STOICHIOMETRY

- ***Stoichiometry*** is the quantitative relationship between the ***reactants and products*** in a balanced ***chemical equation***. Stoichiometry allows chemists to predict how much of a reactant is necessary to form a given amount of product or how much of a reactant is required to completely react with another reactant.

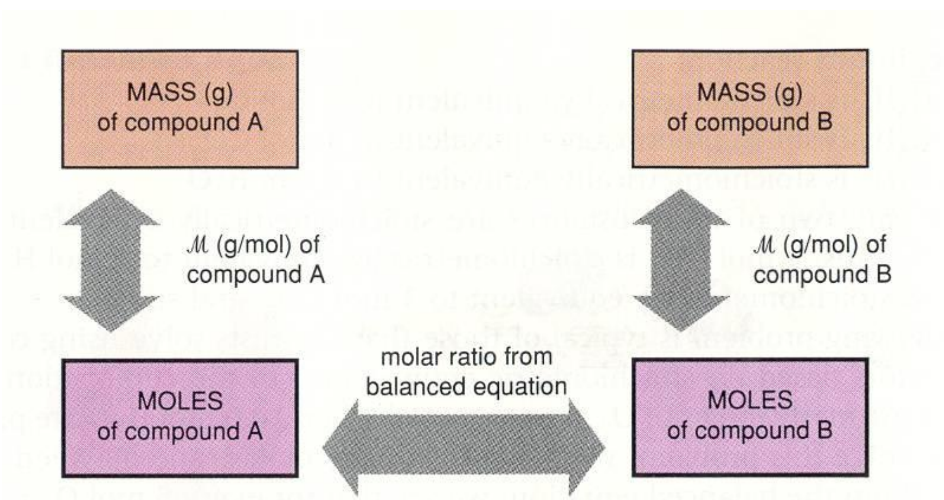
- A balanced chemical equation provides several important information about the reactants and products in a chemical reaction. For example:

$1N_2(g)$	+	$3\ H_2(g)$	$\longrightarrow$	$2\ NH_3(g)$
1 molecule		**3 molecules**		**2 molecules**
100 molecules		**300 molecules**		**200 molecules**
1 million molecules		**3 million molecules**		**2 million molecules**

1 mole N_2 + 3 moles H_2 $\longrightarrow$ 2 moles NH_3

- The coefficient in a chemical equation specify the relative amounts in moles of each of the substances in the reaction.

- Stoichiometry calculations can be classified as one of the following:



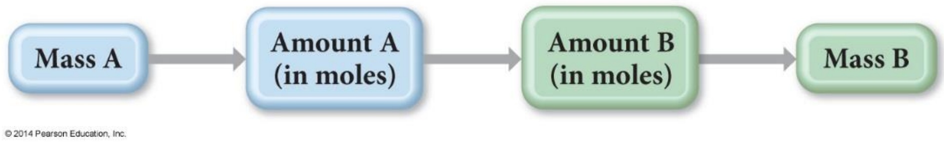
Summary of Stoichiometric Calculations in Chemistry

STOICHIOMETRIC CALCULATIONS

Mass-Mass Calculations:

- The most common stoichiometric calculations in chemistry involve mass-to-mass conversions where the mass reactants and products are related to one another in a chemical equation.

- The general conceptual plan for these calculations is shown below:

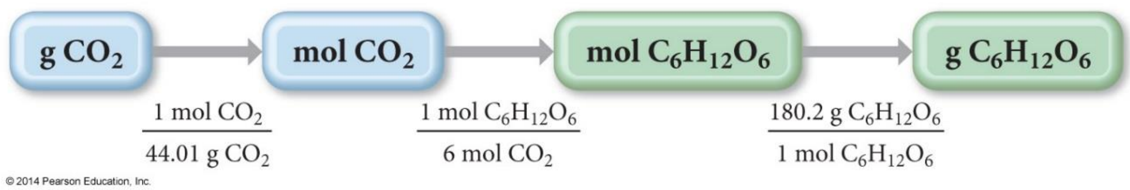


Examples:

1. In photosynthesis, plants convert carbon dioxide and water into glucose according to the reaction below:

$$6\ CO_2\ (g)\ +\ 6\ H_2O\ (g) \xrightarrow{\text{sunlight}} 6\ O_2\ (g) + C_6H_{12}O_6\ (aq)$$

If a plant consumes 37.8 g of CO_2 in a week, what mass of glucose (in g) can the plant synthesize from this reaction? (Assume excess water present)



STOICHIOMETRIC CALCULATIONS

Examples (cont's):

2. One component of acid rain is nitric acid (HNO_3) which is formed when NO_2, a pollutant reacts with oxygen and water according to the reaction shown below:

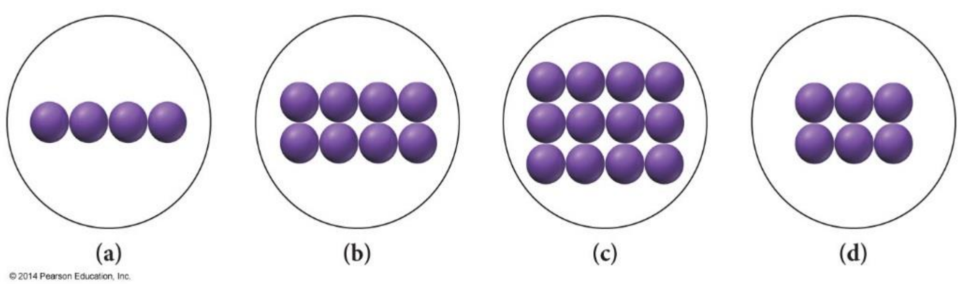
$$4\ NO_2\ (g)\ +\ O_2\ (g)\ +\ 2\ H_2O\ (l)\ \rightarrow\ 4\ HNO_3\ (aq)$$

If a medium-sized home produces 16 kg of NO_2 each year from generation of electricity, what mass of nitric acid (in kg) is produced from the reaction of this amount of NO_2?

3. Sodium reacts with oxygen to form sodium oxide, as shown below:

$$4\ Na\ (s)\ +\ O_2\ (g)\ \rightarrow\ 2\ Na_2O\ (s)$$

A flask contains oxygen, as shown in the diagram to the right. Which of the diagrams below show the proper amount of sodium needed to completely react with this amount of oxygen?



