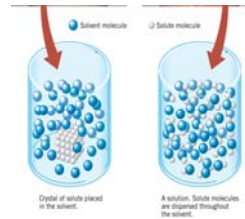


Highlights of Chapter 4:

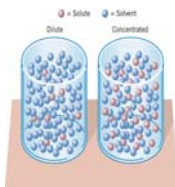
- A **solution** is a homogeneous mixture in which the two or more components mix freely
- The **solvent** is taken as the component present in the largest amount
- A **solute** is any substance dissolved in the solvent



Formation of a solution of iodine molecules in ethyl alcohol. Ethyl alcohol is the *solvent* and iodine the *solute*.

Solutions have variable composition. They may be characterized using a solute-to-solvent ratio called the **concentration**.

- The *relative* amounts of solute and solvent are often given without specifying the actual quantities



The **dilute solution** on the left has less solute per unit volume than the (more) **concentrated solution** on the right.

Concentrated and dilute are *relative* terms.

- There is usually a limit to the amount of solute that can dissolve in a given amount of solvent
 - For example, 36.0 g NaCl is able to dissolve in 100 g of water at 20°C
- A solution is said to be **saturated** when no more solute can be dissolved at the current temperature
- The **solubility** of a solute is the number of grams of *solute* that can dissolve in 100 grams of *solvent* at a given temperature

- Solubilities of some common substances

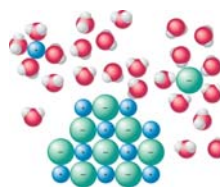
Substance	Formula	Solubility (g/100 g water)
Sodium chloride	NaCl	35.7 at 0°C
		39.1 at 100°C
Sodium hydroxide	NaOH	42 at 0°C
		347 at 100°C
Calcium carbonate	CaCO ₃	0.0015 at 25°C

A solution containing less solute is called **unsaturated** because it is able to dissolve more solute.

- Solubility usually increases with temperature
- **Supersaturated solutions** contain more solute than required for saturation at a given temperature
- They can be formed, for example, by careful cooling of saturated solutions
- Supersaturated solutions are unstable and often result in the formation of a **precipitate**

- A precipitate is the solid substance that separates from solution
- Precipitates can also form from reactions
- Reactions that produce a precipitate are called **precipitation reactions**
- Many ionic compounds dissolve in water
- Solutes that produce ions in solution are called **electrolytes** because their solutions can conduct electricity

- An ionic compounds **dissociates** as it dissolves in water

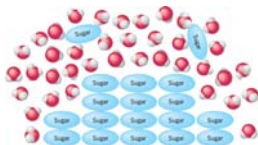


Ions separate from the solid and become **hydrated** or surrounded by water molecules.

The ions move freely and the solution is able to conduct electricity.

Ionic compounds that dissolve completely are **strong electrolytes**

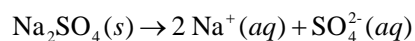
- Most solutions of molecular compounds do not conduct electricity and are called **nonelectrolytes**



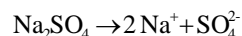
The molecules of a nonelectrolyte separate but stay intact. The solution is nonconducting because no ions are generated.

Some ionic compounds have low solubilities in water but are still strong electrolytes because what does dissolve is 100% dissociated.

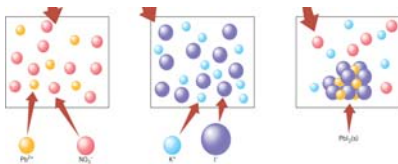
- The dissociation of ionic compounds may be described with chemical equations



- The hydrated ions, with the symbol (aq) , have been written separately
- Since physical states are often omitted, you might encounter the equation as:

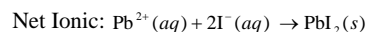
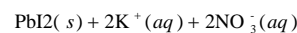
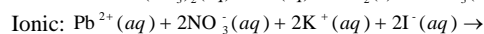
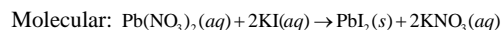


- Ionic compounds often react when their aqueous solutions combine



When a solution of $\text{Pb}(\text{NO}_3)_2$ is mixed with a solution of KI the yellow precipitate PbI_2 rapidly forms.

- This reaction may be represented with a **molecular, ionic, or net ionic equation:**



- The most compact notation is the net ionic equation which eliminates all the non-reacting **spectator ions** from the equation

- Criteria for balanced ionic and net ionic equations:

- 1) Material balance – the same number of each type of atom on each side of the arrow
- 2) Electrical balance – the *net* electrical charge on the left side of the arrow must equal the *net* electrical charge on the right side of the arrow

Remember that the charge on an ion must be included when it is not in a compound. Adding the charges on all the ions on one side of the arrow gives the net electrical charge.

