Types of Chemical Reactions

and

Solution Stoichiometry

Aqueous Solutions

Water is the dissolving medium, or solvent.

Some Properties of Water

- Water is "bent" or V-shaped.
- The O-H bonds are covalent.
- Water is a polar molecule.
- Hydration occurs when salts dissolve in water.

Figure 4.1: (Left) The water molecule is polar. (Right) A space-filling model of the water molecule.

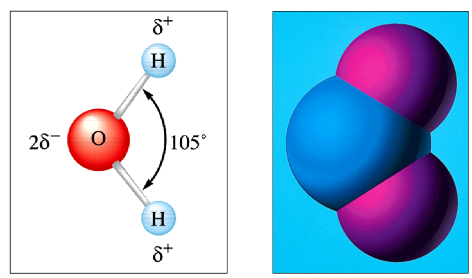
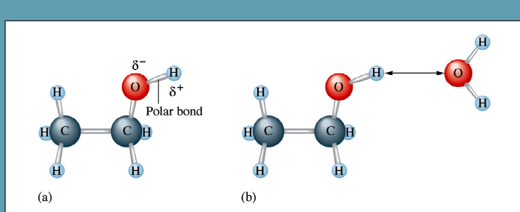


Figure 4.3: (a) The ethanol molecule contains a polar O—H bond similar to those in the water molecule. (b) The polar water molecule interacts strongly with the polar O—H bond in ethanol. This is a case of "like dissolving like."



A Solute

- dissolves in water (or other "solvent")
- changes phase (if different from the solvent)
- is present in lesser amount (if the same phase as the solvent)

A Solvent

- retains its phase (if different from the solute)
- is present in greater amount (if the same phase as the solute)

Ions in Aqueous Solution

Ionic Theory of Solutions

- Some molecular compounds dissolve but do not dissociate into ions.

$$C_6H_{12}O_6(s)\ \text{(glucose)} \xrightarrow{H_2O} C_6H_{12}O_6(aq)$$

- These compounds are referred to as **nonelectrolytes**. They dissolve in water to give a **nonconducting** solution.

Ions in Aqueous Solution

Ionic Theory of Solutions

- Many ionic compounds ***dissociate*** into independent ions when dissolved in water

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

- These compounds that "freely" dissociate into independent ions in aqueous solution are called ***electrolytes.***
- Their aqueous solutions are capable of conducting an electric current.

Figure 4.2: Polar water molecules interact with the positive and negative ions of a salt assisting in the dissolving process.

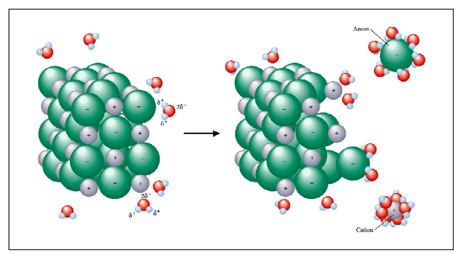
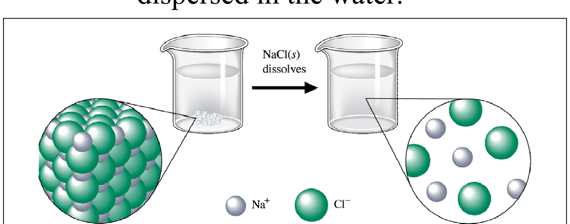


Figure 4.5: When solid NaCl dissolves, the Na^+ and Cl^- ions are randomly dispersed in the water.



Ions in solution conduct electricity. Strong electrolytes in solution conduct electricity!

Electrolytes

Strong - conduct current efficiently

NaCl, HNO_3

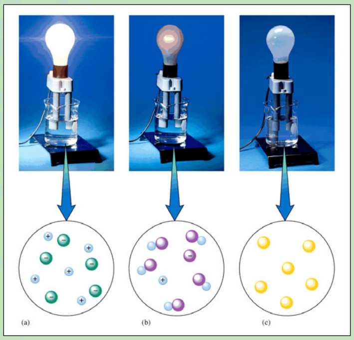
Weak - conduct only a small current

vinegar, tap water

Non - no current flows

pure water, sugar solution

Figure 4.4: Electrical conductivity of aqueous solutions.



