, look up the form in which it appears in either Appendix 4 of *Chemistry: The Science in Context*, or Appendix C of *Chemistry: The Central Science*. Once you have located the form of the element you are dealing with, look at the value in the ΔH°_f column. If the ΔH°_f value for that particular form is zero, then it is in its elemental state.

- 2. The oxidation number for any monoatomic ion always equals the charge on that ion.**
 - i. Thus K^+ has an oxidation number of +1, and S^{2-} has an oxidation number of -2 (when writing charges, put the number before the sign; when writing oxidation numbers, put the sign before the number).
 - ii. The alkali metals (Group 1A) always have an oxidation number of +1; the alkaline earth metals (Group 2A) always have an oxidation number of +2, and aluminum always has an oxidation number of +3.
- 3. Nonmetals usually have negative oxidation numbers, but occasionally have positive ones.**
 - i. Oxygen usually has an oxidation number of -2. The main exception is in compounds containing peroxide (O_2^{2-}), in which each oxygen atom has an oxidation number of -1.
 - ii. Hydrogen has an oxidation number of +1 when bonded to nonmetals and -1 when bonded to metals.
 - iii. Fluorine always has an oxidation number of -1. The other halogens have an oxidation number of -1 in most binary compounds; however, when combined with oxygen, they have positive oxidation states.
- 4. The sum of the oxidation numbers of all atoms in a compound is zero; the sum of oxidation numbers of all atoms in a polyatomic ion equals the charge of the ion.**
 - i. In NaCl, the sum of the oxidation numbers of the sodium and the chloride is zero.
 - ii. In H_3O^+ , the sum of the oxidation numbers of the hydrogen and the oxygen is +1.
 - iii. This rule is very useful for finding the oxidation number of an atom in an ion or compound if you know the oxidation numbers of the other atoms.
- 5. The change in an oxidation number from one side of the equation to another indicates whether a substance has been oxidized or reduced.**
 - i. If the oxidation number of an atom becomes more positive, the atom has lost electrons and been oxidized
 - ii. If the oxidation number of an atom becomes more negative, the atom has gained electrons and been reduced.

Let's look at an example of a problem involving oxidation numbers.



What are the oxidation numbers on the zinc, chlorine, and hydrogen? What is being oxidized, and what is being reduced?

Let's start by assigning oxidation numbers on the reactant side of the equation. The zinc is in its elemental form, so its oxidation number is zero. The hydrogen is bonded to a non-metal, so its oxidation number is +1. Chlorine forms monoatomic ions, so its oxidation number is the same as its charge: -1. We know these numbers for the hydrogen and chlorine are right because HCl is a compound, so the oxidation numbers in it must add up to zero, which -1 and +1 indeed do.

Now let's look at the product side. Here, chlorine is bonded to zinc. Again, its oxidation number equals its charge, and is -1. Now we need to determine the oxidation number of the zinc. We know that ZnCl_2 is a compound, so its oxidation numbers must add up to zero. We also know that the oxidation number on each chlorine atom in the compound is -1. Because there are two chlorine atoms, the total value of the oxidation numbers provided by chlorine is -2. There is only one zinc atom, so if we are to get a value of zero for the compound, the oxidation number on the zinc must be +2. The oxidation number on each hydrogen atom is zero, since the hydrogen is now in its elemental form.

To summarize: the zinc has an oxidation number of 0 on the reactant side, and +2 on the product side; the chlorine has an oxidation number of -1 on each side; and the hydrogen has an oxidation number of +1 on the reactant side, and 0 on the product side. Because the zinc's oxidation number becomes more positive, we know that it has lost electrons and been oxidized. Conversely, because the hydrogen's oxidation number has become more negative, we know that it has gained electrons and been reduced. The chlorine's oxidation number did not change, so it was neither oxidized nor reduced.

Practice Problems

- Predict the products for the following reactions. Write the balanced equations, determine whether or not the reactions are precipitation reactions, and identify the precipitate. Write the complete and net ionic equations for all precipitation reactions that are present.
 - $Ba(NO_3)_{2(aq)} + NaCl_{(aq)} \rightarrow$
 - $Na_2S_{(aq)} + Cu(NO_3)_{2(aq)} \rightarrow$
 - $NH_4Cl_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow$
 - $MgSO_{4(aq)} + BaCl_{2(aq)} \rightarrow$
 - $KNO_{3(aq)} + BaCl_{2(aq)} \rightarrow$
- Complete and balance each of the following equations. Identify the acid and the base in each reaction, and write the complete and net ionic equations.
 - $H_2SO_{4(aq)} + Ca(OH)_{2(aq)} \rightarrow$
 - $HClO_{4(aq)} + NaOH_{(aq)} \rightarrow$
 - $HBr_{(aq)} + KOH_{(aq)} \rightarrow$
 - $HNO_{3(aq)} + Al(OH)_{3(s)} \rightarrow$
 - $HC_2H_3O_{2(aq)} + KOH_{(aq)} \rightarrow$
- Give the oxidation number for chlorine in each of the following molecules:
 - HClO
 - HClO₃
 - ClO₄⁻
- Give the oxidation number for nitrogen in each of the following molecules:
 - N₂
 - N₂H₄
 - NH₄⁺
- Balance each of the following redox reactions. Assign each atom an oxidation number. Determine which is being oxidized and which is being reduced.
 - $Na_{(s)} + S_{(s)} \rightarrow Na_2S_{(s)}$
 - $Mg_{(s)} + O_{2(g)} \rightarrow MgO_{(s)}$
 - $Ca_{(s)} + F_{2(g)} \rightarrow CaF_{2(s)}$
 - $Fe_{(s)} + Cl_{2(g)} \rightarrow FeCl_{3(s)}$
 - $FeO_{(s)} + O_{2(g)} + H_2O_{(l)} \rightarrow Fe(OH)_{3(s)}$