LIMITING REACTANT, THEORETICAL YIELD & PERCENT YIELD

- When ***2 or more reactants*** are combined in ***non-stoichiometric*** ratios, the amount of ***product*** produced is ***limited*** by the reactant that is ***not in excess*** (***limiting reactant***).

- When solving these problems, assume reactant 1 to be the limiting reactant, and calculate the amount of product that can be formed based on this assumption. Then repeat this process with reactant 2. The lesser amount of product formed is considered the correct assumption for the limiting reactant. (**Method 1**)

- Alternately, choose any reactant and determine the amount of the other required to react completely with it. If more of the second reactant are present than calculated, then the first reactant is the limiting reactant. Otherwise, the second reactant is the limiting reactant. (**Method 2**)

Summary:

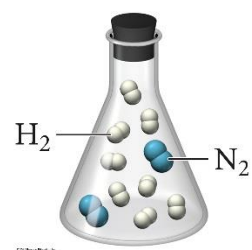
- **Limiting Reactant:** the reactant that is completely used during a chemical reaction.
- **Theoretical Yield**: the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.
- **Actual Yield**: the amount of product actually produced in a chemical reaction.
- **Percent Yield**: calculated as $\frac{\text{actual yield}}{\text{theoretical yield}} x100$

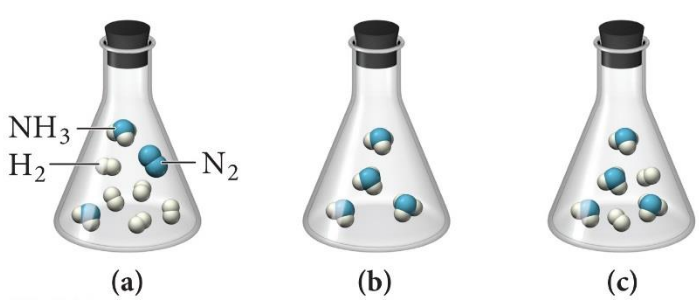
Examples:

1. Ammonia can be produced from the reaction shown below:

$$N_2\,(g) + 3\,H_2\,(g) \rightarrow 2\,NH_3\,(g)$$

If the flask on the right represents the mixture of reactants present at the start, which diagram below represents the flask after the reaction has completely reacted?





LIMITING REACTANT, THEORETICAL YIELD & PERCENT YIELD

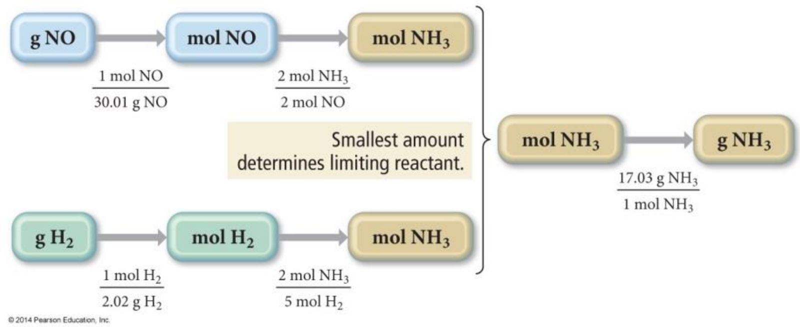
Examples (cont'd)

2. Ammonia can be synthesized by the reaction shown below:

$$2\ NO\ (g) + 5\ H_2\ (g) \rightarrow 2\ NH_3\ (g) + 2\ H_2O\ (g)$$

a) If 86.3 g of NO are combined with 25.6 g of H_2, what is the theoretical yield for this reaction?

Method 1:



b) After the reaction above is completed, how much of the excess reagent remains unreacted? How much additional limiting reactant is needed to completely react with this excess reagent?

Method 2:

LIMITING REACTANT, THEORETICAL YIELD & PERCENT YIELD

Examples (cont'd)

3. Mining companies use the reaction shown below to obtain iron from iron ore:

$$Fe_2O_3\ (s)\ +\ 3\ CO\ (g)\ \rightarrow\ 2\ Fe\ (s)\ +\ 3\ CO_2\ (g)$$

The reaction of 167 g of Fe_2O_3 with 85.8 g of CO produces 72.3 g of Fe. Determine the theoretical and percent yield for this reaction.

4. Elemental phosphorous reacts with chlorine gas according to the equation:

$$P_4\ (s)\ +\ 6\ Cl_2\ (g)\ \rightarrow\ 4\ PCl_3\ (l)$$

A reaction mixture initially contains 45.69 g of P_4 and 131.3 g of Cl_2. Once the reaction has completed, what is the identity of the excess reactant and what mass (in g) of it is left?

SOLUTION CONCENTRATION

- Many chemical reactions involve reactants dissolved in water. A homogeneous mixture of two substances–for example, salt in water– is called a ***solution***.

- The majority component of a solution is ***solvent***, and the minority component of a solution is ***solute.*** An ***aqueous*** solution is one in which water acts as the solvent.

- The amount of solute in a solution is variable and is referred to as ***concentration***. The most common unit of concentration in chemistry is ***molarity***. Molarity is defined as:

$$\text{Molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

- Molarity of a solution can also be used as a conversion factor between moles of solute and volume of solution.

Examples:

1. What mass of KBr (in g) is required to make 250.0 mL of a 1.50 M KBr solution?

2. How many grams of sucrose ($C_{12}H_{22}O_{11}$) are in 1.55L of 0.758 M sucrose solution?

3. Determine the percent by mass of NaCl in a 1.35 M NaCl solution. The density of the solution is 1.05 g/mL.

CONCENTRATION OF IONS IN SOLUTION

- When dealing with substances that form ions in solutions, the concentration of ions in solution depends on the mole ratio between the dissolved substance and the cations and the anions it forms in solution.

- For example:

  in a 1.0 M NaCl solution $[Na^+] = 1.0$ M and $[Cl^-] = 1.0$ M

  whereas,

  in a 1.0 M $CaCl_2$ solution $[Ca^{2+}] = 1.0$ M and $[Cl^-] = 2.0$ M

Examples:

1. What is the concentration of Cl^- (in M) in each solution below:

   a) 0.200 M NaCl ____________

   b) 0.100 M $AlCl_3$ ____________

   c) 0.150 M $SrCl_2$ ____________

2. A solution is prepared by dissolving 25 g of aluminum nitrate in enough water to make 150 mL of solution. Identify the ions present in solution and the concentration of each.

