Chemistry

Chemical Reactions
Lesson 7
Lesson Plan
David V. Fansler

Describing Chemical Change

Objectives: Write equations describing chemical reactions using appropriate symbols; Write balanced chemical equations when given the names or formulas of the reactants and products in a chemical reaction

- Word Equations
  - Every minute of everyday, there are chemical reactions going on all around you – name some
    - Digestion of food – acids breaking down food to a form the body can use for energy
    - Photosynthesis – converting sunlight and CO_2 to O_2 and food
    - Respiration – 0_2 being taken in, CO_2 and water vapor expelled
    - Battery in your car to create electricity
    - Gasoline being used to create motion
  - Not all chemical reactions are good – name some
    - Rust on your car (iron and oxygen)
    - Food molding – food broken down by mold
  - In all cases, you have chemicals reactions, which involve one or more substances (reactants) which are changed into one or more new substances (products)
  - Your task is to learn to describe a chemical reaction in writing.
  - The basic form is Reactants $\rightarrow$ Products
  - Important to note that atoms are neither created nor destroyed in a chemical reaction – they are merely rearranged. – Remember the Law of Conservation of Mass?
  - In describing a chemical reaction you use statements such as
    - Iron reacts with oxygen to produce iron(III) oxide which is called rust
    - Or it could be written iron + oxygen $\rightarrow$ iron(III) oxide
      - Note that the reactants are on the left of the arrow and the product is on the right
      - A plus sign separates the reactants – if there had been more than two reactants, there would have been a + sign between each of them.
      - Had there been more than one product, they also would have been separated by + signs
    - Consider hydrogen peroxide being poured into a cut
      - If you have ever done this, you know that the hydrogen peroxide bubbles – this is oxygen being formed. The reaction would be
        - Hydrogen peroxide $\rightarrow$ water + oxygen

- Another example common to our live is the burning of natural gas for heating – methane is the major component and would give a chemical reaction of
  - Methane + oxygen $\longrightarrow$ carbon dioxide + water
  - Notice that we included oxygen since it is needed as part of the reaction
- Some word equations indicate outside influences that are needed – such as:

$$copper\ \ carbon\ dioxide \xrightarrow{water} copper\ carbonate$$

- $glu\cos e \xrightarrow{yeast} ethanol + carbon\ dioxide$

$$carbon\ dioxide + water \xrightarrow{light} glu\cos e + oxygen$$

  - In these examples we have the green that you see on copper left outside (domes, Statue of Liberty), production of ethanol (which is an alcohol) and photosynthesis

- Chemical Equations
  - To be more precise, we use chemical equations to describe a chemical reaction. As with a word equation, we use an arrow to indicate a reaction, with the reactants on the left, separated by a + sign, and products on the right, also separated by a + sign. Take rust:
    - $Fe + O_2 \rightarrow Fe_2O_3$
    - Such an equation is called a skeleton equation, in that it only shows the products and reactants, but does not show the amount used or created in the reaction.
    - A skeleton equation is the important first step in obtaining a correct chemical equation
  - The second thing we can add is the physical state of the reactants and products – to do this we use abbreviations in parenthesis following the element or compound
    - (s) for solid, (l) for liquid, (g) for gas and (aq) for an aqueous solution (a substance dissolved in water). For rust:
      - $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$
  - In many chemical reactions, a substance is used that will speed up the reaction, but is not used up in the reaction – this substance is called a **catalyst.** Because a catalyst is not used up in a reaction, it is written in over the arrow.
    - Remember that hydrogen peroxide on a cut bubbles forming water and oxygen. Adding manganese(IV) dioxide (MnO_2(s)) catalyzes the decomposition of an aqueous solution of hydrogen peroxide (H_2O_2)(aq) into water and oxygen, we get the skeleton equation of
      - $H_2O_2(aq) \xrightarrow{MnO_2} H_2O(l) + O_2(g)$
- Symbols used in Chemical Equations

Symbols Used in Chemical Equations	
Symbol	Explanation
+	Used to separate two reactants or two products
$\longrightarrow$	"Yields" separates reactants and products
$=$	An alternative to $\longrightarrow$
(s)	Designates a reactant or product in the solid state; placed after the formula
(l)	Designates a reactant or product in the liquid state; placed after the formula
(g)	Designates a reactant or product in the gaseous state; placed after the formula
(aq)	Designates an aqueous solution; the substance is dissolved in water; placed after the formula
$\xrightarrow{\Delta}$ $\xrightarrow{heat}$	Indicates that heat is supplied to the reaction
$\xrightarrow{Pt}$	A formula written above or below the yield sign indicates its use as a catalyst (in this example, platinum)

Sample Problem

Write a skeleton equation for this chemical reaction: Solid sodium, hydrogen carbonate reacts with hydrochloric acid to produce aqueous sodium chloride, water and carbon dioxide gas. Include the appropriate symbols.