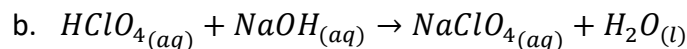
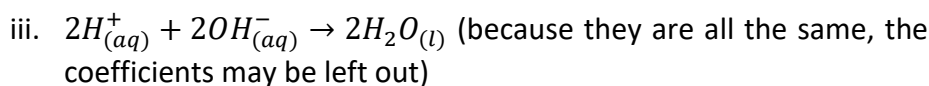
Answer Key

1. Precipitation reactions

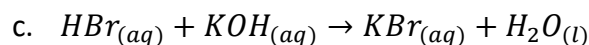
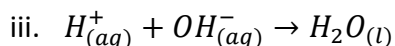
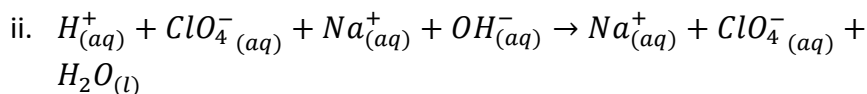
- a. $Ba(NO_3)_2(aq) + 2NaCl(aq) \rightarrow 2NaNO_3(aq) + BaCl_2(aq)$
 - i. This is not a precipitation reaction.
- b. $Na_2S(aq) + Cu(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + CuS(s)$
 - i. This is a precipitation reaction. CuS is the precipitate.
 - ii. $2Na^+_{(aq)} + S^{2-}_{(aq)} + Cu^{2+}_{(aq)} + 2NO_3^-_{(aq)} \rightarrow 2Na^+_{(aq)} + 2NO_3^-_{(aq)} + CuS(s)$
 - iii. $S^{2-}_{(aq)} + Cu^{2+}_{(aq)} \rightarrow CuS(s)$
- c. $2NH_4Cl(aq) + Pb(NO_3)_2(aq) \rightarrow 2NH_4NO_3(aq) + PbCl_2(s)$
 - i. This is a precipitation reaction. PbCl₂ is the precipitate.
 - ii. $2NH_4^+_{(aq)} + 2Cl^-_{(aq)} + Pb^{2+}_{(aq)} + 2NO_3^-_{(aq)} \rightarrow 2NH_4^+_{(aq)} + 2NO_3^-_{(aq)} + PbCl_2(s)$
 - iii. $2Cl^-_{(aq)} + Pb^{2+}_{(aq)} \rightarrow PbCl_2(s)$
- d. $MgSO_4(aq) + BaCl_2(aq) \rightarrow MgCl_2(aq) + BaSO_4(s)$
 - i. This is a precipitation reaction. BaSO₄ is the precipitate.
 - ii. $Mg^{2+}_{(aq)} + SO_4^{2-}_{(aq)} + Ba^{2+}_{(aq)} + 2Cl^-_{(aq)} \rightarrow Mg^{2+}_{(aq)} + 2Cl^-_{(aq)} + BaSO_4(s)$
 - iii. $SO_4^{2-}_{(aq)} + Ba^{2+}_{(aq)} \rightarrow BaSO_4(s)$
- e. $2KNO_3(aq) + BaCl_2(aq) \rightarrow 2KCl(aq) + Ba(NO_3)_2(aq)$
 - i. This is not a precipitation reaction.