SOLUTION DILUTION

- Many solutions in the laboratory are prepared from more concentrated solutions, called ***stock*** solution. This process is called ***dilution***.

- To determine the amount of stock solution required to prepare a more dilute solution, the dilution equation shown below is used:

$$M_1\ V_1 = M_2\ V_2$$

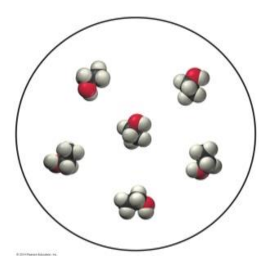
- This relationship is valid because the product of molarity times volume on each side equals the moles of solute, which remains constant during dilution. Molarity and volume, however, are inversely proportional during the dilution process.

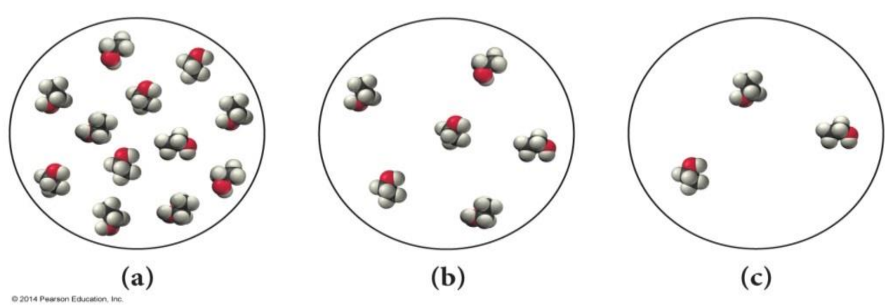
Examples:

1. What is the final volume of a solution prepared by diluting 100.0 mL of 5.00 M $CaCl_2$ solution to obtain a 0.750 M solution?

2. What volume of water (in L) must be added to 50.0 mL of 12 M stock HNO_3 solution to obtain a 0.100 M acid solution?

3. The diagram on the right represents a small volume within 500 mL of an aqueous ethanol solution. (Water molecules have been omitted for clarity). Which image below represents the same volume of solution after an additional 500 mL of water have been added?

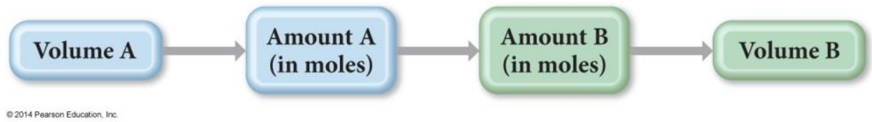




SOLUTION STOICHIOMETRY

- In aqueous solutions, the quantities of reactants and products are often specified in terms of volume and concentration, rather than mass. We can use the volume and concentration of a reactant or product to calculate amount in moles.

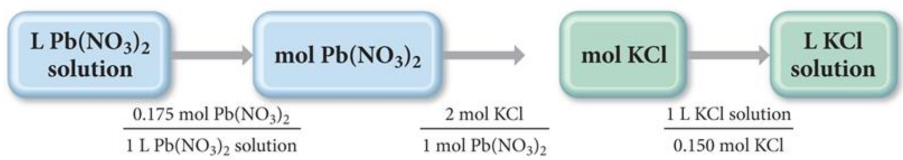
- The general conceptual plan shown below is then used to solve stoichiometric problems with solutions:



Examples:

1. What volume (in L) of 0.150 M KCl solution will completely react with 0.150 L of 0.175 M $Pb(NO_3)_2$ solution according to the balanced equation below:

$$2\ KCl\ (aq)\ +\ Pb(NO_3)_2\ (aq)\ \rightarrow\ PbCl_2\ (s)\ +\ 2\ KNO_3\ (aq)$$

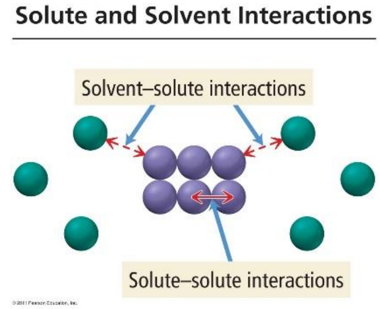


2. A 55.0-mL sample of 0.102 M K_2SO_4 is mixed with 35.0 mL of 0.114 M $Pb(NO_3)_2$ solution, according to the reaction below:

$$K_2SO_4\ (aq)\ +\ Pb(NO_3)_2\ (aq)\ \rightarrow\ PbSO_4\ (s)\ +\ 2\ KNO_3\ (aq)$$

The solid $PbSO_4$ is collected, dried and found to have a mass of 1.01 g. What is the percent yield of this reaction?

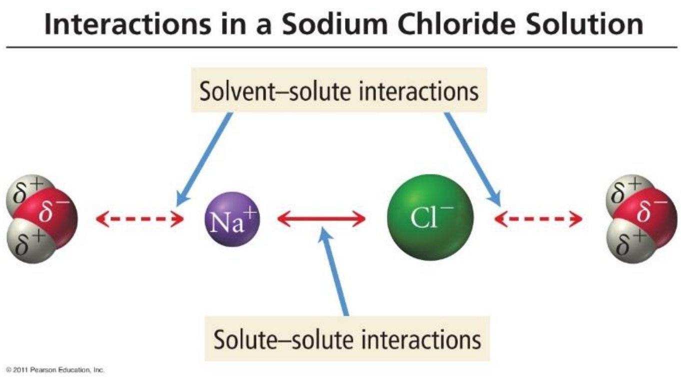
AQUEOUS SOLUTIONS & SOLUBILITY



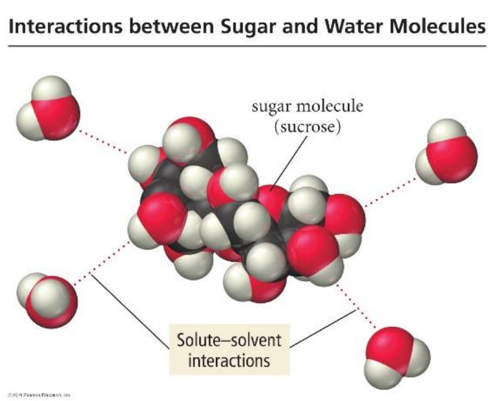
- When a solid is placed into a liquid solvent, the attractive forces that hold the solid together (solute–solute interactions) compete with the attractive forces between solvent particles and the particles of the solid (solvent–solute interaction).