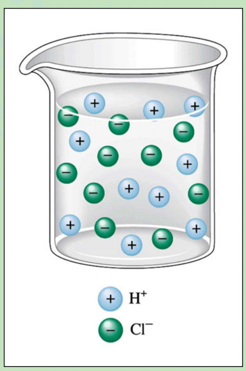
Ions in Aqueous Solution

Ionic Theory of Solutions

- Electrolytes are substances that dissolve in water to give an electrically conducting solution.
  - Thus, in general, ***ionic solids*** that dissolve in water are ***electrolytes***.
  - Some ***molecular compounds***, such as acids, also dissociate in aqueous solution and are considered ***electrolytes.***

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

Figure 4.6: HCl(*aq*) is completely ionized.



Acids

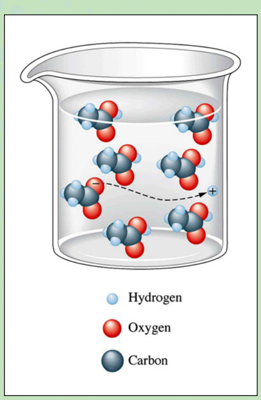
Strong acids - dissociate completely to produce H^+ in solution

hydrochloric and sulfuric acid

Weak acids - dissociate to a slight extent to give H^+ in solution

acetic and formic acid

Figure 4.8: Acetic acid ($HC_2H_3O_2$) exists in water mostly as undissociated molecules. Only a small percentage of the molecules are ionized.



Bases

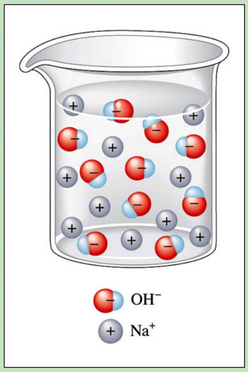
Strong bases - react completely with water to give OH^- ions.

sodium hydroxide

Weak bases - react only slightly with water to give OH^- ions.

ammonia

Figure 4.7: An aqueous solution of sodium hydroxide.



Ions in Aqueous Solution

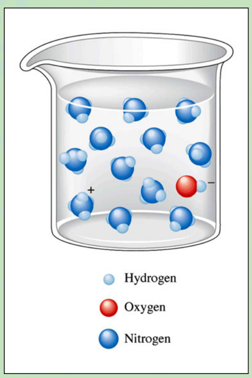
Ionic Theory of Solutions

- Strong and weak electrolytes.
  - A ***weak electrolyte*** is an electrolyte that dissolves in water to give a relatively small percentage of ions.
  - $NH_3\ (g) + H_2O\ (l) \rightleftharpoons NH_4OH\ (aq)$

$$NH_4OH(aq) \rightleftarrows NH_4^+(aq) + OH^-(aq)$$

  - Most soluble molecular compounds are either ***nonelectrolytes*** or ***weak electrolytes***.

Figure 4.9: The reaction of NH_3 in water.



Ions in Aqueous Solution

Ionic Theory of Solutions: Summary

- In summary, substances that dissolve in water are either ***electrolytes*** or ***nonelectrolytes***.
  - ***Nonelectrolytes*** form nonconducting solutions because they ***dissolve as molecules***.
  - ***Electrolytes*** form conducting solutions because they ***dissolve as ions***.

Ions in Aqueous Solution

Ionic Theory of Solutions: Summary

- Electrolytes can be ***strong*** or ***weak***.
  - Almost all ionic substances that dissolve are ***strong electrolytes***.
  - Molecular substances that dissolve are either ***nonelectrolytes*** or ***weak electrolytes***.

Working with Solutions

Molar Concentration

- ***Molar concentration***, or ***molarity (M)***, is defined as the moles of solute dissolved in one liter (cubic decimeter) of solution.

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

Molarity

Molarity (M) = moles of solute per volume of solution in liters:

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

$$3\ M\ \text{HCl} = \frac{\text{6 moles of HCl}}{\text{2 liters of solution}}$$

Common Terms of Solution Concentration

Stock - routinely used solutions prepared in concentrated form.

Concentrated - *relatively* large ratio of solute to solvent. (5.0 M NaCl)

Dilute - *relatively* small ratio of solute to solvent. (0.01 M NaCl)

Working with Solutions

- The majority of chemical reactions discussed here occur in **aqueous solution**.
  - When you run reactions in liquid solutions, it is convenient to dispense the amounts of reactants by measuring out ***volumes*** of reactant solutions.