2. Acid-Base Reactions

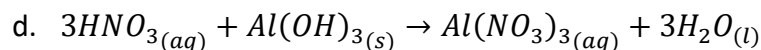
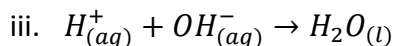
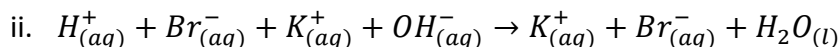
- a. $H_2SO_4(aq) + Ca(OH)_2(aq) \rightarrow CaSO_4(aq) + 2H_2O(l)$
 - i. H₂SO₄ is the acid, and Ca(OH)₂ is the base.
 - ii. $2H^+_{(aq)} + SO_4^{2-}_{(aq)} + Ca^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Ca^{2+}_{(aq)} + SO_4^{2-}_{(aq)} + 2H_2O(l)$



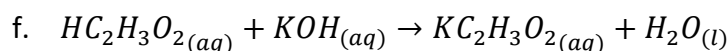
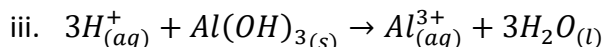
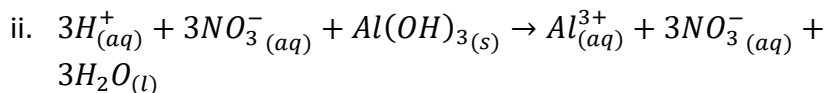
i. $HClO_4$ is the acid, and $NaOH$ is the base.



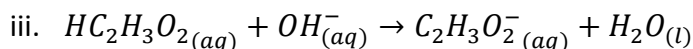
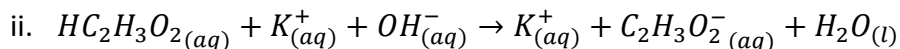
i. HBr is the acid, and KOH is the base



i. HNO_3 is the acid, and $Al(OH)_3$ is the base



i. $HC_2H_3O_2$ is the acid, and KOH is the base



3. Oxidation numbers of chlorine

a. +1 (H= +1, O= -2. All oxidation numbers must add up to zero)

b. +5 (H= +1, O= -2, and there are three O atoms for a total of -6. All numbers must add up to zero)

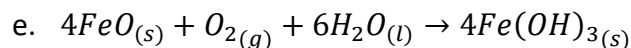
c. +7 (O= -2, and there are four O atoms for a total of -8. All oxidation numbers must add up to -1 since that is the charge on the ion)

4. Oxidation numbers of nitrogen

- a. 0 (It is in its elemental state)
- b. -2 (H= +1, and there are four H atoms for a total of +4. The oxidation numbers must add up to zero, and there are two N atoms, so the two N atoms combined must add up to -4. Therefore, the oxidation number on an individual N atom in this compound is -2)
- c. -3 (H= +1, and there are four H atoms, for a total of +4. The oxidation numbers must add up to +1, which is the charge on the ion, so the nitrogen must be -3)

5. Oxidation-Reduction Reactions

- a. $2Na_{(s)} + S_{(s)} \rightarrow Na_2S_{(s)}$
 - i. Reactant ox. nos.: Na = 0 (elemental state); S = 0 (elemental state)
 - ii. Product ox. nos.: Na = +1 (charge on ion); S = -2 (there are two Na atoms with an ox. no. of +1 each and the numbers must add up to zero)
 - iii. Na is oxidized; S is reduced
- b. $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$
 - i. Reactant ox. nos.: Mg = 0 (elemental state); O = 0 (elemental state)
 - ii. Product ox. nos.: Mg = +2 (charge on ion); O = -2 (oxygen is always -2, unless in its peroxide form; this also makes the numbers add up to zero)
 - Mg is oxidized; O is reduced
- c. $Ca_{(s)} + F_{2(g)} \rightarrow CaF_{2(s)}$
 - i. Reactant ox. nos.: Ca = 0 (elemental state); F = 0 (elemental state)
 - ii. Product ox. nos.: Ca = +2 (charge on ion); F = -1 (fluorine always has a charge of -1)
 - iii. Ca is oxidized; F is reduced
- d. $2Fe_{(s)} + 3Cl_{2(g)} \rightarrow 2FeCl_{3(s)}$
 - i. Reactant ox. nos.: Fe = 0 (elemental state); Cl = 0 (elemental state)
 - ii. Product ox. nos.: Cl = -1 (charge on ion); Fe = +3 (there are three Cl atoms, each with an ox. no. of -1, and the numbers must add up to zero)
 - iii. Fe is oxidized; Cl is reduced



- i. Reactant ox. nos.: O in FeO = -2 (oxygen is always -2, unless in its peroxide form); Fe = +2 (numbers must add up to zero); O in O₂ = 0 (elemental state); H = +1 (hydrogen is always +1 when bonded to a non-metal); O in H₂O = -2 (oxygen is always -2 unless in its peroxide form; this also makes the numbers add up to zero)
- ii. Product ox. nos.: H = +1 (hydrogen is always +1 when bonded to a nonmetal); O = -2 (oxygen is always -2 unless in its peroxide form); Fe = +3 (there are three hydroxide ions- thus, there are three hydrogen ions each with an ox. no. of +1, so the total of the hydrogen ox. nos. is +3. There are also three oxygen atoms, each with an ox. no. of -2, so the total of the oxygen ox. nos. is -6. The oxygen ox. nos., plus the hydrogen ox. nos. is thus -3 (-6+3). The ox. nos. must add up to zero, so iron is +3)
- iii. Fe is oxidized; O from O₂ is reduced