- When the solvent–solute interactions overcome the solute–solute interactions, the solid dissolves in the solvent.

- For example, when NaCl is placed in water, there is a competition between the attraction of Na^+ and Cl^- ions to each other, and the attraction between the Na^+ and Cl^- ions to water molecules. The attraction between the separated ions and the polar water molecule is stronger than the attractions of the ions to each other. As a result, NaCl dissolves in water.



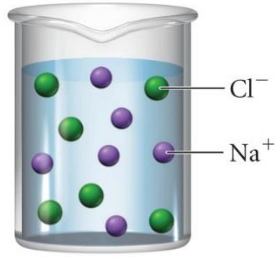
- Similarly, when sugar is added to water, the attraction of the partially charged sugar molecules to the polar water molecule is stronger than the attraction of the sugar molecules to each other. As a result, sugar dissolves in water.



ELECTROLYTES & NONELECTROLYTES

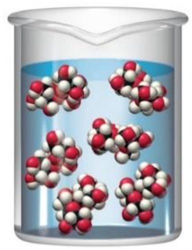
- When salt (an ionic compound) dissolves in water, it produces a conducting solution, whereas when sugar (a molecular compound) dissolves in water, it produces a non-conducting solution.

- This difference is due to the manner in which ionic and molecular substances dissolve in water, and is a fundamental difference between types of solutions.



$NaCl(aq)$

- When soluble ionic solids dissolve in water, they dissociate into ions and produce conducting solutions. These solutions are called ***electrolytes.*** Solutions that dissociate completely into ions are called ***strong electrolytes***.

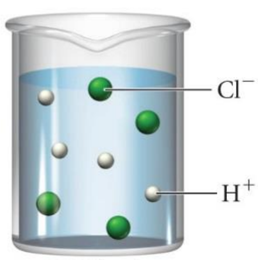


$C_{12}H_{22}O_{11}(aq)$

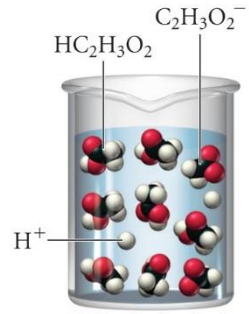
- In contrast, when most molecular compounds dissolve in water, they remain intact and do not dissociate into ions. As a result, they produce non-conducting solutions. These solutions are called ***nonelectrolytes***.

- Acids, are molecular substances that ionize–form ions– when dissolved in water. For example, HCl is a molecular compound that dissociates into H^+ and Cl^- when it dissolves in water. HCl is a ***strong acid***, since it ionizes completely, and is classified a ***strong electrolyte***.

- Many acids are ***weak acids,*** since they do not completely ionize, and only form a few ions. Weak acids are classified as ***weak electrolytes***.



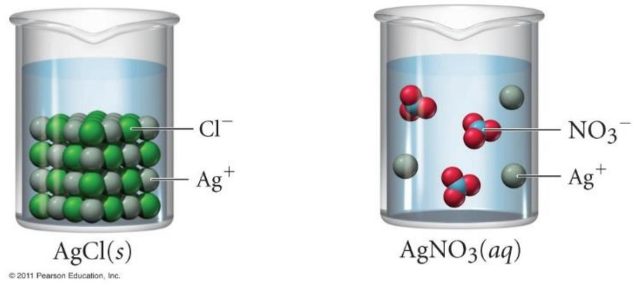
$HCl(aq)$



$HC_2H_3O_2(aq)$

SOLUBILITY OF IONIC COMPOUNDS

- Many ionic solids ***dissolve*** in water and are called ***soluble salts***. However, some ionic solids ***do not dissolve*** in water and do not form ions in solution. These salts are called ***insoluble salts*** and remain solid in solution.



- Chemists use a set of ***solubility rules*** to ***predict*** whether a salt is ***soluble or insoluble***. These rules are summarized below:

TABLE 4.1 Solubility Rules for Ionic Compounds in Water

Compounds Containing the Following Ions Are Generally Soluble	Exceptions
Li^+, Na^+, K^+, and NH_4^+	None
NO_3^- and $C_2H_3O_2^-$	None
Cl^-, Br^-, and I^-	When these ions pair with Ag^+, Hg_2^{2+} or Pb^{2+}, the resulting compounds are insoluble.
SO_4^{2-}	When SO_4^{2-} pairs with Sr^{2+}, Ba^{2+}, Pb^{2+}, Ag^+, or Ca^{2+}, the resulting compound is insoluble.
Compounds Containing the Following Ions Are Generally Insoluble	**Exceptions**
OH^- and S^{2-}	When these ions pair with Li^+, Na^+, K^+, or NH_4^+, the resulting compounds are soluble.
	When S^{2-} pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the resulting compound is soluble.
	When OH^- pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the resulting compound is slightly soluble.
CO_3^{2-} and PO_4^{3-}	When these ions pair with Li^+, Na^+, K^+, or NH_4^+, the resulting compounds are soluble.

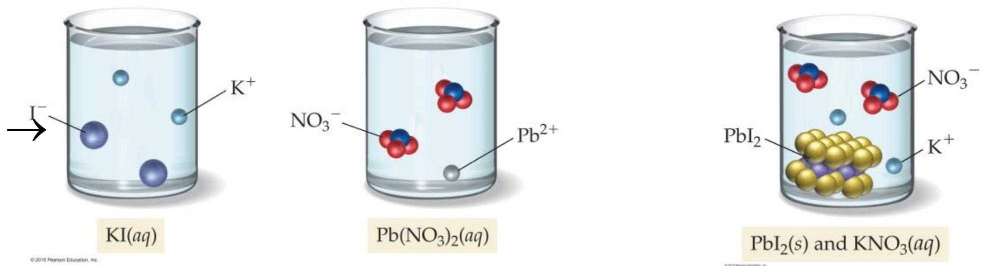
PRECIPITATION REACTIONS

- Solubility rules can be used to predict whether a solid, called a ***precipitate***, can be formed when two solutions of ionic compounds are mixed.