Figure 4.11:
(a) A measuring pipet is graduated and can be used to measure various volumes of liquid accurately.
(b) a volumetric (transfer) pipet is designed to measure one volume accurately.

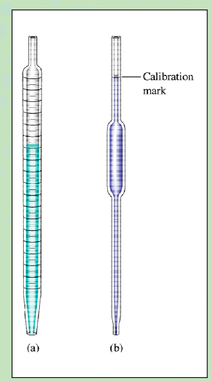
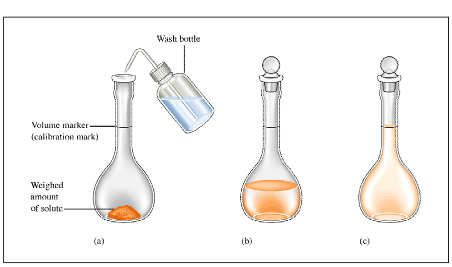


Figure 4.10: Steps involved in the preparation of a standard aqueous solution.



Working with Solutions

Molar Concentration

- When we dissolve a substance in a liquid, we call the substance the **solute** and the liquid the **solvent**.
  - The general term ***concentration*** refers to the quantity of solute in a standard quantity of solution.

Working with Solutions

Molar Concentration

- Let's try an example.
  - A sample of 0.0341 mol iron(III) chloride, $FeCl_3$, was dissolved in water to give 25.0 mL of solution. What is the molarity of the solution?
  - Since $\text{molarity} = \dfrac{\text{moles of } FeCl_3}{\text{liters of solution}}$

then $M = \dfrac{0.0341 \text{ mole of } FeCl_3}{0.0250 \text{ liter of solution}} = 1.36 \text{ M } FeCl_3$

Working with Solutions

How many moles of Cl^- are in 1.00 L of 1.36 M $FeCl_3$ solution?

$\dfrac{1.36 \text{ moles } FeCl_3}{1.00 \text{ L soln.}} \times \dfrac{3 \text{ moles } Cl^-}{1 \text{ mole } FeCl_3} \times 1.00 \text{ L soln} = 4.08 \text{ moles } Cl^-$

$$FeCl_3\,(s) \xrightarrow{H_2O} Fe^{3+}\,(aq) + 3\,Cl^-\,(aq)$$

Working with Solutions

From the equation $\dfrac{\text{moles of solute}}{\text{Liters of solution}} = \text{Molarity}$

$\dfrac{\text{Moles}}{\text{L}} = M$ or, rearranging: L x M = moles

How many moles of NaCl are in 25.0 ml of 0.100 M NaCl solution?

First, convert ml to L: 25.0 ml x 10^{-3} L/ml = 0.0250 ml

(0.0250 L soln) (0.100 moles NaCl/L soln) = 0.00250 moles NaCl

Working with Solutions

Diluting Solutions

- The ***molarity*** of a solution and its ***volume*** are inversely proportional. Therefore, adding water makes the solution less concentrated.
  - This inverse relationship takes the form of:

Moles = Moles

$$M_i \times V_i = M_f \times V_f$$

  - So, as water is added, increasing the final volume, V_f, the final molarity, M_f, decreases.

How many liters of 0.100 M HCl will be needed to prepare 0.500 L of 0.0750 M HCl?

Use $M_iV_i = M_fV_f$

We want to solve for V_i

M_i = 0.100 M HCl

M_f = 0.0750 M HCl

V_f = 0.500 L HCl

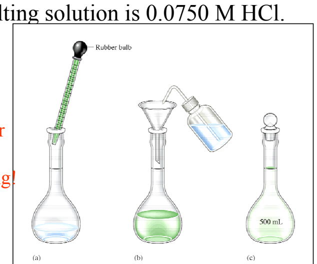
Solving the equation for V_i

$V_i = (M_fV_f)/M_i$ = (0.0750 M x 0.500 L)/(0.100 M) =

0.375 L

Figure 4.12: Dilution Procedure
(a) About 100 mL of water is added to a 500 mL flask.
(b) A measuring pipet is then used to transfer 375 of 0.100 M HCl solution to a volumetric flask.
(c) Water is added to the flask to the calibration mark.
(d) The resulting solution is 0.0750 M HCl.

Always add acids to water to prevent violent boiling!



