Chapter 7

Balancing Molecular Equations

Chemists write chemical equations to illustrate what is happening during a chemical reaction—bonds are broken, atoms are rearranged and new bonds are formed. Every chemical reaction supports the Law of Conservation of Matter. This means that in every reaction, the number of atoms of each type of element contained within the reactants must be the same as the number of atoms of each type of element contained within the products.

Balancing equations is a process which assures that equations are written properly to support the Law of Conservation of Matter; however, balancing cannot be done until each reactant and product formula is written correctly. It is important to properly write the seven elements that are diatomic in their elemental form and also to use subscripts and parentheses appropriately when considering the oxidation numbers of ions. All compounds must be made neutral before beginning to balance atoms.

Balancing is accomplished by adding coefficients that multiply the number of atoms represented by the formula. For example, a coefficient of 2 in front of oxygen (e.g., $2O_2$) means that 4 oxygen atoms are represented. Unlike algebra, in chemistry a coefficient does not need to be outside parentheses or brackets to be distributed. A coefficient applies to the complete substance; however, it no longer applies when a plus sign (+) or arrow ($\longrightarrow$) is encountered. For example,

$3(NH_4)_2CO_3$ shows 6 nitrogen, 24 hydrogen, 3 carbon and 9 oxygen atoms

$3MgCl_2 + NaBr$ indicates 3 magnesium, 6 chlorine, 1 sodium and 1 bromine atom

Tips for Balancing Equations

- Be sure each molecular formula is written correctly and each compound is neutral.
- Mentally count or tally how many of each type of atom is present on each side of the equation.
- Begin by balancing elements that are only found in one substance on each side.
- Balance oxygen and hydrogen *last*—they usually balance out at the end or perhaps only the number of water molecules will need to be adjusted.
- If there is an odd number of an element on one side and an even number on the other, the odd number will need to be evened out—so use a coefficient of 2 for that substance.
- If there are polyatomic ions that remain together as a unit during the reaction, count the polyatomic ion as a unit.
- When tallying, be sure to adjust the count for each and every element that an added coefficient affects.
- Combustion reactions that don't seem to balance will often come out better if a coefficient of 2 is used for the hydrocarbon.

Some Examples

Hydrogen and chlorine gases react to form hydrogen chloride gas.

$$H_2 + Cl_2 \longrightarrow HCl$$

H	2	H	1
Cl	2	Cl	1

Because one hydrogen and one chlorine atom appears to have been lost, a coefficient of 2 is added to the HCl.

$$H_2 + Cl_2 \longrightarrow 2HCl$$

H	2	H	2
Cl	2	Cl	2

Potassium chlorate is heated and decomposes into potassium chloride and oxygen gas.

$$KClO_3 \longrightarrow KCl + O_2$$

K	1	K	1
Cl	1	Cl	1
O	3	O	2

Because there is an odd number of oxygen atoms on the left and an even number on the right, a coefficient of 2 is added to the potassium chlorate.

$$2KClO_3 \longrightarrow KCl + O_2$$

K	2	K	1
Cl	2	Cl	1
O	6	O	2

The oxygen atoms can now be balanced using a coefficient of 3; the potassium and chlorine atoms can be balanced using a coefficient of 2 in front of KCl.

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

K	2	K	2
Cl	2	Cl	2
O	6	O	6

Ammonium nitrate reacts with calcium phosphate to form ammonium phosphate and calcium nitrate.

$$NH_4NO_3 + Ca_3(PO_4)_2 \longrightarrow (NH_4)_3PO_4 + Ca(NO_3)_2$$

Balancing this equation can be simplified greatly by noticing that each of the polyatomic ions remain together as a group throughout the reaction. Note that the tallying below indicates this.