First, write the formula for each substance in the reaction, separate the reactants from the products, and indicate the state of each substance

Reactants:

Solid sodium hydrogen carbonate: $NaHCO_3$(s)

Hydrochloric acid: HCl(aq)

Products:

Aqueous sodium chloride: NaCl(aq)

Water: H_2O(l)

Carbon dioxide gas: CO_2(g)

Then write the equation

$$NaHCO_3(s) + HCl(aq) \to NaCl(aq) + H_2O(l) + CO_2(g)$$

Sulfur burns in oxygen to form sulfur dioxide
$S(s) + O_2(g) \rightarrow SO_2(g)$

Heating potassium chlorate in the presence of the catalyst manganese dioxide produces oxygen gas. Potassium chloride is left as a solid.
$KHClO_3(s) \xrightarrow{\Delta,\, MnO_2} KCl(s) + O_2(g)$

Write a sentence that describes each chemical reaction:
$KOH(aq) + H_2SO_4(aq) \rightarrow H_2O(l) + K_2SO_4(aq)$
Aqueous potassium hydroxide is combined with aqueous sulfuric acid to form water and aqueous potassium sulfate

$Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$
Solid sodium is combined with water to form sodium hydroxide and hydrogen gas

- Balancing an equation
  - So far all we have done in writing a skeleton equation is to identify the elements or compounds that are the reactants and products, and what state they are in (solid, liquid, gas or aqueous)
  - In order to maintain the Law of Conservation of Mass, we must have the same number of atoms of each element on both sides of the equation
  - Sometimes this works out by itself
    - $C(s) + O_2(g) \rightarrow CO_2(g)$
    - Notice that we have 1 C on both sides and 2 oxygens, so the equation is balanced
  - In another equation we have
    - $H_2(g) + O_2(g) \rightarrow H_2O(l)$
    - In this case we have 2 hydrogen on both sides, but one oxygen on the reactant side and 2 oxygens on the product side – this is not a balanced equation, how do we balance it?
    - $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
    - Now it is balanced because we have 4 hydrogen atoms on each side and two oxygen atoms on each side
  - Rules for balancing an equation
    - Determine the correct formula for all the reactants and products in the reaction, using what you think you know about elements and compounds. In some cases, you may also indicate in parentheses the state in which the reactants and products exist.
    - Write the formulas for the reactants on the left side and the formulas for the products on the right side with a yields sign ($\rightarrow$) in between. If two or more reactants or products are involved, separate their formulas with plus signs. When finished, you will have a skeleton equation.

- Count the number of atoms of each element in the reactants and products. For simplicity, a polyatomic ion appearing unchanged on both sides of the equation is counted as a single unit.
- Balance the elements one at the time by using coefficients. When no coefficient is written, it assumed to be 1. It is best to begin the balancing operation with elements that appear only once on each side of the equation. You must not attempt to balance an equation by changing the subscripts in the chemical formula of a substance.
- Check each atom or polyatomic ion to be sure that the equation is balanced.
- Finally, make sure all the coefficients are in the lowest possible ratio that balances

Sample Problems

Hydrogen and oxygen react to form water. Write a balanced equation for this reaction.