Types of Solution Reactions

- Precipitation reactions

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

- Acid-base reactions

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

- Oxidation-reduction reactions

$$Fe_2O_3(s) + 2\,Al(s) \rightarrow 2\,Fe(s) + Al_2O_3(s)$$

Ions in Aqueous Solution

Molecular and Ionic Equations

- A ***molecular equation*** is one in which the reactants and products are written as if they were molecules, ***even though they may actually exist in solution as ions***.

$$Ca(OH)_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2NaOH(aq)$$

  – Note that $Ca(OH)_2$, Na_2CO_3, and NaOH are all soluble compounds but $CaCO_3$ is not.

Simple Rules for Solubility

1. Most nitrate (NO_3^-) salts are soluble.
2. Most alkali (group 1A) salts and NH_4^+ are soluble.
3. Most Cl^-, Br^-, and I^- salts are soluble (NOT Ag^+, Pb^{2+}, Hg_2^{2+})
4. Most sulfate salts are soluble (NOT $BaSO_4$, $PbSO_4$, $HgSO_4$, $CaSO_4$)
5. Most OH^- salts are only slightly soluble (NaOH, KOH are soluble, $Ba(OH)_2$, $Ca(OH)_2$ are marginally soluble)
6. Most S^{2-}, CO_3^{2-}, CrO_4^{2-}, PO_4^{3-} salts are only slightly soluble.

Ions in Aqueous Solution

Molecular and Ionic Equations

- An ***ionic equation***, however, represents strong electrolytes as separate independent ions. This is a more accurate representation of the way electrolytes behave in solution.

$$Ca^{2+}(aq) + 2OH^-(aq) + 2Na^+(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)\downarrow + 2Na^+(aq) + 2OH^-(aq)$$

Ions in Aqueous Solution

Molecular and Ionic Equations

- Complete and net ionic equations
  – A ***complete ionic equation*** is a chemical equation in which strong electrolytes (such as soluble ionic compounds) are written as separate ions in solution.

$$\underset{\text{(strong)}}{Ca(NO_3)_2(aq)} + \underset{\text{(strong)}}{K_2CO_3(aq)} \rightarrow \underset{\text{(insoluble)}}{CaCO_3(s)} + \underset{\text{(strong)}}{2KNO_3(aq)}$$

$$Ca^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s) + 2K^+(aq) + 2NO_3^-(aq)$$

Ions in Aqueous Solution

Molecular and Ionic Equations

- Complete and net ionic equations.
  – A ***net ionic equation*** is a chemical equation from which the spectator ions have been removed.
  – A ***spectator ion*** is an ion in an ionic equation that does not take part in the reaction.

$$Ca^{2+}(aq) + \cancel{2NO_3^-}(aq) + \cancel{2K^+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s) + \cancel{2K^+}(aq) + \cancel{2NO_3^-}(aq)$$

Ions in Aqueous Solution

Molecular and Ionic Equations

- Complete and net ionic equations
  - Separating the strong electrolytes into separate ions, we obtain the complete ionic equation.

$$2H^+(aq) + 2NO_3^-(aq) + Mg(OH)_2(s) \rightarrow$$
$$2H_2O(l) + Mg^{2+}(aq) + 2NO_3^-(aq)$$

  - Note that the nitrate ions did not participate in the reaction. These are ***spectator ions***.

Ions in Aqueous Solution

Molecular and Ionic Equations

- Complete and net ionic equations
  - Eliminating the spectator ions results in the net ionic equation.

$$2H^+(aq) + \cancel{2NO_3^-}(aq) + Mg(OH)_2(s) \rightarrow$$
$$2H_2O(l) + Mg^{2+}(aq) + \cancel{2NO_3^-}(aq)$$

$$2H^+(aq) + Mg(OH)_2(s) \rightarrow 2H_2O(l) + Mg^{2+}(aq)$$

This equation represents the "essential" reaction.

TABLE 4.1 Simple Rules for the Solubility of Salts in Water

1. Most nitrate (NO_3^-) salts are soluble.
2. Most salts containing the alkali metal ions (Li^+, Na^+, K^+, Cs^+, Rb^+) and the ammonium ion (NH_4^+) are soluble.
3. Most chloride, bromide, and iodide salts are soluble. Notable exceptions are salts containing the ions Ag^+, Pb^{2+}, and Hg_2^{2+}.
4. Most sulfate salts are soluble. Notable exceptions are $BaSO_4$, $PbSO_4$, Hg_2SO_4, and $CaSO_4$.
5. Most hydroxide salts are only slightly soluble. The important soluble hydroxides are NaOH and KOH. The compounds $Ba(OH)_2$, $Sr(OH)_2$, and $Ca(OH)_2$ are marginally soluble.
6. Most sulfide (S^{2-}), carbonate (CO_3^{2-}), chromate (CrO_4^{2-}), and phosphate (PO_4^{3-}) salts are only slightly soluble.