NH_4	1	NH_4	3
NO_3	1	NO_3	2
Ca	3	Ca	1
PO_4	2	PO_4	1

Beginning with calcium, add a coefficent of 3 on the right to balance the Ca. This also changes the nitrate count to 6 on the right, so adjust that by adding a coefficient of 6 on the left side.

$$6NH_4NO_3 + Ca_3(PO_4)_2 \longrightarrow (NH_4)_3PO_4 + 3Ca(NO_3)_2$$

NH_4	**6**	NH_4	3
NO_3	**6**	NO_3	**6**
Ca	3	Ca	**3**
PO_4	2	PO_4	1

Finally, a coefficient of 2 is needed in front of ammonium phosphate to complete the balancing.

$$6NH_4NO_3 + Ca_3(PO_4)_2 \longrightarrow 2(NH_4)_3PO_4 + 3Ca(NO_3)_2$$

NH_4	6	NH_4	**6**
NO_3	6	NO_3	6
Ca	3	Ca	3
PO_4	2	PO_4	**2**

As a double check, it is always wise to count the total number of each type of element indicated. For this reaction there is a total of 12 N, 24 H, 26 O, 3 Ca and 2 P atoms on each side.

Now it is time to apply the rules for balancing equations to some chemical reactions. It turns out that many chemical reactions can be organized into various groups—that is the focus of the sections that follow.

Exercise 7-1: Balance the following equations by adding coefficients as needed. Some equations may already be balanced.

1. $Ca + HOH \longrightarrow Ca(OH)_2 + H_2$
2. $Cl_2O_7 + H_2O \longrightarrow HClO_4$
3. $Fe + O_2 \longrightarrow Fe_3O_4$
4. $C_6H_{14} + O_2 \longrightarrow CO_2 + H_2O$
5. $Ca_3(PO_4)_2 + H_2SO_4 \longrightarrow Ca(H_2PO_4)_2 + CaSO_4$
6. $AlCl_3 + AgNO_3 \longrightarrow Al(NO_3)_3 + AgCl$
7. $HCl + CaCO_3 \longrightarrow CO_2 + HOH + CaCl_2$
8. $WO_3 + H_2 \longrightarrow W + H_2O$
9. $Cl_2 + H_2O \longrightarrow HCl + HClO$
10. $Cl_2 + NaI \longrightarrow NaCl + I_2$

ROUND 2

Synthesis Reactions occur when two or more reactants combine to form a single product. There are several common types of synthesis reactions.

A metal combines with a nonmetal to form a binary salt.

example A piece of lithium metal is dropped into a container of nitrogen gas.

$6Li + N_2 \longrightarrow 2Li_3N$

Metallic oxides and water form bases (metallic hydroxides).

example Solid sodium oxide is added to water.

$Na_2O + HOH \longrightarrow 2NaOH$

Solid magnesium oxide is added to water.

$MgO + 2HOH \longrightarrow Mg(OH)_2$

Nonmetallic oxides and water form acids. The nonmetal retains its oxidation number.

example Carbon dioxide is bubbled into water.

$CO_2 + H_2O \longrightarrow H_2CO_3$ (Oxidation number of C is +4)

Dinitrogen pentoxide is bubbled into water.

$N_2O_5 + H_2O \longrightarrow 2HNO_3$ (Oxidation number of N is +5)

Metallic oxides and nonmetallic oxides form salts.

example Solid sodium oxide is added to carbon dioxide.

$Na_2O + CO_2 \longrightarrow Na_2CO_3$

Solid calcium oxide is added to sulfur trioxide.

$CaO + SO_3 \longrightarrow CaSO_4$

Decomposition Reactions occur when a single reactant is broken down into two or more products.

Metallic carbonates decompose into metallic oxides and carbon dioxide.

example A sample of magnesium carbonate is heated.

$MgCO_3 \longrightarrow MgO + CO_2$

Metallic chlorates decompose into metallic chlorides and oxygen.

example A sample of magnesium chlorate is heated.

$Mg(ClO_3)_2 \longrightarrow MgCl_2 + 3O_2$

Ammonium carbonate decomposes into ammonia, water and carbon dioxide.

example A sample of ammonium carbonate is heated.