- A solid is formed when two ions of an insoluble salt come in contact with one another.

- For example, when a solution of $Pb(NO_3)_2$ is mixed with a solution of KI , a yellow insoluble salt PbI_2 is produced.

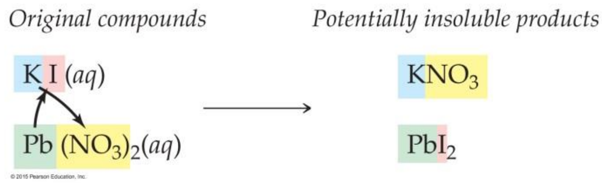
$$\mathbf{2\ KI\ (aq) + Pb(NO_3)_2\ (aq) \rightarrow \underset{\text{precipitate}}{PbI_2\ (s)} + 2\ KNO_3\ (aq)}$$



- Double replacement reactions in which a precipitate is formed are called ***precipitation*** reactions.

- The solubility rules can be used to predict whether a precipitate forms when two solutions of ionic compounds are mixed together.

- For example, when solutions of KI and $Pb(NO_3)_2$ are mixed together, two potentially insoluble products are formed (KNO_3 and PbI_2).



- If the potentially insoluble products are both soluble, then no reaction occurs. If, on the other hand, one of these products is insoluble, then a precipitation reaction occurs.

PRECIPITATION REACTIONS

Examples:

1. Write an equation for the precipitation reaction that occurs (if any) when solutions of potassium carbonate and nickel (II) chloride are mixed.

2. Write an equation for the precipitation reaction that occurs (if any) when solutions of sodium nitrate and lithium sulfate are mixed.

MOLECULAR, COMPLETE IONIC & NET IONIC EQUATIONS

- When writing equations for precipitation reactions, the equation is usually written as a ***molecular equation***, showing each compound in the reaction as a molecule.

$$AgNO_3\,(aq) + NaCl\,(aq) \rightarrow AgCl\,(s) + NaNO_3\,(aq)$$

- This equation can also be written in a way to show the aqueous solutions as they really exist as ions, and is called ***complete ionic equation***.

$$Ag^+\,(aq) + NO_3^-\,(aq) + Na^+\,(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s) + Na^+\,(aq) + NO_3^-\,(aq)$$

- In the equation above, notice that some of the ions appear in the same form on the reactant and product side. These ions do not participate in the reaction and are called ***spectator ions***.

- The complete ionic equation can be simplified by omitting the spectator ions. The resulting equation is called ***net ionic equation***.

$$Ag^+\,(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s)$$

- To summarize:

  - A ***molecular equation*** is a chemical equation showing the complete, neutral formulas for every compound in the reaction.
  - A ***complete ionic equation*** is a chemical equation showing all the species as they are actually present in solution.
  - A ***net ionic equation*** is an equation showing only the species that actually participate in the reaction.

NET IONIC EQUATIONS

Examples:

1. Write complete and net ionic equations for each reaction shown below:

   a) $3\ SrCl_2\ (aq) + 2\ Li_3PO_4\ (aq) \rightarrow Sr_3\ (PO_4)_2\ (s) + 6\ LiCl\ (aq)$

   b) $HC_2H_3O_2\ (aq) + KOH\ (aq) \rightarrow KC_2H_3O_2\ (aq) + H_2O\ (l)$

   c) $2\ HI\ (aq) + Ba(OH)_2\ (aq) \rightarrow BaI_2\ (aq) + 2\ H_2O\ (l)$

2. A solution contains the following ions: Ba^{2+}, Pb^{2+} and Mg^{2+}. What substances do you need to add to this solution in order to separate each ion? (Include NIE equations for each step)

ACID-BASE REACTIONS

- Two other important classes of reactions that occur in aqueous solutions are acid-base reactions and gas-evolution reactions.

Acid-Base Reactions:

- Recall from earlier that acids are substances that form H^+ ion in solution and bases are substances that form OH^- solution. These definitions are called the Arrhenius definitions.

- Some acids–called polyprotic acids–contains more that none ionizable hydrogen and release them sequentially. For example, H_2SO_4 is a polyprotic acid. It is strong when losing its first proton, and weak when losing the second.

$$H_2SO_4\,(aq) \rightarrow H^+\,(aq) + HSO_4^-\,(aq)$$

$$HSO_4^-\,(aq) \rightleftharpoons H^+\,(aq) + SO_4^{2-}\,(aq)$$

- The most important reaction of acids and bases is called **neutralization**. In these reactions an acid combines with a base to form a **salt and water**. For example:

$$\underset{\text{acid}}{HCl\,(aq)} + \underset{\text{base}}{NaOH\,(aq)} \longrightarrow \underset{\text{salt}}{NaCl\,(aq)} + \underset{\text{water}}{H_2O\,(l)}$$

- ***Salts*** are ***ionic*** substances with the ***cation*** donated from the ***base*** and the ***anion*** donated from the ***acid***. In the laboratory, neutralization reactions are observed by an increase in temperature (exothermic reaction).

- Listed below are some common acids and bases used in the laboratory.

TABLE 4.2 Some Common Acids and Bases

Name of Acid	Formula	Name of Base	Formula
Hydrochloric acid	HCl	Sodium hydroxide	$NaOH$
Hydrobromic acid	HBr	Lithium hydroxide	$LiOH$
Hydroiodic acid	HI	Potassium hydroxide	KOH
Nitric acid	HNO_3	Calcium hydroxide	$Ca(OH)_2$
Sulfuric acid	H_2SO_4	Barium hydroxide	$Ba(OH)_2$
Perchloric acid	$HClO_4$	Ammonia*	NH_3 (weak base)
Acetic acid	$HC_2H_3O_2$ (weak acid)		
Hydrofluoric acid	HF (weak acid)		