First, write out the skeleton equation

$H_2(g) + O_2(g) \longrightarrow H_2O(l)$

Next, counting the number of atoms of each element on both sides we see that we have 2 hydrogen on both sides, but two oxygens in the reactants and only one in the products – so if we multiple the product by 2, it will give us two oxygens on both sides

$H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$

But now we have 2 hydrogens in the reactants and 4 in the product, so it we use a coefficient of 2 on the reactant hydrogen

$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$

we now have 4 hydrogens on both sides, and 2 oxygens on both sides – the equation is balanced.

$$AgNO_3 + H_2S \to Ag_2S + HNO_3$$

balances to

$$2AgNO_3 + H_2S \to Ag_2S + 2HNO_3$$

$$MnO_2 + HCl \to MnCl_2 + H_2O + Cl_2$$

balances to

$$MnO_2 + 4HCl \to MnCl_2 + 2H_2O + Cl_2$$

$Zn(OH)_2 + H_3PO_4 \rightarrow Zn_3(PO_4)_2 + H_2O$

balances to

$3Zn(OH)_2 + 2H_3PO_4 \rightarrow Zn_3(PO_4)_2 + 6H_2O$

Write the balanced chemical equation for the reaction of copper metal and an aqueous solution of silver nitrate.

$AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$

balances to

$2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

$CO + Fe_2O_3 \rightarrow Fe + CO_2$

balances to

$3CO + Fe_2O_3 \rightarrow 2Fe + 3CO_2$

Write the balanced chemical equation for the reaction of carbon with oxygen to from carbon monoxide

$C + O_2 \rightarrow CO$

balances to

$2C + O_2 \rightarrow 2CO$

Aluminum reacts with oxygen in the air to form a thing protective coat of aluminum oxide. Balance the equation for this reaction

$Al(s) + O_2(g) \rightarrow Al_2O_3(s)$

balances to

$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

$FeCl_3 + NaOH \rightarrow Fe(OH)_3 + NaCl$

balances to

$FeCl_3 + 3NaOH \rightarrow Fe(OH)_3 + 3NaCl$

$CH_4 + Br_2 \rightarrow CH_3Br + HBr$

balances to

$CH_4 + Br_2 \rightarrow CH_3Br + HBr$

$sodium + water \rightarrow sodium\ hydroxide + hydrogen$

$Na + H_2O \rightarrow NaOH + H_2$

balances to

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

$calcium\ hydroxide + sulfuric\ acid \rightarrow calcium\ sulfate + water$

$Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + H_2O$

balances to

$Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + 2H_2O$

Types of Chemical Reactions

Objectives: Identify a reaction as combination, decomposition, single replacement, double replacement, and combustion reactions; Predict the products of combination, decomposition, single replacement, double replacement, and combustion reactions.

- Classifying Reactions
  - If we consider all the different possible chemical reactions, we would quickly conclude that there are millions of them possible, with millions of possible compounds
  - Just as we learned how to identify compounds (molecular and ionic) there are several ways to categorize chemical reactions
    - Combination – two or more substances form a new compound
    - Decomposition – a single compound is broken down into two or more products
    - Single-Replacement – atoms of one element replace the atoms of another element in a compound
    - Double-Replacement – an exchange of positive ions between two reacting compounds
    - Combustion Reactions – an element or compound reacts with oxygen often producing energy such as heat and light
  - Combination Reaction
    - In this type of reaction, two or more substances combine to form a new substance
    - When a Group A metal combines with a non-metal, you have a ionic reaction. For cations and anions, you should be able to write the correct formula by considering the charges
      - Potassium and Chlorine combine $K(s) + Cl_2(g) = KCl(s)$
      - *balances to* $2K(s) + Cl_2(g) = 2KCl(s)$
    - When two non-metals react by combination, often more than one product is possible
      - $S(s) + O_2(g) \rightarrow SO_2(g)$ *sulfur dioxide*
      - $S(s) + O_2(g) \rightarrow SO_3(g)$ *sulfur trioxide* $2S(s) + 3O_2(g) \rightarrow 2SO_3(g)$ *sulfur trioxide*
    - When a transition metal and a non-metal combine you may also get more than one product

- $Fe(s) + S(s) \rightarrow FeS(s)$ *iron(II) sulfide*
  $2Fe(s) + 3S(s) \rightarrow Fe_2S_3(s)$ *iron(III) sulfide*

- Some non-metal oxides react with water to produce an acid (hydrogen ions in aqueous solution)
  - $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$ *sulfurous acid*
- Some metallic oxides react with water to form a base (compound containing hydroxide ions). Use the ionic charges to derive the formula.
  - $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$ *calcium hydroxide*

Sample Problems

Complete the following combination reactions:

Here we have a metal and a non-metal – a typical ionic reaction

$$Al(s) + O_2 \rightarrow$$
$$Al^{3+}(s) + O^{2-}{}_2 \rightarrow Al_2O_3(s)$$
$$4Al(s) + 3O_2 \rightarrow 2Al_2O_3(s)$$

Here we have transition metal and a non-metal. Remember that transition metals can have multiple ionic charges!

$$Cu(s) + S(s) \rightarrow (\textit{two reactions possible})$$
$$Cu^{2+}(s) + S^{2-}(s) \rightarrow CuS(s) \textit{ balanced}$$
or
$$Cu^{1+}(s) + S^{2-}(s) \rightarrow Cu_2S$$
$$2Cu(s) + S(s) \rightarrow Cu_2S$$

And lastly, we have a non-metal oxide and water, which usually produces an acid when the two molecules combine.

$$SO_3(g) + H_2O(l) \rightarrow$$
$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l) \textit{ balanced}$$

- Decomposition Reactions
  - In decomposition reactions, a single compound is broken down into two or more products. The products can be an element or compounds.
  - Usually the decomposition of a compound can be difficult, except when a simple binary compound breaks down into its base elements.
  - Extremely rapid decomposition can often be accompanied by gases and heat – commonly known as an explosion. These decompositions usually require energy either as heat, light or electricity
  - When calcium carbonate is heated, it decomposes into calcium oxide (lime) and carbon dioxide

$$CaCO_3(s) \xrightarrow{heat} CaO(s) + CO_2(g)$$

  - Another example would be mercury(II) oxide when heated produces mercury and gaseous oxygen

$$2HgO(s) \xrightarrow{heat} 2Hg(l) + O_2(g)$$

- For a rapid decomposition example, how about TNT (trinitrotoluene)

  $$C_7H_5N_3O_6(s) \rightarrow N_2(g) + CO(g) + H_2O(g) + C(s)$$
  $$2C_7H_5N_3O_6(s) \rightarrow 3N_2(g) + 7CO(g) + 5H_2O(g) + 7C(s)$$

  - Notice that for every 2 mol of TNT that you have 15 moles of expanding gas

Sample Problems

Write a balanced equation for each decomposition:

$$H_2O(l) \xrightarrow{electricity}$$
$$H_2O(l) \xrightarrow{electricity} H_2(g) + O_2(g)$$
$$2H_2O(l) \xrightarrow{electricity} 2H_2(g) + O_2(g)$$

$$lead(IV)\ oxide \xrightarrow{heat}$$
$$PbO_2(s) \xrightarrow{heat} Pb(s) + O_2(g)\ balanced$$

- Single Replacement
  - In a single- replacement (also called single-displacement) reaction you start with an element and a compound. The element displaces an element in the compound, leaving you with that element and a new compound
  - Example would be dropping potassium into water
    - Result is fire on the water
    - What could be happening? Potassium replaces the hydrogen in the water, also liberating lots of heat. The hydrogen burns, and a base is produce – KOH
    - $K(s) + H_2 0(l) \rightarrow KOH(aq) + H_2(g) + heat$
      $2K(s) + 2H_2 0(l) \rightarrow 2KOH(aq) + H_2(g) + heat$
  - Whether a metal will displace another metal in a compound depends on the relative relativities of the two metals – a list of the activity series of metals is below
  - According to the list, a reactive metal will replace any metal listed below it
    - Iron will replace copper from a copper compound in solution
    - Magnesium will replace zinc in a zinc compound in solution or silver from a silver compound in solution
    - However, magnesium will never replace lithium or calcium in aqueous solutions of their compounds
    - $Mg(s) + Zn(NO_3)_2(aq) \rightarrow Mg(NO_2)_2(aq) + Zn(s)$
      $Mg(s) + 2AgNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2Ag(s)$
      $Mg(s) + LiNO_3(aq) \rightarrow no\ reaction$
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Activity Series of Metals		
	Name	**Symbol**
↓	Lithium	Li
	Potassium	K
	Calcium	Ca
	Sodium	Na
	Magnesium	Mg
	Zinc	Zn
	Iron	Fe
	Lead	Pb
	(Hydrogen)	(H)*
	Copper	Cu
	Mercury	Hg
	Silver	Ag

* Metals from Li to Na will replace H from acids and water. From Mg to Pb they will replace H from acids Only.