$(NH_4)_2CO_3 \longrightarrow 2NH_3 + H_2O + CO_2$

Sulfurous acid decomposes into sulfur dioxide and water.

example A sample of sulfurous acid is heated.

$H_2SO_3 \longrightarrow H_2O + SO_2$

Carbonic acid decomposes into carbon dioxide and water.

example A sample of carbonic acid is heated.

$$H_2CO_3 \longrightarrow H_2O + CO_2$$

A binary compound may break down to produce two elements.

example Molten sodium chloride is electrolyzed.

$$2NaCl \longrightarrow 2Na + Cl_2$$

Hydrogen peroxide decomposes into water and oxygen.

example $2H_2O_2 \longrightarrow 2H_2O + O_2$

Ammonium hydroxide decomposes into ammonia and water.

example $NH_4OH \longrightarrow NH_3 + HOH$

Exercise 7–2: Predict and balance the following synthesis and decomposition reactions. Use abbreviations to indicate the phase of reactants and products where possible [i.e., (aq) (s) (l) (g)].

1. A sample of calcium carbonate is heated.

2. Sulfur dioxide gas is bubbled through water.

3. Solid potassium oxide is added to a container of carbon dioxide gas.

4. Liquid hydrogen peroxide is warmed.

5. Solid lithium oxide is added to water.

6. Molten aluminum chloride is electrolyzed.

7. A pea-sized piece of sodium is added to a container of iodine vapor.

8. A sample of carbonic acid is heated.

9. A sample of potassium chlorate is heated.

10. Solid magnesium oxide is added to sulfur trioxide gas.

Notes

Chapter 8

Single Replacement Reactions

Single replacement reactions are reactions that involve an element replacing one part of a compound. The products include the displaced element and a new compound. An element can only replace another element that is less active than itself.

#memorize

General activity series for metals

(most active) Li Ca Na Mg Al Zn Fe Pb [H_2] Cu Ag Pt *(least active)*

General activity series for nonmetals

(most active) F_2 Cl_2 Br_2 I_2 *(least active)*

Here are some common types of single replacement reactions.

Active metals replace less active metals from their compounds in aqueous solution.

example Magnesium turnings are added to a solution of iron(III) chloride.

$3Mg + 2FeCl_3 \longrightarrow 2Fe + 3MgCl_2$

Active metals replace hydrogen in water.

example Sodium is added to water.

$2Na + 2HOH \longrightarrow H_2 + 2NaOH$

Active metals replace hydrogen in acids.

example Lithium is added to hydrochloric acid.

$2Li + 2HCl \longrightarrow H_2 + 2LiCl$

Active nonmetals replace less active nonmetals from their compounds in aqueous solution.

example Chlorine gas is bubbled into a solution of potassium iodide.

$Cl_2 + 2KI \longrightarrow I_2 + 2KCl$

If a less reactive element is combined with a more reactive element in compound form, there will be no resulting reaction.

example Chlorine gas is bubbled into a solution of potassium fluoride.

$Cl_2 + KF \longrightarrow$ no reaction

example Zinc is added to a solution of sodium chloride.

$Zn + NaCl \longrightarrow$ no reaction

ROUND 3

Exercise 8–1: Using the activity series, predict and balance the following single replacement reactions. Use abbreviations to indicate the appropriate phase of reactants and products where possible. *Note: Not all of the reactions will occur. For those that do not, write no reaction.*

1. A piece of copper is dropped into a container of water.

2. Liquid bromine is added to a container of sodium iodide crystals.

3. An aluminum strip is immersed in a solution of silver nitrate.

4. Zinc pellets are added to a sulfuric acid solution.

5. Fluorine gas is bubbled into a solution of aluminum chloride.

6. Magnesium turnings are added to a solution of lead(II) acetate.

7. Iodine crystals are added to a solution of sodium chloride.

8. Calcium metal is added to a solution of nitrous acid.

9. A pea-sized piece of lithium is added to water.

10. A solution of iron(III) chloride is poured over a piece of platinum wire.

Note: On the AP reaction prediction section, all reactions "work"; in other words there will be no "No reactions" on the AP Exam.