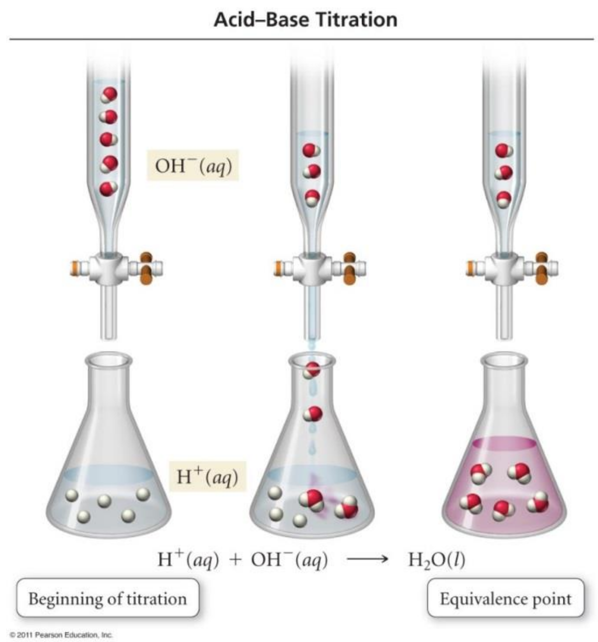
*Ammonia does not contain OH^-, but it produces OH^- in a reaction with water that occurs only to a small extent: $NH_3\,(aq) + H_2O(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$.

ACID-BASE TITRATIONS

- Principles of acid-base neutralization and stoichiometry can be applied to a common laboratory technique called a titration.

- In a titration, a substance of known concentration is reacted with another substance of unknown concentration. For example, consider the acid-base reaction between HCl and NaOH shown below:

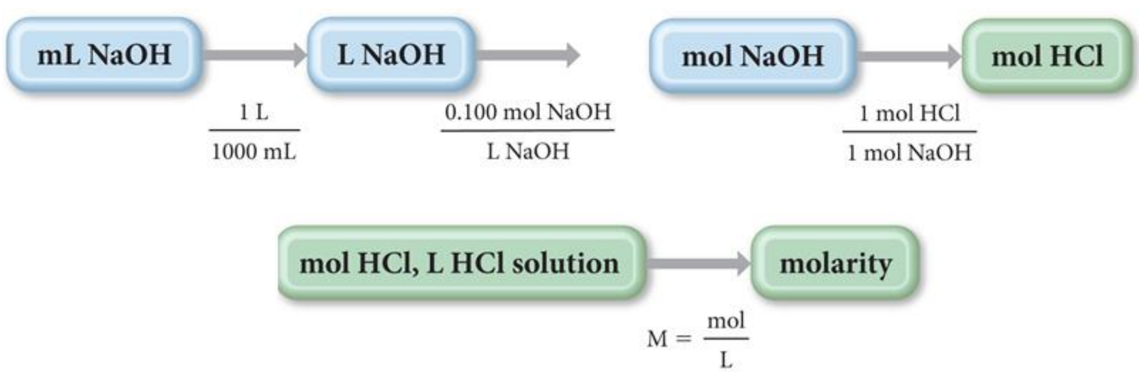
  $HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$



- When a solution of NaOH with a known concentration is added slowly to a known volume of HCl with an unknown concentration, the H^+ and OH^- ions combine to form water. (Note that the Na^+ and Cl^- ions forming NaCl are omitted from this discussion, since they are spectator ions).

- After addition of enough OH^- to neutralize all of the H^+ present (***equivalence point***), the solution becomes neutral and the titration is complete. The equivalence point is usually detected by addition of an ***indicator***, a dye that changes color based on the acidity or basicity of the solution.

- Stoichiometric calculations based on concentration of know substance (NaOH) and volume of both solutions can yield the concentration of the unknown substance (HCl).



ACID-BASE TITRATIONS

Examples:

1. The titration of a 10.00 mL sample of an HCl solution of unknown concentration requires 12.54 mL of a 0.100 M NaOH solution to reach the equivalence point. What is the concentration of the unknown HCl solution?

2. The titration of a 20.00 mL sample of an H_2SO_4 solution of unknown concentration requires 22.87 mL of a 0.158 M KOH solution to reach the equivalence point. What is the concentration of the unknown H_2SO_4 solution?

3. What volume (in mL) of 0.200 M NaOH solution is required to titrate 35.00 mL of 0.140 M HBr solution to the equivalence points?

GAS-EVOLUTION REACTIONS

- Some chemical reactions ***produce gas*** because one of the products formed in the reaction is ***unstable***. Three such products are listed below:

Carbonic acid	$H_2CO_3\ (aq) \rightarrow CO_2\ (g) + H_2O\ (l)$
Sulfurous acid	$H_2SO_3\ (aq) \rightarrow SO_2\ (g) + H_2O\ (l)$
Ammonium hydroxide	$NH_4OH\ (aq) \rightarrow NH_3\ (g) + H_2O\ (l)$

- When any of these products appears in a chemical reaction, they should be replaced with their decomposition products.

$$2\ HCl\ (aq) + Na_2CO_3\ (aq) \rightarrow 2\ NaCl\ (aq) + \underset{\textbf{(unstable)}}{H_2CO_3\ (aq)}$$

$$2\ HCl\ (aq) + Na_2CO_3\ (aq) \rightarrow 2\ NaCl\ (aq) + CO_2\ (g) + H_2O\ (l)$$

Examples:

1. Write molecular equation for the reaction that occurs when you mix aqueous nitric acid and sodium carbonate.

2. Write molecular and net ionic equations for the reaction that occurs when you mix aqueous hydrobromic acid and aqueous potassium sulfite.

OXIDATION-REDUCTION (REDOX) REACTIONS

- Some of the most important reaction in chemistry are ***oxidation-reduction (redox)*** reactions. In these reactions, electrons transfer from one reactant to the other.

- The rusting of iron, and combustion of octane in automobiles are two examples of redox reactions.

$$4\ Fe\ (s) + 3\ O_2\ (g) \rightarrow 2\ Fe_2O_3\ (s)$$

$$2\ C_8H_{18}\ (l) + 25\ O_2\ (g) \rightarrow 16\ CO_2\ (g) + 18\ H_2O\ (g)$$

- Many redox reactions involve reaction of a substance with oxygen. However, not all redox reactions need to involve oxygen. For example, reaction of sodium and chlorine to form sodium chloride is also a redox reaction.