Figure 4.16: Precipitation of silver chloride by mixing solutions of silver nitrate and potassium chloride. The K+ and NO_3- ions remain in solution.

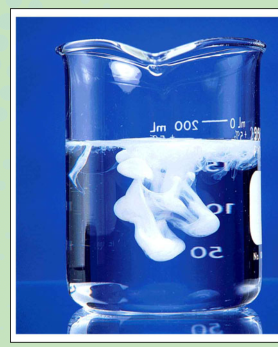
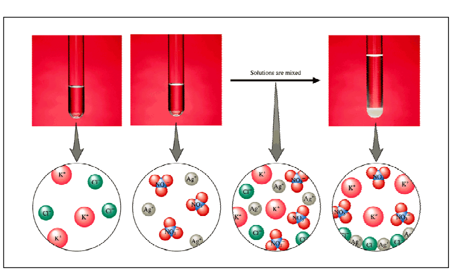


Figure 4.17: The reaction of KCl(*aq*) with $AgNO_3$ to form AgCl(*s*).



Describing Reactions in Solution

1. Molecular equation (reactants and products as compounds)

$$AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$$

2. Complete ionic equation (all strong electrolytes shown as ions)

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

Describing Reactions in Solution
(continued)

3. Net ionic equation (show only components that actually react)

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

K^{+} and NO_3^{-} are spectator ions.

Types of Chemical Reactions

Acid-Base Reactions

- The Arrhenius Concept
  - The Arrhenius concept defines ***acids*** as ***substances that produce hydrogen ions, H^{+}, when dissolved in water.***
  - An example is nitric acid, HNO_3, a molecular substance that dissolves in water to give H^{+} and NO_3^{-}.

$$HNO_3(aq) \xrightarrow{H_2O} H^{+}(aq) + NO_3^{-}(aq)$$

Types of Chemical Reactions

Acid-Base Reactions

- The Arrhenius Concept
  - The Arrhenius concept defines ***bases*** as ***substances that produce hydroxide ions, OH^{-}, when dissolved in water.***
  - An example is sodium hydroxide, NaOH, an ionic substance that dissolves in water to give sodium ions and hydroxide ions.

$$NaOH(s) \xrightarrow{H_2O} Na^{+}(aq) + OH^{-}(aq)$$

Types of Chemical Reactions

Acid-Base Reactions

- The Arrhenius Concept
  - The molecular substance ammonia, NH_3, is a base in the Arrhenius view,

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^{+}(aq) + OH^{-}(aq)$$

because it yields hydroxide ions when it reacts with water.

Types of Chemical Reactions

Acid-Base Reactions

- The Brønsted-Lowry Concept
  - The **Brønsted-Lowry** concept of acids and bases involves the transfer of a proton (H^{+}) from the acid to the base.
  - In this view, acid-base reactions are ***proton-transfer reactions.***

Types of Chemical Reactions

Acid-Base Reactions

- The Brønsted-Lowry Concept
  - The **Brønsted-Lowry** concept defines an **acid** as the ***species (molecule or ion) that donates a proton (H^{+}) to another species in a proton-transfer reaction.***
  - A ***base*** is defined as the ***species (molecule or ion) that accepts the proton (H^{+}) in a proton-transfer reaction.***

Types of Chemical Reactions

Acid-Base Reactions

- The Brønsted-Lowry Concept

In the reaction of ammonia with water,

$$NH_3(aq) + H_2O(l) \leftrightarrows NH_4^+(aq) + OH^-(aq)$$

H^+

the H_2O molecule is the acid because it donates a proton. The NH_3 molecule is a base, because it accepts a proton.

Types of Chemical Reactions

Acid-Base Reactions

- The Brønsted-Lowry Concept

The $H^+_{(aq)}$ ion associates itself with water to form $H_3O^+_{(aq)}$.

$$H^+(aq) + H_2O(l) \rightarrow H_3O^+(aq)$$

This "mode of transportation" for the H^+ ion is called the **hydronium ion**.