- In the reaction of $\text{Pb}(\text{NO}_3)_2$ with KI the cations and anions changed partners
- This is an example of a “**swapping partners**” or double **replacement reaction**
- **Solubility rules** allows the prediction of when a precipitation reaction will occur
- For many ionic compounds the solubility rules correctly predict whether the ionic compound is soluble or insoluble

- Solubility rules for ionic compounds in water:

– Soluble Compounds

- 1) All compounds of the alkali metals (Group IA) are soluble.
- 2) All salts containing NH_4^+ , NO_3^- , ClO_4^- , ClO_3^- , and $\text{C}_2\text{H}_3\text{O}_2^-$ are soluble.
- 3) All salts containing Cl^- , Br^- , or I^- are soluble *except* when combined with Ag^+ , Pb^{2+} , and Hg_2^{2+} .
- 4) All sulfates are soluble *except* those of Pb^{2+} , Ca^{2+} , Sr^{2+} , Hg_2^{2+} , and Ba^{2+} .

– Insoluble compounds

- 5) All metal hydroxides and oxides are insoluble *except* those of Group IA and of Ca^{2+} , Sr^{2+} , and Ba^{2+} . When metal oxides do dissolve, they react with water to form hydroxides. The oxide ion, O^{2-} , does not exist in water.
 - 6) All salts that contain PO_4^{3-} , CO_3^{2-} , SO_3^{2-} , and S^{2-} are insoluble, *except* those of Group IA and NH_4^+ .
- A knowledge of these rules will allow you to predict a large number of precipitation reactions

- **Acids** and **bases** are another important class of compounds
- Acids and bases affect the color of certain natural dye substances
- They are called **acid-base indicators** because they indicate the presence of acids or bases with their color
- The first comprehensive theory of acids, bases, and electrical conductivity appeared in 1884 in the Ph.D. thesis of Savante Arrhenius

- He proposed that acids form hydrogen ions and bases released hydroxide ions in solution
- The characteristic reaction between acids and bases is neutralization

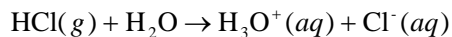
$$\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)$$
- In general, the reaction of an acid and a base produces water and a **salt**
- We can state the Arrhenius definition of acids and bases in updated form

Arrhenius Definition of Acids and Bases

An *acid* is a substance that reacts with water to produce the hydronium ion, H_3O^+ .

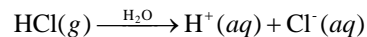
A *base* is a substance that produces hydroxide ion in water.

- In general, acids are molecular compounds that react with water to produce ions
- This is called **ionization**:



- It is common to encounter the hydrogen ion (H^+) instead of the hydronium ion

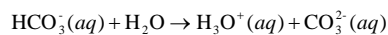
- The previous ionization is also written as



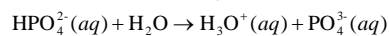
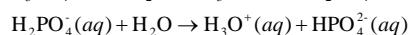
- **Monoprotic** acids are capable of furnishing only one hydrogen ion per molecule
- Acids that can furnish more than one hydrogen ion per molecule are called **polyprotic** acids

Monoprotic : $\text{HCl}(aq) + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)$

Diprotic : $\text{H}_2\text{CO}_3(aq) + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+(aq) + \text{HCO}_3^-(aq)$



Triprotic : $\text{H}_3\text{PO}_4(aq) + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+(aq) + \text{H}_2\text{PO}_4^-(aq)$



- Binary compounds of many nonmetals and hydrogen are acidic

- In water solution these are referred to as **binary acids**

- They are named by adding the prefix *hydro-* and the suffix *-ic* to the stem of the nonmetal name, followed by the word *acid*

Molecular Compound	Binary Acid
$\text{HCl}(g)$ hydrogen chloride	$\text{HCl}(aq)$ <i>hydrochloric acid</i>
$\text{H}_2\text{S}(g)$ hydrogen sulfide	$\text{H}_2\text{S}(aq)$ <i>hydrosulfuric acid</i>

- Acids that contain hydrogen, oxygen, plus another element are called **oxoacids**
- They are named according to the number of oxygen atoms in the molecule and do not take the prefix *hydro-*
- When there are two oxoacids, the one with the larger number of oxygens takes the suffix *-ic* and the one with the fewer oxygen atoms takes the suffix *-ous*

H_2SO_4 *sulfuric acid* HNO_3 *nitric acid*
 H_2SO_3 *sulfurous acid* HNO_2 *nitrous acid*

- The halogen can occur with up to four different oxoacids
- The oxoacid with the most oxygens has the prefix *per-* the one with the least has the prefix *hypo-*

HClO *hypochlorous acid* HClO_3 *chloric acid*
 HClO_2 *chlorous acid* HClO_4 *perchloric acid*

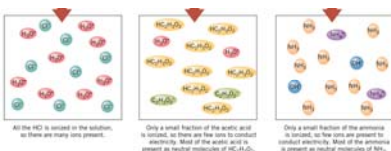
- Anions are produced when oxoacids are neutralized
- There is a simple relationship between the name of the polyatomic ion and the parent acid
 - 1) *-ic* acids give *-ate* ions
 - 2) *-ous* acids give *-ite* ions
- In naming polyatomic anions, the prefixes *per-* and *hypo-* carry over from the parent acid