- A non-metal can also replace a non-metal from a compound. This is usually limited to the halogens (F_2, Cl_2, Br_2, and I_2). The activity decreases as you go down the group.

Sample Problems

Write a balanced chemical equation for each single-replacement reaction:

$$Zn(s) + H_2SO_4(aq) \rightarrow$$
$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

$$Na(s) + H_2O(l) \rightarrow$$
$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

$$Sn(s) + NaNO_3(aq) \rightarrow$$
$$Sn(s) + NaNO_3(aq) \rightarrow no\ reaction$$

$$Cl_2(g) + NaBr(aq) \rightarrow$$
$$Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(l)$$

- Double-Replacement Reactions
  - When dealing with two ionic compounds dissolved in water to form a homogenous mixture, there are two possibilities when they are mixed together.
    - They form a new homogenous mixture
    - A chemical reaction will occur

- Mixing solutions of NaCl and $CaNO_2$ results in a homogenous mixture
- Mixing solutions of aqueous solutions of barium chloride ($BaCl_2$) and potassium carbonate (KCO_3) results in a chemical reaction in which the barium and the potassium exchange places forming an aqueous solution of potassium chloride (K_2Cl) and a precipitate of barium carbonate ($BaCO_3$).
  - $K_2CO_3(aq) + BaCl_2(aq) \to KCl + BaCO_3$
    $K_2CO_3(aq) + BaCl_2(aq) \to 2KCl + BaCO_3 \; balanced$

- Generally double-replacement reactions involve the exchange of positive ions between two reacting compounds.
  - Typically the compounds are in aqueous solution and are often characterized by the production of a precipitate
  - For a double-replacement to take place on of the following is usually true
    - One product is only slightly soluble and precipitates from solution. For example sodium sulfide reacts with cadmium nitrate to produce a yellow precipitate of cadmium sulfide in a solution of sodium nitrate:
      - $Na_2S(aq) + Cd(NO_3)_2(aq) \to CdS(s) + 2NaNO_2(aq)$
    - One product is a gas that bubbles out of the mixture. An example would be hydrogen cyanide gas is produced when an aqueous sodium cyanide is mix with sulfuric acid
      - $2NaCN(aq) + H_2SO_4(aq) \to 2HCN(g) + Na_2SO_4(aq)$
    - One product is a molecular compound such as water. Combining solutions of calcium hydroxide and hydrochloric acid produces water as one of the products
      - $Ca(OH)_2(aq) + 2HCl(aq) \to CaCl_2(aq) + 2H_2O(l)$

Sample Problems:

Write a balanced equation for the each double replacement reaction:

$BaCl_2(aq) + K_2CO_3(aq) \to (A\ parcipitate\ of\ barium\ carbonate\ is\ formed)$

$BaCl_2(aq) + K_2CO_3(aq) \to BaCO_3(s) + KCl(aq)$

$BaCl_2(aq) + K_2CO_3(aq) \to BaCO_3(s) + 2KCl(aq)$

$FeS(s) + HCl(aq) \to (hydrogen\ sulfide\ gas\ (H_2S)\ is\ formed)$

$FeS(s) + HCl(aq) \to H_2S(g) + FeCl_2(aq)$

$FeS(s) + 2HCl(aq) \to H_2S(g) + FeCl_2(aq)$

- Combustion Reactions
  - In a combustion reaction, an element or compound reacts with oxygen, often producing energy as heat and light.
  - Some of the most common combustion reactions involve hydrocarbons – that is compounds of hydrogen and carbon.
    - These reactions always produce CO_2 and H_2O
    - In addition carbon (C) and CO may be additional products
    - Lack of oxygen may lead to incomplete combustion
    - Complete combustion of a hydrocarbon releases a large amount of energy as heat.
      - Methane (CH_4), Propane (C_3H_8) and butane (C_4H_{10}) are important fossil fuels
        - $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Sample Problems