Chapter 9

Double Replacement (Metathesis) Reactions

In many reactions between two compounds in aqueous solution, the cations and anions appear to switch partners according to the following equation:

$$AX + BY \longrightarrow AY + BX$$

The two compounds react to form two new compounds. No changes in oxidation numbers occur. Reactions of this type are known as *double replacement* or *metathesis reactions*. An example of such a reaction would be the mixing of aqueous solutions of potassium bromide and silver nitrate forming insoluble silver bromide (precipitate) and aqueous potassium nitrate:

$$KBr(aq) + AgNO_3(aq) \longrightarrow AgBr(s) + KNO_3(aq)$$

Note that each cation pairs up with the anion in the other compound, thus switching partners. Anions do not pair up with anions and cations do not pair up with cations. Likes repel; opposites attract!

All double replacement reactions must have a "driving force" or a reason why the reaction will occur or "go to completion." The "driving force" in metathesis reactions is the removal of at least one pair of ions from solution.

This removal of ions can occur in one of three ways:

1. **Formation of a precipitate:** A precipitate is an insoluble substance (solid) formed by the reaction of two aqueous substances. It is the result of ions bonding together so strongly that the solvent (water) cannot pull them apart. The insoluble solid (or solids if a double precipitate occurs) will settle out (precipitate) from the solution and this results in the removal of ions from the solution.
2. **Formation of a gas:** Gases may form directly in a double replacement reaction or from the decomposition of one of the products. The gases will bubble off or evolve from the solution.
3. **Formation of primarily molecular species:** The formation of primarily unionized molecules in solution removes ions from the solution and the reaction "works" or is said to go to completion. Unionized or partially ionized molecules give solutions that are known as *nonelectrolytes* or *weak electrolytes*. The best known nonelectrolyte is water formed in acid–base neutralization reactions. Acetic acid is an example of an acid that is primarily molecular (weak electrolyte) when placed in water.

Reversible Reactions

If a double replacement reaction does not go to completion (no precipitate, gas or molecular species is formed), then the reaction is reversible (no ions have been removed). Reversible reactions are at equilibrium and have both forward and reverse reactions taking place. In a reversible reaction, evaporation of the water solvent will result in solid residues of both reactants and products. The reaction is not driven to completion (products) because no ions have been removed. A double arrow is used to designate a reversible reaction at equilibrium.

$$BaCl_2(aq) + 2NaNO_3(aq) \rightleftharpoons Ba(NO_3)_2(aq) + 2NaCl(aq)$$

Solubility Rules Table

The solubility classification of ionic substances according to their solubility in water is difficult. Nothing is completely "insoluble" in water. The degree of solubility varies from one "soluble" substance to another. Nevertheless, a solubility classification scheme is useful even though it must be regarded as an approximate guideline.

MAINLY WATER SOLUBLE	
NO_3^-	All nitrates are soluble.
CH_3COO^- or $C_2H_3O_2^-$	All acetates are soluble except $AgCH_3COO$*.
ClO_3^-	All chlorates are soluble.
Cl^-	All chlorides are soluble except AgCl, Hg_2Cl_2 $PbCl_2$*.
Br^-	All bromides are soluble except AgBr, $PbBr_2$*, Hg_2Br_2 and $HgBr_2$*.
I^-	All iodides are soluble except AgI, Hg_2I_2, HgI_2 and PbI_2.
SO_4^{2-}	All sulfates are soluble except $BaSO_4$, $PbSO_4$, Hg_2SO_4, $CaSO_4$, Ag_2SO_4* and $SrSO_4$*.
Alkali metal cations (Group IA) and NH_4^+	All are soluble.
H^+	All common inorganic acids and low molecular mass organic acids are soluble.
MAINLY WATER INSOLUBLE	
CO_3^{2-}	All carbonates are insoluble except those of the IA elements and NH_4^+.
CrO_4^{2-}	All chromates are insoluble except those of of the IA elements, NH_4^+, $CaCrO_4$* and $SrCrO_4$*.
OH^-	All hydroxides are insoluble except those of the IA elements, NH_4^+, $Ba(OH)_2$, $Sr(OH)_2$*, and $Ca(OH)_2$*.
PO_4^{3-}	All phosphates are insoluble except those of the IA elements and NH_4^+.
SO_3^{2-}	All sulfites are insoluble except those of the IA elements and NH_4^+.
S^{2-}	All sulfides are insoluble except those of the IA and IIA elements and NH_4^+.