Types of Chemical Reactions

Acid-Base Reactions

- The Brønsted-Lowry Concept

The dissolution of nitric acid, HNO_3, in water is therefore a proton-transfer reaction

$$HNO_3(aq) + H_2O(l) \rightarrow NO_3^-(aq) + H_3O^+(aq)$$

H^+

where HNO_3 is an acid (proton donor) and H_2O is a base (proton acceptor).

Types of Chemical Reactions

Acid-Base Reactions

- In summary, the Arrhenius concept and the Brønsted-Lowry concept are essentially the same in aqueous solution.
  - The Arrhenius concept

acid: proton (H^+) donor

base: hydroxide ion (OH^-) donor

Types of Chemical Reactions

Acid-Base Reactions

- In summary, the Arrhenius concept and the Brønsted-Lowry concept are essentially the same in aqueous solution.
  - The Brønsted-Lowry concept

acid: proton (H^+) donor

base: proton (H^+) acceptor

Types of Chemical Reactions

Acid-Base Reactions

- Strong and Weak Acids and Bases
  - A **strong acid** is an acid that ionizes ***completely*** in water; it is a ***strong*** electrolyte.

$$HNO_3(aq) + H_2O(l) \rightarrow NO_3^-(aq) + H_3O^+(aq)$$

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

  - Table 4.3 lists the common strong acids.

Types of Chemical Reactions

Acid-Base Reactions

- Strong and Weak Acids and Bases
  - A **strong base** is a base that is present ***entirely*** as ions, one of which is OH^-; it is a ***strong*** electrolyte.

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

  - The hydroxides of Group IA and IIA elements, except for beryllium hydroxide, are strong bases. (see Table 4.3)

Table 4.3
Common Strong Acids and Bases

Strong Acids	Strong Bases
$HClO_4$	LiOH
H_2SO_4	NaOH
HI	KOH
HBr	$Ca(OH)_2$
HCl	$Sr(OH)_2$
HNO_3	$Ba(OH)_2$

Return to Slide 43

Types of Chemical Reactions

Acid-Base Reactions

- Strong and Weak Acids and Bases
  - You will find it important to be able to identify an acid or base as strong or weak.
  - When you write an ionic equation, ***strong acids and bases are represented as separate ions***.
  - Weak acids and bases ***are represented as undissociated "molecules" in ionic equations***.

Types of Chemical Reactions

Acid-Base Reactions

- Strong and Weak Acids and Bases
  - A **weak acid** is an acid that only ***partially*** ionizes in water; it is a ***weak*** electrolyte.
  - The hydrogen cyanide molecule, HCN, reacts with water to produce a small percentage of ions in solution.

$$HCN(aq) + H_2O(l) \leftrightarrows CN^-(aq) + H_3O^+(aq)$$

Types of Chemical Reactions

Acid-Base Reactions

- Strong and Weak Acids and Bases
  - A **weak base** is a base that is only ***partially*** ionized in water; it is a ***weak*** electrolyte.
  - Ammonia, NH_3, is an example.

$$NH_3(aq) + H_2O(l) \leftrightarrows NH_4^+(aq) + OH^-(aq)$$

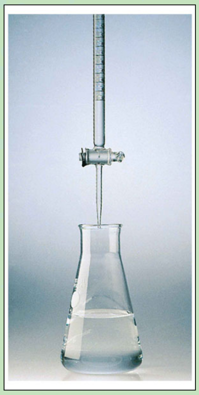
Types of Chemical Reactions

Acid-Base Reactions

- Neutralization Reactions
  - One of the chemical properties of acids and bases is that ***they neutralize one another***.
  - A ***neutralization reaction*** is a reaction of an acid and a base that results in an ionic compound and water.
  - The ionic compound that is the product of a neutralization reaction is called a *salt*.

$$\underset{\text{acid}}{HCN(aq)} + \underset{\text{base}}{KOH(aq)} \rightarrow \underset{\text{salt}}{KCN(aq)} + H_2O(l)$$

Figure 4.18a: The titration of an acid with a base.