Write a balanced equation for the complete combustion of these compounds

$benzene\ (C_6H_6)(l)$

$C_6H_6(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

$2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(g)$

$methanol(CH_3OH)(l)$

$CH_3OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$

- Predicting Products in a Chemical Reaction
  - With the five major type of reactions covered, we should mention a 5th type of reaction – oxidation - reduction, commonly known as redox
    - Oxidation is the combination of an element with oxygen to produce an oxide (iron(III) oxide – rust)
    - Reduction is the removal of oxygen from a compound (to make iron from iron ore:
      $2Fe_2O_3(s) + 3C(s) \rightarrow 4Fe(s) + 3CO_2(g)$ Fe2O3 + CO2
  - Five types summarized
    - Combination $R + S \rightarrow RS$
    - Decomposition $RS \rightarrow R + S$
    - Single-Replacement $T + RS \rightarrow TS + R$
    - Double-Replacement $R^+S^- + T^+U^- \rightarrow T^+S^- + R^+U^-$
    - Combustion $C_xH_y + (\frac{x+y}{4})\,O_2 \rightarrow xCO_2 + (y/2)H_2O$
- Reactions in Aqueous Solutions
  - Net Ionic Equations
    - With the world being 70% water, and even the adult human body being ~66% water, it is not surprising that many

important chemical reactions take place in water – that is an aqueous solution.

- When we looked at a double-replacement reaction, we spoke of the exchange of positive ions.
  - $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
  - This is more realistically written if we were to recognize that most ionic compounds dissociate or separate into anions and cations when they dissolve in water.
  - With this new information we can write a **complete ionic equation**
  - $$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow \\ AgCl(s) + Na^+(aq) + NO_3^-(aq)$$
  - We can simplify the equation by canceling ions that do not participate in the reaction
  - $$Ag^+(aq) + \cancel{NO_3^-(aq)} + \cancel{Na^+(aq)} + Cl^-(aq) \rightarrow \\ AgCl(s) + \cancel{Na^+(aq)} + \cancel{NO_3^-(aq)}$$
  - Leaving us with $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
  - Ions that are not directly involved in the reaction are called **spectator ions**. The reduced equation is called a **net ionic equation** – it only includes those particles actually taking part in the reaction
  - A net ionic equation still must be balanced – number of moles and the charge
  - Consider the following reaction: $Pb(s) + AgNO_3(aq) \rightarrow Ag(s) + Pb(NO_3)_2(aq)$
  - $Pb^{+2}(s) + Ag^+(s) + NO_3^-(aq) \rightarrow Ag^+(s) + Pb(NO_3)_2(aq)$
    $Pb^{2+}(s) + 2Ag^+(s) \rightarrow 2Ag^+(s) + Pb^{2-}(aq)$

Sample Problems

Identify the spectator ions and write the balanced net ionic equation for these reactions

$$HCl(aq) + ZnS(aq) \rightarrow H_2S(g) + ZnCl_2(aq)$$

$$H^+(aq) + Cl^-(aq) + Zn^+(aq) + S^-(aq) \rightarrow \\ H_2^+(aq) + S^-(g) + Zn^+(aq) + Cl_2^-(aq)$$

zinc and cholrine are the spectator ions

$$H^+(aq) + S^{2-}(aq) \rightarrow H_2S(g)$$

$$2H^+(aq) + S^{2-}(aq) \rightarrow H_2S(g)$$

$$Cl_2(g) + NaBr(aq) \rightarrow Br_2(l) + NaCl(aq)$$
$$Cl_2^-(g) + Na^+(aq) + Br^-(aq) \rightarrow Br_2^-(l) + Na^+(aq) + Cl^-(aq)$$
$$Spectator\ ion\ is\ Na^+(aq)$$
$$Cl_2^-(g) + Br^-(aq) \rightarrow Br_2^-(l) + Cl^-(aq)$$
$$Cl_2^-(g) + 2Br^-(aq) \rightarrow Br_2^-(l) + 2Cl^-(aq)$$

- Predicting the Formation of a Precipitate
  - As seen in double replacement, when you mix two ionic compounds you often get an aqueous solution and a solid which falls out of solution – the precipitate
  - Precipitates are often insoluble salts, and can be predicted with the use of a few general rules for solubility of ionic compounds

Solubility Rules for Ionic Compounds		
Compound	Solubility	Exceptions
1. Salts of alkali metals (Group 1A) and ammonia	Soluble	Some lithium compounds
2. Nitrate salts and chlorate salts (NO_3, ClO_3)	Soluble	Few exceptions
3. Sulfate salts (SO_4)	Soluble	Compounds of Pb, Ag, Hg, Ba, Sr, and Ca
4. Chloride salts	Soluble	Compounds of Ag and some compounds of Hg and Pb
5. Carbonates, phosphates, chromates, sulfides, and hydroxides	Most are insoluble	Compounds of the alkali metals and of ammonia

- Using the table – lets take an example of two aqueous solutions of sodium sulfate (Na_2SO_4) and barium nitrate ($Ba(NO_3)_2$) are mixed together
  - First we need to recognize that while there are essentially 2 elements and 2 polyatomic ions, we could have up to 4 combinations, except that the charges limit us to two combinations
  $$NaSO_4(aq) + Ba(NO_3)_2(aq) \rightarrow ?$$
  - $$2Na^+ + SO_4^{2-}(aq) + Ba^{2+} + 2NO_3^-(aq) \rightarrow ?$$
  $$2Na^+ + SO_4^{2-}(aq) + Ba^{2+} + 2NO_3^-(aq) \rightarrow$$
  $$NaNO_3(?) + BaSO_4(?)$$

- Here we see the two possible combinations in double replacement, so which one will be the precipitate?
  - According to rule 1, salts of alkali metals (Group 1A) are soluble – so that means that $NaNO_3$ will stay aqueous due to the Na
  - According to rule 2, nitrate salts and chlorate salts are soluble – again proof that $NaNO_3$ will stay aqueous due to the NO_3.
  - Ba is not covered by any of the rules – so that means it is insoluble (will not stay as an aqueous solution) and therefore falls out of solution.
  - This means that Na+ and NO3- are spectator ions and the net ionic equation can be written as

$$Ba^{2+}(aq) + SO_4^{\ 2-}(aq) \to + BaSO_4(s)$$

Sample Problems

Using the solubility rules, identify the precipitates formed and write the net ionic equation for the reaction of aqueous potassium carbonate with aqueous strontium chloride

First set up the skeleton equation

$$K_2CO_3(aq) + SrCl_2(aq) \to KCl(?) + SrCO_3(?)$$

break it apart to the ionic equation on the reac tan *t side*

$$K_2^{\ +}(aq) + CO_3^{\ 2-}(aq) + Sr^{2+}(aq) + Cl_2^{\ -}(aq) \to KCl(?) + SrCO_3(?)$$

Looking at the rules for solubility, rule 1 would apply to the potassium in potassium chloride. Rule 2 has no applicability, nor does rule 3. Rule 4 says that chloride salts are soluble, so again KCl is soluble. Rule 5 says that carbonates are insoluble, so that means that $SrCO_3$ would be the precipitate. This means that potassium and chloride are the spectator ions

$$Sr^{2+}(aq) + CO_3^{\ 2-}(aq) \to SrCO_3(s)$$

Identify the precipitate formed, and write the net ionic equation for mixing a solutions of NH_4Cl(aq) and $Pb(NO_3)_2$(aq)

$$NH_4Cl(aq) + Pb(NO_3)_2(aq) \to$$

$$NH_4^{\ +}(aq) + Cl^-(aq) + Pb^{2+}(aq) + NO_3^{\ -}(aq) \to$$

$$NH_4^{\ +}(aq) + Cl^-(aq) + Pb^{2+}(aq) + NO_3^{\ -}(aq) \to$$
$$NH_4NO_3(?) + PbCl_2(?)$$

Looking at our rules, Rule 1 covers ammonia; rule 2 covers nitrates (that's two hits for NH_4NO_3 being soluble); rule 3 does not cover anything, rule 4 covers chloride salts with lead as being soluble), and rule 5 is not used.

$$Pb^{2+}(aq) + Cl^-(aq) \to PbCl_2(s)$$

$$Pb^{2+}(aq) + 2Cl^-(aq) \to PbCl_2(s)$$

Write a complete ionic equation for the reaction of aqueous solutions of iron(III) nitrate and sodium hydroxide

$$Fe(NO_3)_3(aq) + NaOH(aq) \rightarrow ?$$

$$Fe^{3+}(aq) + 3NO_3^-(aq) + 3Na^+(aq) + 3OH^-(aq) \rightarrow$$
$$3Na^+(aq) + 3NO_3^-(aq) + Fe(OH)_3(aq)$$

$$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$$

$Fe(OH)_3$ forms a precipitate by rule 5 – hydroxides are insoluble