*_Soluble_ compounds dissolve to the extent of at least 10 g/L at 25 °C. _Slightly soluble_ compounds (marked with an *) dissolve in the range of from 1 g/L to 10 g/L at 25 °C. Those compounds that have a solubility of less than 1 g/L are considered to be _insoluble_. These standards are common but arbitrary.

ROUND 4

Formation of a Precipitate

In order to predict double replacement reactions yielding precipitates, one must memorize the solubility rules listed on page 48.

Exercise 9–1: Predict and balance the following metathesis reactions based on the solubility of the products. Use the abbreviations (aq) and (s) for the reactants and products. All reactants are aqueous. *Note:* Some of these reactions do not go to completion.

1. silver nitrate + potassium chromate

2. ammonium chloride + cobalt(II) sulfate

3. lithium hydroxide + sodium chromate

4. zinc acetate + cesium hydroxide

5. ammonium sulfide + lead(II) nitrate

6. iron(III) sulfate + barium iodide

7. chromium(III) bromide + sodium nitrate

8. rubidium phosphate + titanium(IV) nitrate

9. ammonium carbonate + nickel(II) chloride

10. tin(IV) nitrate + potassium sulfite

Note: Correct molecular formulas must be written for both the reactants and products before an equation may be balanced.

ROUND 5

Formation of a Gas

Common gases formed in metathesis reactions are listed in the table below.

	Common Gases
H_2S	Any sulfide (salt of S^{2-}) plus any acid form $H_2S(g)$ and a salt.
CO_2	Any carbonate (salt of CO_3^{2-}) plus any acid form $CO_2(g)$, HOH and a salt.
SO_2	Any sulfite (salt of SO_3^{2-}) plus any acid form $SO_2(g)$, HOH and a salt.
NH_3	Any ammonium salt (salt of NH_4^+) plus any soluble strong hydroxide react upon heating to form $NH_3(g)$, HOH and a salt.

Reactions that produce three of the gases (CO_2, SO_2, and NH_3) involve the initial formation of a substance that breaks down to give the gas and HOH.

Example 1. The reaction of Na_2SO_3 and HCl produces H_2SO_3:

$$Na_2SO_3(aq) + 2HCl(aq) \longrightarrow H_2SO_3(aq) + 2NaCl(aq)$$

Bubbling is observed in this reaction because the H_2SO_3 (sulfurous acid) is unstable and immediately decomposes to give HOH and SO_2 gas:

$$H_2SO_3 \longrightarrow HOH(l) + SO_2(g)$$

The molecular equation for the complete reaction, therefore, is:

$$Na_2SO_3(aq) + 2HCl(aq) \longrightarrow HOH(l) + SO_2(g) + 2NaCl(aq)$$

Example 2. A typical reaction of a carbonate and an acid is:

$$K_2CO_3(aq) + 2HNO_3(aq) \longrightarrow HOH(l) + CO_2(g) + 2KNO_3(aq)$$

Bubbling is also observed in this reaction. Theoretically H_2CO_3, carbonic acid, is formed, but the acid is unstable and immediately decomposes to form carbon dioxide gas and water according to the following equation:

$$H_2CO_3(aq) \longrightarrow HOH(l) + CO_2(g)$$

Example 3. Ammonium salts and soluble bases react as follows (particularly when the solution is warmed):