- Naming bases is much less complicated
- Ionic compounds containing metal ions are named like any other ionic compound
- Molecular bases are specified by giving the name of the molecule
- Acids and bases can be classified as strong or weak and so as strong or weak electrolytes
- Strong acids are strong electrolytes

- The most common strong acids are:
 - $\text{HClO}_4(aq)$ perchloric acid
 - $\text{HCl}(aq)$ hydrochloric acid
 - $\text{HBr}(aq)$ hydrobromic acid
 - $\text{HI}(aq)$ hydroiodic acid
 - $\text{HNO}_3(aq)$ nitric acid
 - $\text{H}_2\text{SO}_4(aq)$ sulfuric acid
- Strong bases are the soluble metal hydroxides

- These include:

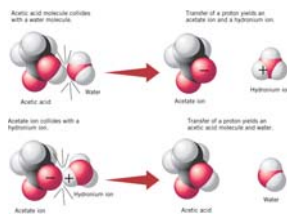
Group IA	Group IIA
LiOH lithium hydroxide	
NaOH sodium hydroxide	
KOH potassium hydroxide	$\text{Ca}(\text{OH})_2$ calcium hydroxide
RbOH rubidium hydroxide	$\text{Sr}(\text{OH})_2$ strontium hydroxide
CsOH cesium hydroxide	$\text{Ba}(\text{OH})_2$ barium hydroxide
- Most acids are not completely ionized in water
- They are classified as weak electrolytes



The brightness of light is experimental verification of the classification as a strong or weak electrolyte.

Weak acids and bases are weak electrolytes because less than 100% of the molecules ionize.

- Weak acids and bases are in *dynamic equilibrium* in solution
- Consider the case of acetic acid:



Two opposing reactions occur in solution: the ionization of the acid, called the **forward reaction**, and the recombination of ions into molecules, called the **reverse reaction**.

Chemical or dynamic equilibrium results when the rate of the forward and reverse reaction are equal.

- Solutions are characterized by their concentration
- The **molar concentration** or **molarity (M)** is defined as

$$\text{molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

- The molarity of a solution gives an equivalence relation between the moles of solute and volume of solution

- Solutions provide a convenient way to combine reactants in many chemical reactions
 - Example: How many grams of AgNO_3 are needed to prepare 250 mL of 0.0125 M AgNO_3 solution?

ANALYSIS: Find moles, then mass of solute.

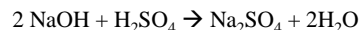
SOLUTION:

$$0.250 \text{ L AgNO}_3 \text{ sol} \times \frac{0.0125 \text{ mol AgNO}_3}{1.00 \text{ L AgNO}_3 \text{ sol}} \times \frac{169.9 \text{ g AgNO}_3}{1 \text{ mol AgNO}_3} = 0.531 \text{ g AgNO}_3$$

- Solutions of high concentration can be **diluted** to make solutions of lower concentration
- Conservation of solute mass requires:

$$V_{\text{dil}} \times M_{\text{dil}} = V_{\text{concd}} \times M_{\text{concd}}$$
- Where *dil* labels the diluted and *concd* the concentrated solution
- Stoichiometry problems often require working with volumes and molarity

- Example: How many mL of 0.124 M NaOH are required to react completely with 15.4 mL of 0.108 M H_2SO_4 ?

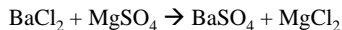


ANALYSIS: Use the mole-to-mole ratio to convert.

SOLUTION:

$$0.0154 \text{ L H}_2\text{SO}_4 \text{ sol} \times \frac{0.108 \text{ mol H}_2\text{SO}_4}{1.00 \text{ L H}_2\text{SO}_4 \text{ sol}} \times \frac{2 \text{ mol NaOH}}{1 \text{ mol H}_2\text{SO}_4} \times \frac{1.00 \text{ L NaOH sol}}{0.124 \text{ mol NaOH}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 26.8 \text{ mL NaOH sol}$$

- Limiting reagent problems are also common
 - Example: How many moles of BaSO_4 will form if 20.0 mL of 0.600 M BaCl_2 is mixed with 30.0 mL of 0.500 M MgSO_4 ?



ANALYSIS: This is a limiting reagent problem.

SOLUTION:

$$0.0200 \text{ L BaCl}_2 \text{ sol} \times \frac{0.600 \text{ mol BaCl}_2}{1.00 \text{ L BaCl}_2 \text{ sol}} \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol BaCl}_2} = 0.0120 \text{ mol BaSO}_4$$

$$0.0300 \text{ L MgSO}_4 \text{ sol} \times \frac{0.500 \text{ mol MgSO}_4}{1.00 \text{ L MgSO}_4 \text{ sol}} \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol MgSO}_4} = 0.0150 \text{ mol BaSO}_4$$

\therefore 0.0120 mol BaSO_4 formed

- Titration** is a technique used to make quantitative measurements of the amounts of solutions
- The *end-point* is often determined visually



The long tube is called the **buret**. The valve at the bottom of the buret is called the **stopcock**. The titration is complete when the indicator changes color.

- Paths for working stoichiometry problems may be summarized with a flowchart:

