

Reactions

Word	Definition
Coefficient	A number placed in front of a formula to balance a chemical reaction.
Decomposition	A redox reaction in which a compound breaks up to form two elements.
Double replacement	A solution reaction in which the positive ion of one compound combines with the negative ion of the other compound to form a precipitate, and the other ions remain dissolved in solution.
Law of conservation of charge	Charge cannot be created or destroyed by physical or chemical change. This is the basis for writing chemical formulas and half-reactions, and balancing redox ionic reactions.
Law of conservation of energy	Energy cannot be created or destroyed by physical or chemical change. This is the basis for calculating the heat of reaction.
Law of conservation of mass	Matter cannot be created or destroyed by physical or chemical change. This is the basis for balancing chemical reactions.
Mole ratio	The whole-number ratio between components of a balanced chemical reaction.
Precipitate	An insoluble solid that is formed either in a double-replacement reaction or as excess solute added to a saturated solution.
Product	The substances that are formed by a chemical reaction, designated as the right side of a chemical equation.
Reactant	The substances that are reacted together, designated as the left side of a chemical equation.
Redox reaction	A reaction in which one element is oxidized and another element is reduced.
Single replacement	A redox reaction in which an element replaces an ion in a compound.
Spectator ion	An ion that does not participate in the chemical reaction. In a redox reaction, it is the ion whose charge does not change. In a double replacement reaction, they are the ions that remain dissolved in solution.
Stoichiometry	The mathematics of mole relationships.
Synthesis	A redox reaction in which two elements combine to form a compound.

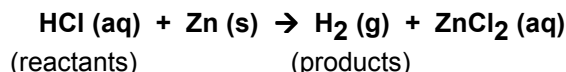
1) What is a Chemical Equation? (HW: p. 19, 20)

Objective: You will determine the reactant and product species in a reaction, balance chemical reactions, write and balance reactions given names and find the missing mass in reactions.

CHEMICAL EQUATION: symbolic representation of a chemical reaction. Includes the substances being reacted (reactants), the substances being formed (products), the phases of each of the substances, the number of moles of each substance, and the resultant energy change.

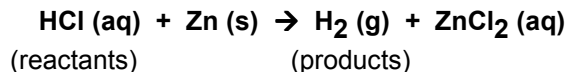
Reactants → Products

Coefficients are placed in front of the substance symbols to denote a mole ratio that is in accordance with the Law of Conservation of Mass.



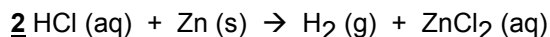
This says that hydrochloric acid reacts with zinc metal to form hydrogen gas and zinc chloride.

THE LAW OF CONSERVATION OF MASS – mass cannot be created or destroyed by physical or chemical change. The elements found in the reactants can be the only elements found in the products, and there must be equal numbers of moles of those elements on both sides.



This equation breaks the **Law of Conservation of Mass**, there are unequal moles of H and Cl on both sides.

Balancing equations involves placing coefficients that act as multipliers in front of a substance's formula.



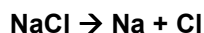
This shows that 2 moles of HCl are required to react with 1 mole of Zn. H₂ is formed because hydrogen exists in diatomic form (BrINClHOF).

RULES FOR WRITING CHEMICAL EQUATIONS (given the names)

- 1) Write the formulas of the compounds
- 2) Balance the equation.
 - a) Write in pencil
 - b) Write coefficients one element at a time
 - c) Only coefficients may be used...you may not change chemical formulas in order to balance.
 - d) Revise where necessary

SOMETHING TO REMEMBER - the difference between coefficients and subscripts (2 Cl vs Cl₂)

2 Cl means that there are **TWO ATOMS of chlorine**. **Cl₂** means that there is **one molecule of diatomic chlorine**. Diatomic molecules (Br₂, I₂, N₂, Cl₂, H₂, O₂, F₂) exist whenever these elements are not in a compound with another element. In NaCl, there is one Cl⁻¹ ion (since Na is charged +1), but if that chlorine is separated from that compound:



Then the Cl's thus formed will pair up diatomically, which throws off the balancing:

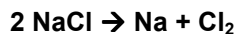


How do we balance this reaction? Look at the next page!

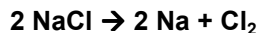


* There is one Na on the left and one Na on the right. Na is balanced...for now.

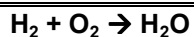
* There is one Cl on the left and one Cl on the right. Multiply them together: $1 \times 2 = 2$. This means that when balanced, there should be two chlorine atoms on each side. There are already two on the right, so put a 2 coefficient in front of NaCl:



* This messes up the balancing of Na, so place a 2 in front of the Na on the right side to balance this off:

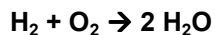


* There are now 2 Na's on both sides and 2 Cl's on both sides. This reaction is balanced.

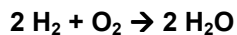


* There are two H's on the left and two H's on the right. H is balanced...for now.

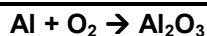
* There are two O's on the left and only one on the right. Multiply: $2 \times 1 = 2$. There should be two O's on either side to be balanced. Put a 2 in front of H₂O to accomplish this:



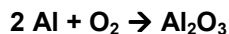
* This messes up the balancing of H. To remedy this, put a 2 in front of the H₂ on the left:



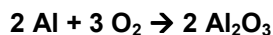
* There are now 4 H's on both sides and 2 O's on both sides. This reaction is balanced.



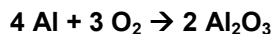
* There is one Al on the left and two on the right. Multiply $1 \times 2 = 2$. There should be 2 on each side, so put a 2 in front of Al on the left:



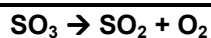
* There are two O's on the left and three on the right. Multiply $2 \times 3 = 6$. When balanced, there should be six O's on each side. On the left, there are two O's. Two times what equals six? Three! Put a 3 in front of the O₂ on the left. Three times what is six? Two! Put a 2 in front of the Al₂O₃ on the right. The reaction now looks like:



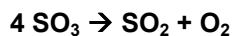
* But this now throws off the Al! There are now 2 Al on the left and four on the right! Easily fixed...Put a 4 in front of the Al on the left. Now it looks like:



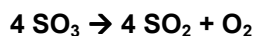
* There are now 4 Al's on each side and 6 O's on both sides. Bravo! It's balanced.



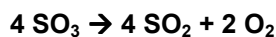
- * This one is strange, because you find O in two places on the left. No problem, follow!
- * There is one S on either side. S is balanced for now...but it won't stay that way for long.
- * There are three O's on the left, but FOUR on the right! Two in the SO_2 and two in the O_2 . Multiply: $3 \times 4 = 12$. There should be 12 O's on each side when balanced. There are three on the left. 3 times what equals 12? FOUR! Place a 4 in front of the SO_3 . The reaction now looks like



- * Now place a 4 in front of the SO_2 to balance off the S:



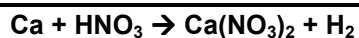
- * Now, there are 8 O's in the SO_2 and two in the O_2 . Remember, we need twelve O's on each side. What We need two more O's on the right side. How can we do that without messing up S again? Right! Put a 2 in front of the O_2 .



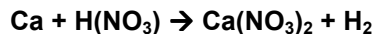
- * Now there are 4 S's on both sides and 12 O's. It's balanced! Note, however, that the coefficients can be simplified to



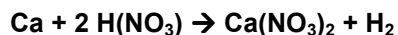
- * If a reaction's coefficients can be simplified in this way, you have to do it...otherwise, the Balancing Police will come and drag you away in handcuffs (well, covalent bonds, really...but you get the point).



- * Now we have a reaction with a polyatomic ion! Treat the whole ion like one element. To make it easier, put parentheses around the polyatomic ion if it doesn't already have them:

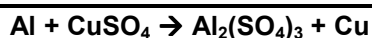


- * There is one Ca on either side. Ca is good. Will it stay that way? The suspense is killing me!
- * There is one H on the left and two on the right. $1 \times 2 = 2$, so there should be 2 H's on either side. Place a 2 in front of $\text{H}(\text{NO}_3)$:

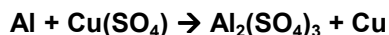


- * There are two NO_3 's on the left and also two on the right! No balancing of NO_3 is necessary

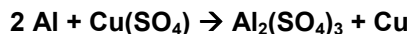
Let's see... 1 Ca on both sides, 2 H's on both sides and 2 NO_3 's on both sides? I'd say this sucker is balanced! ☺



* Again, put parentheses around any polyatomic ion that doesn't have them already:

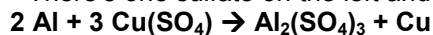


* There is one Al on the left, two on the right. $1 \times 2 = 2$, so there should be two Al's on both sides. Place a 2 in front of the Al on the left:

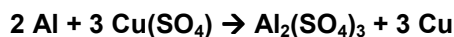


* There is one Cu on both sides. Leave it alone...for now.

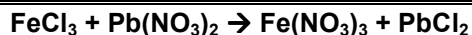
* There's one sulfate on the left and three on the right! $1 \times 3 = 3$, so place a 3 in front of the CuSO_4 on the left:



* Now there are three sulfates on both sides, but now the Cu is messed up. No biggie...there are 3 Cu's on the left, one on the right, so slap a 3 in front of the Cu on the right:



* There are now 2 Al's on each side, 3 Cu's on each side and 3 SO_4 's on each side. The reaction is now balanced!

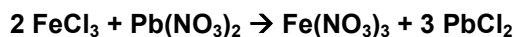


* This is the nastiest it will get for you. There are four things that have to be balanced. Be patient and do on at a time.

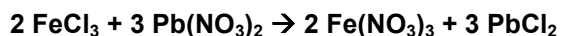
* There is one Fe on each side. Let it be for now.

* There's one Pb on each side. Let that be for now as well.

* There are three Cl's on the left and two on the right. $2 \times 3 = 6$, so there has to be six Cl's on each side. On the left, three times what equals six? Two! Put a 2 in front of FeCl_3 . On the right, two times what equals six? Three! Put a 3 in front of PbCl_2 . The reaction should now look like:



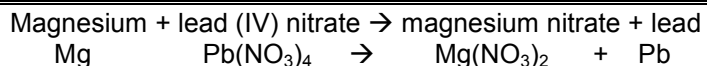
* Now let's deal with the nitrate. There are two NO_3 's on the left and three on the right. $2 \times 3 = 6$, so there should be six on each side. On the left, $2 \times 3 = 6$, so put a 3 in front of the $\text{Pb}(\text{NO}_3)_2$. On the right, $3 \times 2 = 6$, so put a 2 in front of the $\text{Fe}(\text{NO}_3)_3$. Now there are six nitrates on each side:



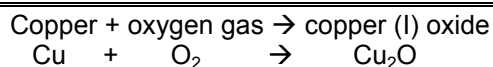
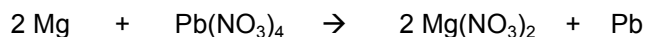
* Now check out what happened to Fe and Pb! There are now 2 Fe's on both sides and 3 Pb's on each side. Balancing the chloride and the nitrate automatically balanced the Fe and the Pb! Nice, how things work out sometimes. This baby be balanced!!!!

Writing Equations given the names

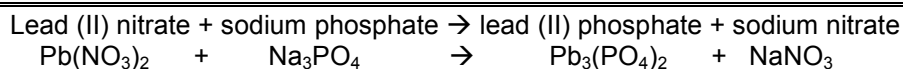
If you are given names of the compounds instead of the formulas, use the rules for writing formulas in the previous unit to write the formulas, and then balance the reaction.



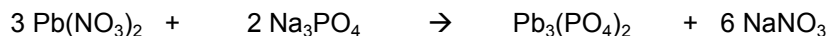
Then balance!



Then balance!



Then balance!



Missing Mass In Equations

The mass on the reactants side and products side have to equal each other because the Law of Conservation of Mass states that mass cannot be created or destroyed in a physical or chemical change.

If 35 grams of nitrogen gas are reacted with hydrogen gas to produce 47 grams of ammonia gas. How many grams of hydrogen gas were reacted?

$$35 + X = 47, \text{ so } X = \mathbf{12 \text{ grams of hydrogen gas}}$$

How many grams of aluminum are formed when 45 grams of aluminum oxide are decomposed into aluminum and 35 grams of oxygen?

$$45 = X + 35, \text{ so } X = \mathbf{10. \text{ grams of aluminum are formed}}$$

How many grams of iron (II) sulfide must be decomposed to form 33 grams of iron and 17 grams of sulfur?

$$X = 33 + 17, \text{ so } \mathbf{50. \text{ grams of iron (II) sulfide must be decomposed.}}$$

2) Oxidation and Reduction Reactions (HW: p. 21, 22)

Objective: You will identify oxidation and reduction reactions, complete reactions, identify the oxidized and reduced species in the reactions and identify spectator ions.

Driving Force: The “motivation” of a reaction to occur: In nature, changes that require the least amount of energy will be the ones that happen. After all, when you let go of a bowling ball, it falls down. The motivation is gravity. It would take more energy to make the ball go up than down, so the ball falls. In order to get the ball to go up, energy has to be added. This motivation is called a driving force.

Redox Reactions: driven by the loss of electrons (oxidation) and the gain of electrons (reduction).

In a redox reaction, one species gains electrons and one species loses them. Any ions that are not involved in this process are called “spectator ions”.

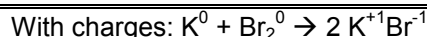
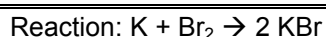
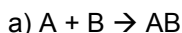
Redox (oxidation-reduction) - a reaction prompted by an exchange in electrons between two elements, resulting in a change of oxidation number of the two elements.

- a) The more electronegative element (usually a nonmetal atom) gains the electron, resulting in a decrease in oxidation number (going more negative). This is called **reduction**.
- b) The less electronegative element (usually a metal atom) loses the electron, resulting in an increase in oxidation number. This is called **oxidation**.

TYPES OF REDOX REACTIONS

1) Synthesis (synthesizer, synthetic, to make something from individual parts)

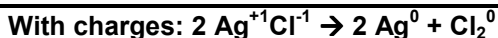
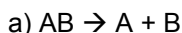
Two elements combine to form a compound. **This is the easiest way to make a compound, and is often used in industry for that purpose.**



K^0 goes from 0 to +1, so it is oxidized. K^0 gives electrons to Br^0 . Br^0 goes from 0 to -1, so it is reduced

2) Decomposition (reverse of synthesis), a compound decomposes into its original elements.

This reaction is very rare to occur on its own. It takes less energy for most compounds to form than to decompose. Water forms on its own from hydrogen and oxygen gas, but it takes a constant supply of energy to get water to decompose into hydrogen and oxygen. Imagine if the water in your glass suddenly decomposed into hydrogen and oxygen! Or if the salt on your plate decomposed suddenly into sodium (explosive metal) and chlorine (poisonous, corrosive gas)! Compounds exist because it requires less energy to exist in compound form. This is why the diatomic molecules exist...hydrogen has less energy as H_2 than as just H...so whenever H atoms are hanging around, they will form diatomic molecules. **This reaction is used to get highly reactive elements out of compounds and into a pure form. It is often done using electric current in a process called electrolytic decomposition.**

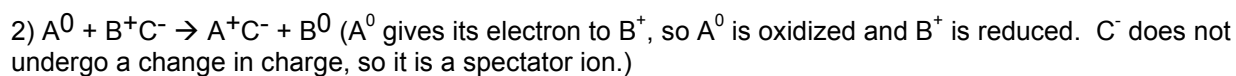
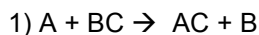


Ag^+ goes from +1 to 0, so it is reduced. It gains electrons from Cl^- , which goes from -1 to 0 and is oxidized.

3) Single Replacement: this reaction has one element and one compound on each side.

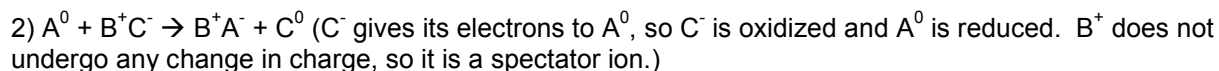
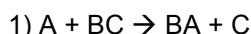
Two possibilities.

A) A metal plus a compound. The metal replaces the positive ion in the compound. **This reaction can be used to create electricity, and so is often found in batteries. This reaction can also be used to get less reactive elements out of compounds. React a more active metal with a compound containing the less active metal, and the more active metal will drive the less active metal out, leaving the less active metal in its pure form. This is also called extraction from ore.**



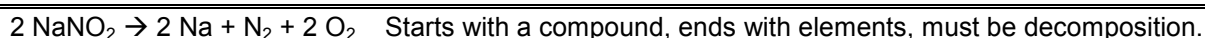
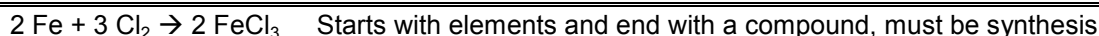
Zn^0 goes from 0 to +2, so it is oxidized. It loses electrons to Cu^{+2} , which goes from +2 to 0 and is reduced. NO_3^{-1} does not change charge, so it is the **spectator ion**.

B) A nonmetal plus a compound. The nonmetal replaces the negative ion in the compound. **This is a very rare reaction, and is a less economical way to produce a pure nonmetal than electrolytic decomposition..**



F_2^0 goes from 0 to -1, so it is reduced. It gains electrons from Cl^{-1} , which goes from -1 to 0 and is oxidized. Zn^{+2} does not change charge, so it is the **spectator ion**.

How Do You Identify The Type Of Reaction?



WHAT YOU HAVE TO BE ABLE TO DO:

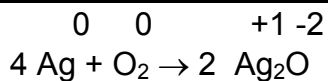
1) How Do You Determine The Charge Of Each Species?

a) Elements by themselves have no charge. Al^0 , Cl_2^0 , Fe^0

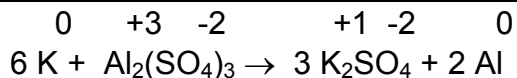
b) The charge of the ions in the compound can be looked up on Table E (for polyatomic ions) or the Periodic Table (for element ions). If the element has more than one charge listed, use the negative ion's charge to determine the charge of the positive ion.