$$NH_4Cl(aq) + NaOH(aq) \longrightarrow NH_3(g) + HOH(l) + NaCl(aq)$$

The odor of ammonia gas is noted and moist blue litmus paper held near the mouth of the container will turn blue. Theoretically NH_4OH, ammonium hydroxide, is produced (also known as ammonia water). The compound is unstable and decomposes into ammonia gas and water:

$$NH_4OH(aq) \longrightarrow NH_3(g) + HOH(l)$$

Example 4. The odor of rotten eggs and bubbling are noted when an acid is added to a sulfide. A typical reaction producing hydrogen sulfide gas is:

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

Helpful Tip: Be aware of reactions involving the formation of carbon dioxide, sulfur dioxide, ammonia, and hydrogen sulfide gases on the AP Chemistry Examination. Over the years these reactions have appeared many, many times. Know these four gases and how they are produced!

Exercise 9–2: Predict and balance the following metathesis reactions. Use the abbreviations (s), (l), (g), and (aq) for the reactants and products. All reactants are aqueous unless otherwise stated.

1. ammonium sulfate and potassium hydroxide are mixed together

2. ammonium sulfide is reacted with hydrochloric acid

3. cobalt(II) chloride is combined with silver nitrate

4. solid calcium carbonate is reacted with sulfuric acid

5. potassium sulfite is reacted with hydrobromic acid

6. potassium sulfide is reacted with nitric acid

7. ammonium iodide + magnesium sulfate

8. solid titanium(IV) carbonate + hydrochloric acid

9. solid calcium sulfite + acetic acid

10. strontium hydroxide + ammonium sulfide

ROUND 6

Formation of a Molecular Species *(Weak or Nonelectrolytes)*

Metathesis reactions that produce primarily molecules in the form of partially dissociated or ionized molecules (weak electrolytes) or molecules that do not ionize or dissociate at all (nonelectrolytes) serve as the driving force in some aqueous reactions. Forming molecular products in double replacement reactions results in the removal of ions from solution. Such reactions tend to go to completion (shift to the right) and form primarily products.

Simplified list of rules:

A. The common strong acids and thus strong electrolytes are $HClO_4$, $HClO_3$, HCl, HBr, HI, HNO_3, and H_2SO_4. (Memorize these seven strong acids!) All other common acids are weak acids and thus weak electrolytes (CH_3COOH, H_3PO_4, HF, and HNO_2 are examples of weak acids. (*Note:* All organic acids (R–COOH) are weak electrolytes.) All strong acids in their pure form (as opposed to dilute aqueous form) are nonelectrolytes (molecular). When water is added, the action of the solvent water with a strong acid produces a hydrated proton (hydronium ion) and a negatively charged anion. The process of making ions from molecular species is known as ionization. Strong acids ionize 100% in water. An example of a strong electrolyte undergoing ionization is as follows:

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

This reaction may be abbreviated as:

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

B. The common strong bases are the soluble hydroxides (those of Group IA elements and Ba^{2+}) and the slightly soluble hydroxides (those of Ca^{2+} and Sr^{2+}). Strong bases, like strong acids, are strong electrolytes. Memorize the strong bases! NH_4OH is a soluble weak electrolyte which normally decomposes into $NH_3(g)$ and HOH(l). Technically speaking, the pure compound ammonium hydroxide has never been isolated and the substance is more correctly known as aqueous ammonia. Most other hydroxides are insoluble. Pure *liquid* hydroxides are strong electrolytes because they already contain ions. The action of the solvent water releasing the ions of a base into solution is known as dissociation. Acids ionize in water; bases dissociate!