Key Titration Terms

Titrant - solution of known concentration used in titration

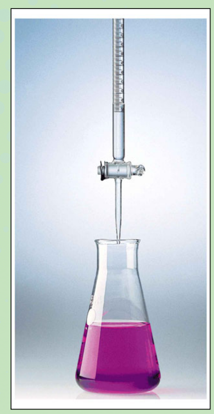
Analyte - substance being analyzed

Equivalence point - enough titrant added to react exactly with the analyte

Endpoint - the indicator changes color so you can tell the equivalence point has been reached.

Figure 4.18b: The titration of an acid with a base

Figure 4.18c: The titration of an acid with a base.



Quantitative Analysis

Volumetric Analysis

- Consider the reaction of sulfuric acid, H_2SO_4, with sodium hydroxide, NaOH:

$$\mathbf{H_2SO_4(aq) + 2NaOH(aq) \rightarrow 2H_2O(l) + Na_2SO_4(aq)}$$

   - Suppose a beaker contains 35.0 mL of 0.175 *M* H_2SO_4. How many milliliters of 0.250 *M* NaOH must be added to completely react with the sulfuric acid?

Quantitative Analysis

Volumetric Analysis

- After balancing the equation, we must convert the 0.0350 L (35.0 mL) to moles of H_2SO_4 (using the molarity of the H_2SO_4).

- Then, convert to moles of NaOH (from the balanced chemical equation).
- Finally, convert to volume of NaOH solution (using the molarity of NaOH).

$$\mathbf{(0.0350L) \times \frac{0.175\ mole\ H_2SO_4}{1\ L\ H_2SO_4\ solution} \times \frac{2\ mol\ NaOH}{1\ mol\ H_2SO_4} \times \frac{1\ L\ NaOH\ soln.}{0.250\ mol\ NaOH} =}$$

0.0490 L NaOH solution (or 49.0 mL of NaOH solution)

Performing Calculations for Acid-Base Reactions

1. List initial species and predict reaction.
2. Write balanced net ionic reaction.
3. Calculate moles of reactants.
4. Determine limiting reactant.
5. Calculate moles of required reactant/product.
6. Convert to grams or volume, as required.

Types of Chemical Reactions

Acid-Base Reactions

- Neutralization Reactions
  - The net ionic equation for each acid-base neutralization reaction involves a transfer of a proton.
  - Consider the reaction of the strong acid , HCl(*aq*) and a strong base, LiOH(*aq*).

$$HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$$

Types of Chemical Reactions

Acid-Base Reactions

- Neutralization Reactions
  - Writing the strong electrolytes in the form of ions gives the complete ionic equation.

$$H^+(aq) + Cl^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + Cl^-(aq) + H_2O(l)$$

Types of Chemical Reactions

Acid-Base Reactions

- Neutralization Reactions
  - Canceling the spectator ions results in the net ionic equation. Note the proton transfer.

$$H^+(aq) + \cancel{Cl}^-(aq) + \cancel{K}^+(aq) + OH^-(aq) \rightarrow \cancel{K}^+(aq) + \cancel{Cl}^-(aq) + H_2O(l)$$

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

H^+

Types of Chemical Reactions

Acid-Base Reactions

- Neutralization Reactions
  - In a reaction involving HCN(*aq*), a weak acid, and KOH(*aq*), a strong base, the product is KCN, a strong electrolyte.
  - The net ionic equation for this reaction is

$$HCN(aq) + OH^-(aq) \rightarrow CN^-(aq) + H_2O(l)$$

H^+

Note the proton transfer.

Types of Chemical Reactions

Acid-Base Reactions

- Acid-Base Reactions with Gas Formation
  - Sulfides react with acids to form H_2S, hydrogen sulfide gas.

$$Na_2S + 2HCl \rightarrow 2NaCl + H_2S \uparrow$$

  - These reactions are summarized in Table 4.4.

Types of Chemical Reactions

Acid-Base Reactions

- Acid-Base Reactions with Gas Formation
  - Carbonates react with acids to form CO_2, carbon dioxide gas.