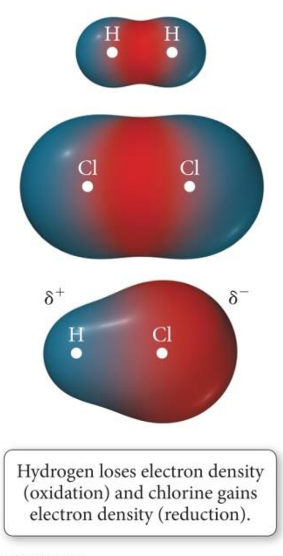
$$2\ Na\ (s) + Cl_2\ (g) \rightarrow 2\ NaCl\ (s)$$

- The transfer of electrons in redox reactions need not be a complete transfer (as occurs in the formation of an ionic compound). For example, consider the reaction between hydrogen gas and chlorine gas to form HCl.

$$H_2\ (g) + Cl_2\ (g) \rightarrow 2\ HCl\ (g)$$

- Even though HCl is a molecular compounds with a covalent bond , and even though hydrogen has not transferred its electrons to chlorine, hydrogen has lost some of its electron density due to the greater electronegativity of chlorine.

- In redox reactions, loss of electron is defined as ***oxidation***, and gain of electrons is defined as ***reduction***. For example, in the reaction of hydrogen and chlorine, since hydrogen has lost electron density, it is oxidized, and since chlorine has gained electron density, it is reduced.



OXIDATION NUMBERS OR STATES

- Identifying the oxidation-reduction nature of reactions between metals and non-metals is straight forward because of ion formation. However, redox reactions that occur between two non-metals are more difficult to characterize since no ions are formed.

- Chemists have devised a scheme to track electrons before and after a reaction in order to simplify this process. In this scheme, a number (oxidation state or number) is assigned to each element assuming that the shared electrons between two atoms belong to the one with the most attraction for these electrons. The oxidation number of an atom can be thought of as the "charge" the atom would have if all shared electrons were assigned to it.

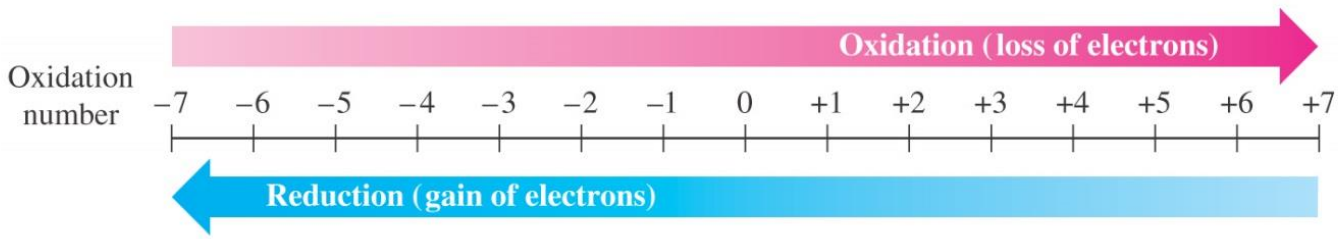
- The following rules are used to assign oxidation numbers to atoms in elements and compounds. (Note: these rules are hierarchical. If any two rules conflict, follow the rule that is higher on the list)

1. The oxidation number of an atom in a free element is 0.
2. The oxidation number of a monatomic ion is equal to its charge.
3. The sum of the oxidation number of all atoms in:
    - A neutral molecule or formula is equal to 0.
    - An ion is equal to the charge of the ion.
4. In their compounds, metals have a positive oxidation number.
    a. Group 1A metals always have an oxidation number of +1.
    b. Group 2A metals always have an oxidation number of +2.
5. In their compounds non-metals are assigned oxidation numbers according to the table at right. Entries at the top of the table take precedence over entries at the bottom of the table.

Nonmetal	Oxidation State	Example
Fluorine	–1	MgF_2 –1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	–2	CO_2 –2 ox state
Group 7A	–1	CCl_4 –1 ox state
Group 6A	–2	H_2S –2 ox state
Group 5A	–3	NH_3 –3 ox state

OXIDIDATION-REDUCTION REACTIONS

- Oxidation numbers (or states) can be used to identify redox reactions and determine which substance is oxidized and which is reduced.
- To do so, assign oxidation numbers to all elements in the reactants and products, then track which elements change oxidation numbers from reactants to products.
- Elements that increase their oxidation numbers lose electrons, and are therefore oxidized. Elements that decrease their oxidation numbers gain electrons, and are therefore reduced.



Examples:

1. Assign oxidation number of each element underlined in the substances below:

   a) $K\underline{F}$ __________ c) $\underline{C}O_2$__________

   b) $\underline{S}O_4^{2-}$ __________ d) $K_2\underline{O}_2$ ___________

2. For each reaction below, use oxidation numbers to determine which element is oxidized and which is reduced and the number of electrons transferred in the reaction:

   a) $Mg\ (s)\ +\ 2\ H_2O\ (l)\ \rightarrow\ Mg(OH)_2\ (aq)\ +\ H_2\ (g)$

   b) $Sn\ (s)\ +\ 4\ HNO_3\ (aq)\ \rightarrow\ SnO_2\ (s) + 4\ NO_2\ (g)\ +\ 2\ H_2O\ (g)$

OXIDIZING & REDUCING AGENTS

- Oxidation and reduction reactions occur simultaneously. If one substance loses electrons (oxidation) then another substance must gain electrons (reduction).
- A substance that causes oxidation of another substance is called ***oxidizing agent***. In a redox reaction, the oxidizing agent is always reduced.
- A substance that causes reduction of another substance is called ***reducing agent***. In a redox reaction, the reducing agent is always oxidized.

Examples:
Determine whether each reaction below is a redox reaction. If yes, identify the oxidizing and reducing agents.

1) $2\,Mg\,(s) + O_2\,(g) \rightarrow 2\,MgO\,(s)$

2) $2\,HBr\,(aq) + Ca(OH)_2\,(aq) \rightarrow CaBr_2\,(aq) + 2\,H_2O\,(g)$

3) $Zn\,(s) + Fe^{2+}\,(aq) \rightarrow Zn^{2+}\,(aq) + Fe\,(s)$

OXIDATION & REDUCTION HALF-REACTIONS

3. The overall oxidation-reduction reaction can be thought of as the sum of an ***oxidation "half-reaction"*** and a ***reduction "half-reaction"***, each of which involves either the loss or the gain of electrons.