C. Most common (soluble) salts are strong electrolytes and thus dissociate into ions when placed in water.

D. Water is a weak electrolyte which is typically produced in acid–base neutralization reactions.

Some Examples of Weak Electrolytes Forming as Products (shown in bold):

$$Ca(CH_3COO)_2(aq) + 2HCl(aq) \longrightarrow CaCl_2(aq) + \mathbf{2CH_3COOH(aq)}$$

$$2Na_3PO_4(aq) + 3H_2SO_4(aq) \longrightarrow 3Na_2SO_4(aq) + \mathbf{2H_3PO_4(aq)}$$

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + \mathbf{HOH(l)}$$

Exercise 9–3: Predict and balance the following reactions. Use the abbreviations (s), (l), (g), and (aq) for the reactants and products. All reactants are aqueous unless otherwise stated.

1. carbon dioxide gas is bubbled through a solution of lithium hydroxide

2. sodium nitrite is reacted with hydrochloric acid

3. ammonium bromide + sodium hydroxide

4. carbon dioxide gas is reacted with solid potassium oxide

5. solid magnesium oxide is reacted with hydrochloric acid

6. equal numbers of moles of potassium hydroxide and phosphoric acid react

7. sodium fluoride reacts with dilute nitric acid

8. ammonium carbonate + potassium bromide

9. oxalic acid (0.1 M) reacts with an equal volume of cesium hydroxide (0.1 M)

10. silver nitrate + sodium chromate

Acid–Base Neutralization Reactions

Acids react with bases to produce salts and water. One mole of hydrogen ions will react with one mole of hydroxide ions to produce one mole of water. Learn which acids are strong acids (written in ionic form) and which are weak acids (written in molecular form). Check the solubility rules for the solubility of the salt produced. If it is soluble, it is written in ionic form; if it is insoluble it is written in molecular form. This will be covered further in Chapter 10.

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water}$$

(A salt consists of a cation from a base and an anion from an acid. e.g., the salt sodium sulfate contains sodium ions from sodium hydroxide and sulfate ions from sulfuric acid.)

Example 1: Hydrogen sulfide gas is bubbled through excess potassium hydroxide solution.

$$H_2S(g) + 2KOH(aq) \longrightarrow K_2S(aq) + 2HOH(l)$$

Polyprotic acids can be tricky when it comes to predicting neutralization reactions. Sulfuric acid and phosphoric acid are classic examples frequently encountered on AP Examinations. If the base is in excess, all hydrogen ions will react with strong base to produce water.

Example 2: Dilute sulfuric acid is reacted with excess sodium hydroxide.

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2HOH(l)$$

If, however, the reaction in Example 2 stated that equal numbers of moles of sulfuric acid and sodium hydroxide react, then the coefficients for both reactants must be one and the salt that forms is sodium hydrogen sulfate.

Example 3: Equal number of moles of sulfuric acid and sodium hydroxide react.

$$H_2SO_4(aq) + NaOH(aq) \longrightarrow NaHSO_4(aq) + HOH(l)$$

Take into account information dealing with the quantity of each reactant.

Example 4: Equal volumes of 0.1 M phosphoric acid and 0.2 M sodium hydroxide are reacted together.

$$H_3PO_4(aq) + 2NaOH(l) \longrightarrow Na_2HPO_4(aq) + 2HOH(l)$$

Watch for substances that react with water before reacting with an acid or a base. The acid and basic anhydrides covered in Chapter 7 behave in such a manner. These are really two-step reactions.

Example 5: Excess sulfur dioxide gas is bubbled into a saturated solution of calcium hydroxide.

$$SO_2(g) + Ca(OH)_2(aq) \longrightarrow CaSO_3(s) + HOH(l)$$

(Remember, $SO_2(g)$ is an acid anhydride)

If an acid + base yields a salt + water, then an **acid anhydride + basic anhydride will yield a salt.**

Example 6: Sulfur dioxide gas and solid calcium oxide are reacted together.

$$SO_2(g) + CaO(s) \longrightarrow CaSO_3(s)$$

(*Note:* $SO_2(g)$ is the acid anhydride for sulfurous acid and $CaO(s)$ is the basic anhydride for calcium hydroxide.)