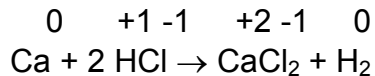
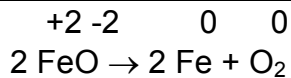
$Fe_3(PO_4)_2$: Fe has 2 charges listed (+2 and +3). PO_4 is listed on Table E as being -3. To make the charges add up to zero, Fe must have a +2 charge: $Fe_3^{+2}(PO_4)_2^{-3}$. For Fe: $3 \times +2 = +6$. For PO_4 : $2 \times -3 = -6$. +6 and -6 add up to ZERO.



Note the elements that are by themselves have a charge of 0



Note that the sum of the ion charges in each of the compounds is ZERO.

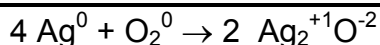


2) How Do You Determine Which Species Is Oxidized And Which Is Reduced And Which Is The Spectator Ion?

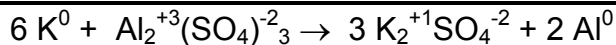
The species that is OXIDIZED has lost electrons, and its oxidation number has become more POSITIVE

The species that is REDUCED has gained electrons, and its oxidation number has become more NEGATIVE

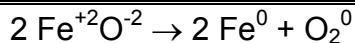
The species that is the SPECTATOR ION neither gains nor loses electrons, and its charge REMAINS THE SAME.



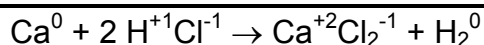
OX (charge more +): Ag^0 RD (charge more -): O_2^0 SI (no change): none



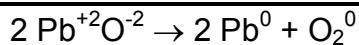
OX (charge more +): K^0 RD (charge more -): Al^{+3} SI (no change): SO_4^{-2}



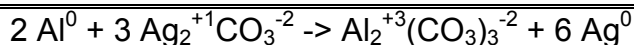
OX (charge more +): O^{-2} RD (charge more -): Fe^{+2} SI (no change): none



OX (charge more +): Ca^0 RD (charge more -): H^{+1} SI (no change): Cl^{-1}



OX (charge more +): O^{-2} RD (charge more -): Pb^{+2} SI (no change): none



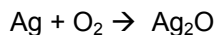
OX (charge more +): Al^0 RD (charge more -): Ag^{+1} SI (no change): CO_3^{-2}

3) How Do You Complete And Balance The Reaction?

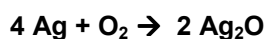
- 1) Identify the reaction type based on the given products or reactants.
- 2) Complete the rearrangement of elements, using the reaction type as your guide.
- 3) Write proper formulas for compounds, keep in mind that Br, I, N, Cl, H, O and F are diatomic when alone.
- 4) Balance the reaction.

Ag + O₂ → ?

- 1) You start with two elements. This must be a synthesis reaction.
- 2) The compound contains Ag and O. Look up the charges. Ag is +1 and O is -2.
- 3) Write the formula, Ag₂O

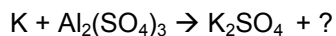


- 4) Balance one element at a time, like you did yesterday:

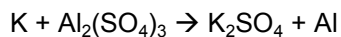


K + Al₂(SO₄)₃ → ?

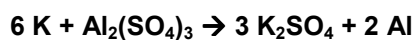
- 1) You start with an element and a compound. This must be a single replacement reaction.
- 2) K is a metal and will therefore replace the positive ion in the compound, which is Al. Your new compound will now contain K and SO₄. Look up their charges. K is +1 and SO₄ is -2.
- 3) Write the formula, K₂SO₄



- 4) This leaves the Al, which goes off on its own. The reaction is now

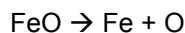


- 5) Which can now be balanced to form the complete equation:



FeO → ?

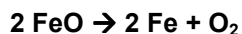
- 1) You start with just a compound. This must be a decomposition reaction.
- 2) The compound breaks up into its component elements, Fe and O.



- 3) Remember that O will form diatomic molecules when not bonded to another element.



- 4) Now balance the reaction:



3) Double Replacement (HW: p. 23, 24)

Objective: You will determine the relative solubility of ionic compounds in water, define and identify precipitates in double replacement reactions, complete reactions and identify the spectator ions.

Ionic compounds generally ionize in water. When the soluble ionic compound is placed into water, the molecule-ion interaction rips the ionic compound into its component ions. The hydrogen end of the water molecule ($\delta +$) attaches to the $-$ ion. The oxygen end of the water molecule ($\delta -$) attaches to the $+$ ion. The water molecules, which are in constant motion, tear the ions off of the crystal and keep them apart, floating forever in solution.

Insoluble ions remain together. This is because the attractions between the ions are too strong for water molecules to tear apart. The ions come together and form crystals, which make the solution cloudy. The crystals are pulled to the bottom of the solution by gravity, forming a **PRECIPITATE**.

Which ionic compounds dissociate (dissolve)? Refer to Reference Table F:

Ions That Form Soluble Compounds	Exceptions	Ions That Form Insoluble Compounds	Exceptions
Group 1 ions (Li^+ , Na^+ , etc.)		carbonate (CO_3^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
ammonium (NH_4^+)		chromate (CrO_4^{2-})	when combined with Group 1 ions, Ca^{2+} , Mg^{2+} , or ammonium (NH_4^+)
nitrate (NO_3^-)		phosphate (PO_4^{3-})	when combined with Group 1 ions or ammonium (NH_4^+)
acetate ($\text{C}_2\text{H}_3\text{O}_2^-$ or CH_3COO^-)		sulfide (S^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
hydrogen carbonate (HCO_3^-)		hydroxide (OH^-)	when combined with Group 1 ions, Ca^{2+} , Ba^{2+} , Sr^{2+} , or ammonium (NH_4^+)
chlorate (ClO_3^-)			
perchlorate (ClO_4^-)			
halides (Cl^- , Br^- , I^-)	when combined with Ag^+ , Pb^{2+} , and Hg_2^{2+}		
sulfates (SO_4^{2-})	when combined with Ag^+ , Ca^{2+} , Sr^{2+} , Ba^{2+} , and Pb^{2+}		

Table F
Solubility Guidelines for Aqueous Solutions

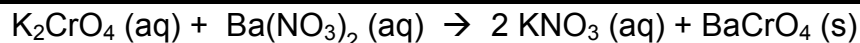
Compound	Soluble or Insoluble?	Why?
Li_2CO_3	Soluble	Li^+ is a group 1 ion, and Group 1 ions are always soluble, with no exceptions.
$\text{Pb}(\text{NO}_3)_2$	Soluble	NO_3^- is listed as being always soluble, no exceptions.
ZnCl_2	Soluble	Cl^- is a halide, which are listed as being soluble, with a few exceptions. However, Zn^{2+} is not one of those exceptions.
BaSO_4	Soluble	SO_4^{2-} is listed as being soluble, with a few exceptions. However, Ba^{2+} is not one of those exceptions.
CuCO_3	Inoluble	CO_3^{2-} is listed as being insoluble, with a few exceptions. However, Cu^{2+} is not one of those exceptions, because it is not a Group 1 ion.
$\text{Pb}(\text{CrO}_4)_2$	Inoluble	CrO_4^{2-} is listed as being insoluble, with a few exceptions. However, Pb^{2+} is not one of those exceptions, because it is not a Group 1 ion.
Na_3PO_4	Soluble	Na^+ is a group 1 ion, and Group 1 ions are always soluble, with no exceptions.
NH_4OH	Soluble	NH_4^+ is listed as being always soluble, no exceptions.
$\text{Mg}(\text{OH})_2$	Inoluble	OH^- is listed as being insoluble, with a few exceptions. However, Mg^{2+} is not one of those exceptions, because it is not a Group 1 ion.

Double Replacement reaction

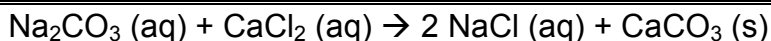
The positive ion of one compound swaps places with the negative ion from the other compound. This reaction is carried out with aqueous solutions of each reactant so that the ions can more easily mix and react.



Solution AB (which is made of separated ions A^+ and B^-) is mixed into solution CD (which is made of separated ions C^+ and D^-). Instantly upon mixing, the A^+ ions of the first solution seek out and bond very tightly to the D^- ions of the second solution, so tightly that water molecules cannot separate them. They form small crystals of compound AD, which turn the solution cloudy. Finally, gravity pulls the crystals of AD down and they settle to the bottom as a **precipitate**.



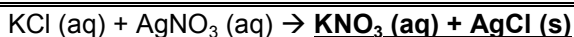
Taking a look at the two products on Table F: K^{+1} is a Group 1 ion, so it is always soluble, no exceptions. NO_3^{-1} is also always soluble, no exceptions. On the other hand, CrO_4^{-2} is listed as being insoluble with exceptions, but Ba^{+2} is not one of those exceptions. BaCrO_4 is the precipitate in this reaction.



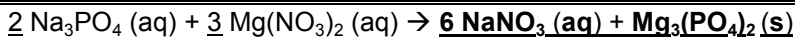
Taking a look at the two products on Table F: Na^{+1} is a Group 1 ion, so it is always soluble, no exceptions. Cl^{-1} is usually soluble, and Na^{+1} is not an exception. On the other hand, CO_3^{-2} is listed as being insoluble with exceptions, but Ca^{+2} is not one of those exceptions. CaCO_3 is the precipitate in this reaction.

How Can Double Replacement Reactions be Completed?

The same way redox reactions are completed! For the missing compounds, write the positive ion first and the negative ion second. Write the formulas, then balance the reaction. Then use Table F to identify the precipitate!

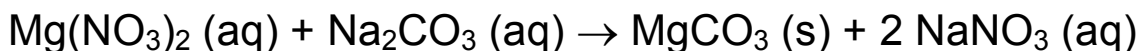


K^{+1} and NO_3^{-1} bond, Ag^{+1} and Cl^{-1} bond. KNO_3 is soluble (aq), and the AgCl is the insoluble precipitate (s).



Na^{+1} and NO_3^{-1} bond, Mg^{+2} and PO_4^{-3} bond. NaNO_3 is soluble and $\text{Mg}_3(\text{PO}_4)_2$ is the insoluble precipitate.

The spectator ions in a double replacement reaction are the two that remain dissolved in water on both sides of the reaction. In the following double replacement reaction, identify the spectator ions:



Mg^{+2} : is found in the precipitate

NO_3^{-1} : is in (aq) compounds on both sides

Na^{+1} : is in (aq) compounds on both sides

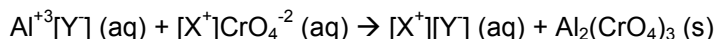
CO_3^{-2} : is found in the precipitate

Therefore, NO_3^{-1} and Na^{+1} are the spectator ions for this reaction. They are not involved in the formation of the precipitate.

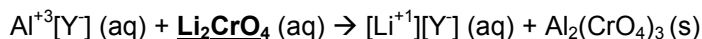
How can we create a double replacement reaction that will make a desired precipitate?

Let's say that our chemical process requires the insoluble compound aluminum chromate, $\text{Al}_2(\text{CrO}_4)_3$. How can we design a double replacement reaction that will produce this precipitate?

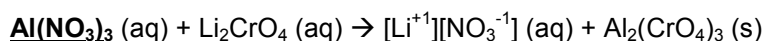
1) We need two soluble compounds, one of which contains the Al^{+3} ion and the other the CrO_4^{-2} ion. We can choose any soluble ions, here called X^+ and Y^- to make those two compounds.