4. For example, consider the reaction of zinc with hydrochloric acid:

$$Zn\ (s) + 2\ HCl\ (aq) \rightarrow ZnCl_2\ (aq) + H_2\ (g)$$

Eliminating Cl^- as the spectator ion, the net ionic equation is:

$$Zn\ (s) + 2\ H^+ \rightarrow Zn^{2+} + H_2\ (g)$$

5. This redox reaction can then be written as two separate half-reactions:

$$Zn\ (s) \rightarrow Zn^{2+} + 2\ e^- \quad \text{(oxidation ½-reaction)}$$

$$2\ H^+ + 2\ e^- \rightarrow H_2\ (g) \quad \text{(reduction ½-reaction)}$$

6. Note that in oxidation half-reaction electrons are written as a product, indicating loss, while in reduction half-reaction, electrons are written as a reactant, indicating gain.

7. When the two half-reactions are added together, the electrons cancel each other and the overall reaction between Zn and H^+ is obtained.

Examples:

1. Identify each half-reaction below as oxidation or reduction, and complete by placing the proper number of electrons on the proper side:

a) $Cl_2\ (g) \rightarrow 2\ Cl^-$ ____________________________

b) $Ca\ (s) \rightarrow Ca^{2+}$ ____________________________

c) $Fe^{3+} \rightarrow Fe^{2+}$ ____________________________

d) $2\ Br^- \rightarrow Br_2\ (l)$ ____________________________

BALANCING REDOX REACTIONS BY HALF-REACTION METHOD

8. Some redox reactions can be difficult to balance by the traditional inspection method. These reactions are more easily balanced by alternate methods.

9. One such method is the half-reaction method. To balance equations by this method, follow the steps below:

    1. Assign oxidation numbers to all elements in each half-reaction.
    2. From the changes in oxidation numbers, identify elements oxidized and reduced.
    3. Determine the number of electrons lost in oxidation and gained in reduction from the oxidation number changes.
    4. Multiply one or both of these numbers by the appropriate factors to make the electrons lost and gained equal, and use the factor as balancing coefficients for the appropriate substances.
    5. Complete the balancing of each half-reaction by inspection.
    6. Add the two balanced half-reactions by cancelling the electrons and other similar substances to obtain the overall equation.
    7. Check the final equation for balance of atoms and charges.

Examples:

Balance each half-reaction shown below and use to obtain the balanced overall equation:

1) $Fe^{2+} \rightarrow Fe^{3+}$

$H^{+} + MnO_4^{-} \rightarrow Mn^{2+} + H_2O$

BALANCING REDOX REACTIONS BY OXIDATION NUMBER METHOD

Examples (cont'd):

2) $Cr_2O_7^{2-} + H^+ \rightarrow Cr^{3+} + H_2O$

$I^- \rightarrow I_2$

3) $H^+ + BrO_3^- \rightarrow Br^- + H_2O$

$N_2H_4 \rightarrow N_2 + H^+$

Answers to In-Chapter Problems:

Page	***Example No.***	***Answer***
2	1	25.8 g
3	2	22 kg
	3	c
4	1	c
5	2	49.0 g
6	3	TY=114 g; % yield = 63.4
	4	Excess reactant is phosphorus; after reaction 7.45 g is left
7	1	44.6 g
	2	402 g
	3	7.52%
8	1	667 mL
	2	5.95 L
	3	c
9	1	0.350 L
	2	83.5%
14	1	K_2CO_3 (aq) + $NiCl_2$ (aq) $\rightarrow$ $NiCO_3$ (s) + 2 KCl (aq)
	2	No reaction (both possible products are soluble)
16	1	3 Sr^{2+} + 6 Cl^- + 6 Li^+ + 2 PO_4^{3-} $\rightarrow$ $Sr_3(PO_4)_2$ (s) + 6 Li^+ + 6 Cl^- 3 Sr^{2+} + 2 PO_4^{3-} $\rightarrow$ $Sr_3(PO_4)_2$ (NIE)
	2	$HC_2H_3O_2$ + K^+ + OH^- $\rightarrow$ K^+ + $C_2H_3O_2^-$ + H_2O $HC_2H_3O_2$ + OH^- $\rightarrow$ $C_2H_3O_2^-$ + H_2O (NIE)
	3	2H^+ + 2 I^- + Ba^{2+} + 2 OH^- $\rightarrow$ Ba^{2+} + 2I^- + 2 H_2O H^+ + OH^- $\rightarrow$ H_2O (NIE)
19	1	0.1254 M
	2	0.0904 M
	3	24.5 mL
20	1	2 HNO_3 (aq) + Na_2CO_3 (aq) $\rightarrow$ 2 $NaNO_3$ (aq) + CO_2 (g) + H_2O (l)
	2	2 HBr (aq) + K_2SO_3 (aq) $\rightarrow$ 2 KBr (aq) + SO_2 (g) + H_2O (l) 2 H^+ + SO_3^{2-+} $\rightarrow$ SO_2 (g) + H_2O (l) (NIE)

Page	***Example No.***	***Answer***
23	1	a) –1 b) +6 c) +4 d) –1
	2a	Mg is oxidized; H is reduced; 2 electrons transferred
	2b	Sn is oxidized; N is reduced; 4 electrons transferred
24	1	Yes; Mg is reducing agent (oxidized); O_2 is oxidizing agent (reduced)
	2	No, not a redox reaction
	3	Yes; Zn is reducing agent (oxidized); Fe^{2+} is oxidizing agent (reduced)
25	1a	Reduction: $Cl_2 + 2\ e^- \rightarrow 2\ Cl^-$
	1b	Oxidation: $Ca\ (s) \rightarrow Ca^{2+} + 2\ e^-$
	1c	Reduction: $Fe^{3+} + e^- \rightarrow Fe^{2+}$
	1d	Oxidation: $2\ Br^- \rightarrow Br_2\ (l) + 2\ e^-$
26	1	$8\ H^+ + 5\ Fe^{2+} + MnO_4^- \rightarrow 5\ Fe^{3+} + Mn^{2+} + 4\ H_2O$
27	2	$Cr_2O_7^{2-} + 6\ I^- + 14\ H^+ \rightarrow 2\ Cr^{3+} + 3\ I_2 + 7\ H_2O$
	3	$2\ BrO_3^- + 3\ N_2H_4 \rightarrow 2\ Br^- + 3\ N_2 + 6\ H_2O$