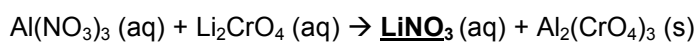
2) Looking at table F, we can notice that Group 1 ions make nice, soluble + ions. We can use Li^{+1} , Na^{+1} or K^{+1} for our X^+ . Hey! Ammonium would also work for that. Let's select...hmmm...ahhh... Li^{+1} . We could have chosen any of the others, though, with the same results. Let's write the formula for Li^{+1} and CrO_4^{-2} :



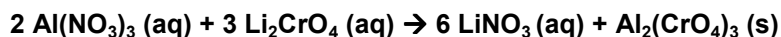
3) Now, we need a nice, soluble, negative ion to pair up with Al^{+3} . What's this I see on Table F? Nitrate, acetate, chlorate and perchlorate are always soluble? Let's pick one of those! I have a special preference for the nitrate ion...it is easily available and is deliciously soluble. Any of the others would have been just as good and given us the same results. Let's write the formula of Al^{+3} and NO_3^{-1} :



4) Now, let's write the formula of the compound between Li^{+1} and NO_3^{-1} :



5) All that's left to do is balance the reaction, and we're done:



Now, go out and buy some aluminum nitrate and lithium chromate, dissolve them in water and mix them, and you are in business! How much aluminum nitrate and lithium chromate? Depends on how much aluminum chromate you need. We'll do that tomorrow!

4) Stoichiometry of Equations (HW: p. 25-27)

Objective: You will calculate moles of a product or reactant, given moles of any substance in a reaction, using stoichiometric ratios from the balanced equation.

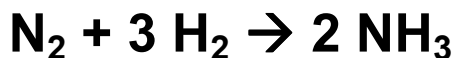
It's time to make pancakes! The recipe calls for 1 egg, 1 cup of mix, 1 cup of milk, 2 tablespoons of oil, 1 ½ tablespoons of vanilla extract and 1/3 teaspoon of nutmeg. This will make sixteen 4" diameter pancakes. This will serve four people quite nicely, and make for a delicious breakfast. I recommend real maple syrup, as it is delicious and actually has some nutritional value.

But wait! Your grandparents and an aunt and uncle are coming to visit! You now have 8 people coming to breakfast, and the recipe only gives you instructions for four servings!!! Whatever shall you do? Easy! Use RATIOS to double the recipe and make enough to satisfy everyone's morning hunger!

Original recipe (serves four, makes 16 pancakes)	Doubled recipe (serves 8, makes 32 pancakes)
1 egg	2 eggs
1 cup mix	2 cups of mix
1 cup milk	2 cups of milk
2 tablespoons of oil	4 tablespoons of oil
1 ½ tablespoon of vanilla extract	3 tablespoons vanilla extract
1/3 teaspoon of nutmeg	2/3 teaspoon of nutmeg

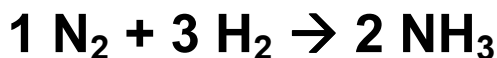
Notice how the proportions of the ingredients AND the pancakes are preserved? Eat hearty!

This is also true of chemical reactions. The coefficients in front of each species lets you know what the proportional number of moles of reactant needed is to make a proportional number of moles of product. For example, in the manufacture of ammonia (called the Haber Process), a simple synthesis reaction is used:



What this means is that if nitrogen and hydrogen are reacted in a 1:3 mole ratio, the amount of ammonia you will produce from the reaction is 2 moles. This is important to know, because if your company gets a call for 10 moles of ammonia, you can use that 1:3 → 2 proportion to determine how many moles of nitrogen and hydrogen you have to react together in order to make the 10 moles of ammonia. Let's see how this scales up:

5 moles 15 moles 10 moles are needed!



If you scale the ratio up by a factor of five, you will see that to make 10 moles of ammonia, you will have to combine 5 moles of nitrogen with 15 moles of hydrogen. This maintains the 1:3 → 2 ratio that the coefficients give you!

What if you needed 34.338 moles of NH₃? Is there an easy way to make use of the 1:3 → 2 ratio to figure that out? In fact, there is a very simple equation you can use:

How many moles of X are formed when n moles of Y are reacted?

$$\text{Moles of given} \quad \times \quad \frac{\text{Coefficient of target}}{\text{Coefficient of given}} = \text{Moles of target}$$

Notice that for each of the following problems, the given is written in ~~strikethrough~~ mode, indicating that this unit cancels out to leave the target.

For the reaction $\text{N}_2 (\text{g}) + 3 \text{H}_2 (\text{g}) \rightarrow 2 \text{NH}_3 (\text{g})$, how many moles of NH_3 will be formed if 6.0 moles of N_2 are completely reacted with H_2 ?

$$\text{Moles of given} \quad \times \quad \frac{\text{Coefficient of target}}{\text{Coefficient of given}} = \text{Moles of target}$$

$$6.0 \text{ moles } \text{N}_2 \times (2 \text{ NH}_3 / 1 \text{ N}_2) = \mathbf{12 \text{ moles of N}_2}$$

For the reaction $\text{N}_2 (\text{g}) + 3 \text{H}_2 (\text{g}) \rightarrow 2 \text{NH}_3 (\text{g})$, how many moles of H_2 are needed to completely react with N_2 to form 1000. moles of NH_3 ?

$$\text{Moles of given} \quad \times \quad \frac{\text{Coefficient of target}}{\text{Coefficient of given}} = \text{Moles of target}$$

$$1000. \text{ moles } \text{NH}_3 \times (3 \text{ H}_2 / 2 \text{ NH}_3) = \mathbf{1500. \text{ moles H}_2}$$

For the reaction $2 \text{Na} + 2 \text{H}_2\text{O} \rightarrow 2 \text{NaOH} + \text{H}_2$, how many moles of Na are needed to make 4.0 moles of H_2 ?

$$\text{Moles of given} \quad \times \quad \frac{\text{Coefficient of target}}{\text{Coefficient of given}} = \text{Moles of target}$$

$$4.0 \text{ moles } \text{H}_2 \times (2 \text{ Na} / 1 \text{ H}_2) = \mathbf{8.0 \text{ moles Na}}$$

For the reaction $4 \text{Al} + 3 \text{O}_2 \rightarrow 2 \text{Al}_2\text{O}_3$, how many moles of Al_2O_3 will form if 6.0 moles of Al are completely reacted with O_2 ?

$$\text{Moles of given} \quad \times \quad \frac{\text{Coefficient of target}}{\text{Coefficient of given}} = \text{Moles of target}$$

$$6.0 \text{ moles } \text{Al} \times (2 \text{ Al}_2\text{O}_3 / 4 \text{ Al}) = \mathbf{3.0 \text{ moles Al}_2\text{O}_3}$$

For the reaction $2 \text{NaCl} (\text{aq}) + \text{Pb}(\text{NO}_3)_2 (\text{aq}) \rightarrow 2 \text{NaNO}_3 (\text{aq}) + \text{PbCl}_2 (\text{s})$, how many moles of PbCl_2 precipitate will form when 5.0 moles of $\text{Pb}(\text{NO}_3)_2$ are completely reacted with NaCl ?

$$\text{Moles of given} \quad \times \quad \frac{\text{Coefficient of target}}{\text{Coefficient of given}} = \text{Moles of target}$$

$$5.0 \text{ moles } \text{Pb}(\text{NO}_3)_2 \times (1 \text{ PbCl}_2 / 1 \text{ Pb}(\text{NO}_3)_2) = \mathbf{5.0 \text{ moles PbCl}_2}$$

Advanced Stoichiometry:

In the last unit, you learned how to convert from grams, liters and number of particles to moles and back again. In this unit, we will add an additional feature, the mole relationship in a balanced equation. If you know how much product you will need, you will be able to calculate the amount of reactant needed. You will also be able to determine the theoretical amount of product that should form when a specific quantity of reactant is used.

If you called a company and told them you need 1000. moles of NH_3 , they would be quite annoyed with you. They make things by the gram or liter, their materials cost dollars per gram or liter, so they would like to know how many grams or liters you need.

This is easy to do. The following steps should be taken:

- 1) Identify the given and the target
- 2) Convert the given to moles by dividing by the appropriate conversion factor
- 3) Multiply the moles of given by the mole ratio in the balanced chemical equation (moles target/moles given). This will cancel out the given and leave you with the target
- 4) Convert moles of target to the relevant unit by multiplying by the appropriate conversion factor.

Given the reaction $\text{Zn (s)} + 2 \text{HCl (aq)} \rightarrow \text{ZnCl}_2 \text{(aq)} + \text{H}_2 \text{(g)}$, How many grams of Zn are needed to produce 20.0 grams of $\text{ZnCl}_2 \text{(g)}$?

- 1) Identify the given and the target

Given: 20.0 g ZnCl_2 Target: ? g of Zn

- 2) Convert the given to moles by dividing by the appropriate conversion factor

$$(20.0 \text{ g ZnCl}_2) / (136 \text{ g/mol}) = 0.15 \text{ moles ZnCl}_2$$

- 3) Multiply the moles of given by the mole ratio in the balanced chemical equation (moles target/moles given). This will cancel out the given and leave you with the target

$$0.15 \text{ moles ZnCl}_2 \times (1 \text{ Zn} / 1 \text{ ZnCl}_2) = 0.15 \text{ moles Zn}$$

- 4) Convert moles of target to the relevant unit by multiplying by the appropriate conversion factor.

$$(0.15 \text{ moles Zn}) \times (65.4 \text{ g/mole}) = \underline{9.8 \text{ grams Zn.}}$$

Given the reaction $\text{Zn (s)} + 2 \text{HCl (aq)} \rightarrow \text{ZnCl}_2 \text{(aq)} + \text{H}_2 \text{(g)}$, How many grams of HCl are needed to produce 20.0 liters of $\text{H}_2 \text{(g)}$?

- 1) Identify the given and the target

Given: 20.0 L H_2 Target: ? g of HCl

- 2) Convert the given to moles by dividing by the appropriate conversion factor

$$(20.0 \text{ L H}_2) / (22.4 \text{ L/mole}) = 0.893 \text{ moles H}_2$$

- 3) Multiply the moles of given by the mole ratio in the balanced chemical equation (moles target/moles given). This will cancel out the given and leave you with the target

$$0.893 \text{ moles H}_2 \times (2 \text{ HCl} / 1 \text{ H}_2) = 1.79 \text{ moles HCl}$$

- 4) Convert moles of target to the relevant unit by multiplying by the appropriate conversion factor.

$$(1.79 \text{ moles HCl}) \times (36.5 \text{ g/mole}) = \underline{65.3 \text{ grams HCl}}$$

Student Name: _____ Grades: _____, _____, _____, _____

1) What is a Chemical Equation? Homework

A) Balance the following equations/reactions by placing small whole-number coefficients in front of the formulas. You may NOT change the formula of a compound. A number one (1) does not need to be written, but is helpful to do.

Reaction (fill in the coefficients)	Sum of Coefficients
_____ C(s) + _____ H ₂ (g) → _____ CH ₄	
_____ Fe(s) + _____ O ₂ (g) → _____ Fe ₂ O ₃	
_____ NaI(s) → _____ Na(s) + _____ I ₂ (s)	
_____ C ₆ H ₁₂ O ₆ (s) → _____ C(s) + _____ H ₂ O(l)	
_____ AgNO ₃ (aq) + _____ Cu(s) → _____ Ag(s) + _____ Cu(NO ₃) ₂ (aq)	
_____ Na ₂ CO ₃ (aq) + _____ HCl(aq) → _____ NaCl(aq) + _____ H ₂ O(l) + _____ CO ₂ (g)	
_____ H ₂ (g) + _____ Cl ₂ (g) → _____ HCl(g)	
_____ N ₂ (g) + _____ O ₂ (g) → _____ N ₂ O ₄ (g)	
_____ CH ₄ (g) + _____ O ₂ (g) → _____ CO ₂ (g) + _____ H ₂ O(g)	
_____ N ₂ (g) + _____ H ₂ (g) → _____ NH ₃ (g)	
_____ H ₂ O ₂ (l) → _____ H ₂ O(l) + _____ O ₂ (g)	
_____ Al ₂ O ₃ → _____ Al(s) + _____ O ₂ (g)	
_____ C(g) + _____ O ₂ (g) → _____ CO ₂ (g)	
_____ CuO(s) + _____ C(s) → _____ Cu(s) + _____ CO ₂ (g)	
_____ Ca(OH) ₂ (aq) + _____ HCl(aq) → _____ CaCl ₂ (aq) + _____ H ₂ O(l)	

B) What does the following mean in an equation?

(s) _____ (l) _____ (g) _____ (aq) _____

C) Write the formulas for the following compounds:

Name	Formula	Name	Formula
sodium oxide		lead (II) carbonate	
potassium sulfate		lead (IV) carbonate	
iron (II) chloride		zinc phosphate	
iron (III) chloride		zinc phosphide	

D) Write the formulas given the names below and balance the reactions:

1) iron + nitrogen gas \rightarrow iron (II) nitride

2) lead + copper (II) nitrate \rightarrow lead (II) nitrate + copper

3) calcium + dihydrogen monoxide \rightarrow calcium hydroxide + hydrogen gas

4) potassium phosphate + iron (II) nitrate \rightarrow potassium nitrate + iron (II) phosphate

E) Find the missing mass!

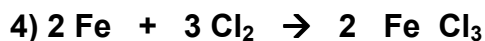
1) If 15 grams of nitrogen are reacted with 21 grams of hydrogen gas to form ammonia, how many grams of ammonia will be formed?

2) If 37 grams of zinc react with hydrochloric acid (HCl) to form 26 grams of zinc chloride and 16 grams of hydrogen gas, how many grams of HCl were reacted?

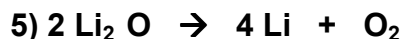
3) How many grams of aluminum oxide must decompose to form 112 grams of aluminum metal and 18 grams of oxygen gas?



OXIDIZED: _____ Reduced: _____ Spectator Ion: _____

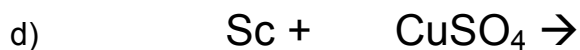


OXIDIZED: _____ Reduced: _____ Spectator Ion: _____

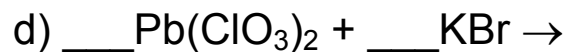
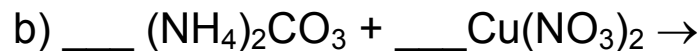
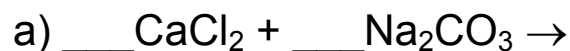


OXIDIZED: _____ Reduced: _____ Spectator Ion: _____

D) Complete the following reactions by writing the appropriate formula(s) and balancing:



D) Complete and balance the following double-replacement reactions, then identify the precipitate with (s):



E) Create a double replacement reaction that will form $\text{Zn}(\text{OH})_2$ as a precipitate. Show all your work, step by step.

5) Explain why ion exchange reactions are not considered to be redox reactions.

6) Explain the difference between spectator ions in redox reactions and spectator ions in ion exchange reactions.

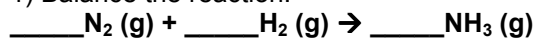
What is a spectator ion in a redox reaction?	
What is a spectator ion in a double replacement reaction?	

4) Stoichiometry of Equations Homework

For ALL mole conversion problems, show ALL of your work, including showing which units cancel out by putting a slash through them.

A) For the reaction $\text{N}_2 (\text{g}) + \text{H}_2 (\text{g}) \rightarrow \text{NH}_3 (\text{g})$:

1) Balance the reaction:



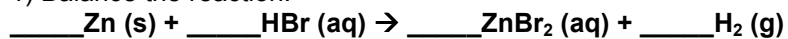
2) How many moles of N_2 are needed to make 5.0 moles of NH_3 ?

3) How many moles of N_2 are needed to completely react with 10.0 moles of H_2 ?

4) How many moles of NH_3 should form if 6.0 moles of H_2 are completely reacted with N_2 ?

B) For the reaction $\text{Zn} (\text{s}) + \text{HBr} (\text{aq}) \rightarrow \text{ZnBr}_2 (\text{aq}) + \text{H}_2 (\text{g})$

1) Balance the reaction:



2) How many moles of Zn are needed to make 8.0 moles of ZnBr_2 ?

3) How many moles of HBr are needed to make 4.0 moles of H_2 ?

4) How many moles of ZnBr_2 should form if 5.0 moles of HBr are completely reacted with Zn ?

C) Using your knowledge of stoichiometric relationships and the factor-label method, find the solutions to the following chemical applications problems:

1) The space industry uses liquid oxygen as a fuel for its rocket boosters. How many grams of H_2O must be decomposed to make 32000. grams of O_2 fuel? ($2 \text{H}_2\text{O} \rightarrow 2 \text{H}_2 + \text{O}_2$)

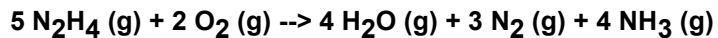
2) Calcium carbide, CaC_2 , is used in the production of acetylene gas, C_2H_2 . This gas is used in welding applications. How many grams of CaC_2 must be reacted to form 448 L of acetylene?



3) Mercury is essential to the manufacture of thermometers, barometers and mercury switches used in thermostats. It can be obtained from cinnabar ore (HgS). How many grams of HgS must be decomposed to produce 100.5 grams of mercury (a liquid metal)? ($\text{HgS} (\text{s}) \rightarrow \text{Hg} (\text{l}) + \text{S} (\text{s})$)

4) Fermentation is a process whereby an sugar is converted to carbon dioxide and an alcohol by the anaerobic respiration of yeast. When making wine, this process takes place in a sealed bottle, trapping the CO_2 (g) along with the alcohol. When the bottle is opened, how many liters of CO_2 (g) will be released if 90.0 grams of alcohol ($\text{C}_2\text{H}_6\text{O}$) were produced? ($\text{C}_6\text{H}_{12}\text{O}_6 (\text{aq}) \rightarrow 2 \text{CO}_2 (\text{g}) + 2 \text{C}_2\text{H}_6\text{O} (\text{l})$)

5) Hydrazine (N_2H_4) is injected into the steam flow in a fossil-fuel plant. Its purpose is to remove dissolved oxygen from the steam, so that the turbine blades don't react with the oxygen. Given the equation



Calculate how many L of N_2H_4 (g) must be injected into the steam tube to react with 320. g of O_2 .

6) Calcium hydroxide ($\text{Ca}(\text{OH})_2$) is used to remove excess fluoride from waste water in sewage treatment plants. Given the equation



Calculate how many grams of $\text{Ca}(\text{OH})_2$ are needed to completely react with 400. grams of HF in order to eliminate it from the wastewater.