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2 \uparrow$$

  - Sulfites react with acids to form SO_2, sulfur dioxide gas.

$$Na_2SO_3 + 2HCl \rightarrow 2NaCl + H_2O + SO_2 \uparrow$$

Acid/Base Reactions

Table 4.4
Some Ionic Compounds That Evolve Gases When Treated with Acids

Ionic Compound	Gas	Example
Carbonate (CO_3^{2-})	CO_2	$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + H_2O + CO_2$
Sulfite (SO_3^{2-})	SO_2	$Na_2SO_3 + 2HCl \longrightarrow 2NaCl + H_2O + SO_2$
Sulfide (S^{2-})	H_2S	$Na_2S + 2HCl \longrightarrow 2NaCl + H_2S$

Return to Slide 52

Types of Chemical Reactions

- Oxidation-Reduction Reactions
  - ***Oxidation-reduction reactions*** involve the transfer of electrons from one species to another.
  - ***Oxidation*** is defined as the loss of electrons.
  - ***Reduction*** is defined as the gain of electrons.
  - Oxidation and reduction always occur simultaneously.

Figure 4.19: The reaction of solid sodium and gaseous chlorine to form solid sodium chloride.

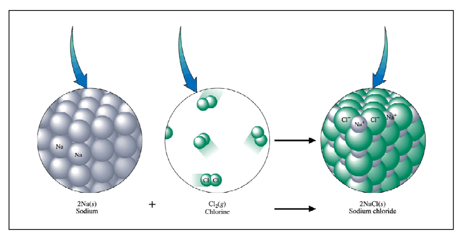
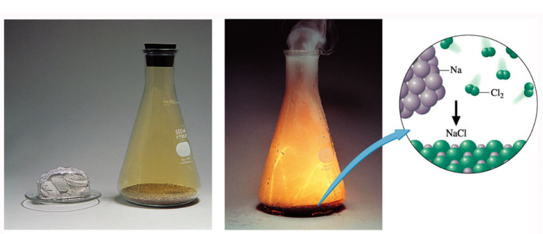


Figure 4.13: Combination reaction.



Return to Slide 63

Figure 4.20: A summary of an oxidation-reduction process, in which M is oxidized and X is reduced.

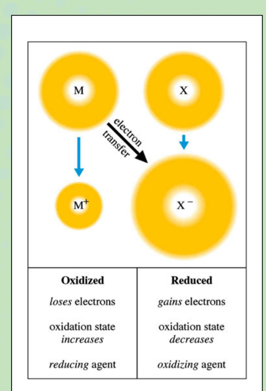


Figure 4.10: Reaction of iron with Cu^{2+} *(aq)*. Photo courtesy of American Color.

Return to Slide 54

Types of Chemical Reactions

- Oxidation-Reduction Reactions
  - The reaction of an iron nail with a solution of copper(II) sulfate, $CuSO_4$, is an oxidation-reduction reaction (see Figure 4.10).
  - The molecular equation for this reaction is:

$$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$$

Types of Chemical Reactions

- Oxidation-Reduction Reactions
  - The net ionic equation shows the reaction of iron metal with $Cu^{2+}_{(aq)}$ to produce iron(II) ion and copper metal.

Loss of 2 e^{-1} oxidation

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

Gain of 2 e^{-1} reduction

Types of Chemical Reactions

Oxidation-Reduction Reactions

- Oxidation Numbers
  - The concept of oxidation numbers is a simple way of keeping track of electrons in a reaction.
  - The ***oxidation number*** (or oxidation state) of an atom in a substance is the actual ***charge*** of the atom if it exists as a monatomic ion.
  - Alternatively, it is ***hypothetical charge*** assigned to the atom in the substance by simple rules.

Types of Chemical Reactions

Oxidation-Reduction Reactions

- Oxidation Number Rules

Rule	Applies to	Statement
1	Elements	The oxidation number of an atom in an element is zero. This includes molecular elements.
2	Monatomic ions	The oxidation number of an atom in a monatomic ion equals the charge of the ion.
3	Oxygen	The oxidation number of oxygen is –2 in most of its compounds. (An exception is O in H_2O_2 and other peroxides, where the oxidation number is –1.)

Types of Chemical Reactions

Oxidation-Reduction Reactions

- Oxidation Number Rules

Rule	Applies to	Statement
4	Hydrogen	The oxidation number of hydrogen is +1 in its covalent compounds.
5	Halogens	Fluorine is –1 in all its compounds. The other halogens are –1 unless the other element is another halogen or oxygen.
6	Compounds and ions	The sum of the oxidation numbers of the atoms in a compound is zero. The sum in a polyatomic ion equals the charge on the ion.

Types of Chemical Reactions

Oxidation-Reduction Reactions

- Describing Oxidation-Reduction Reactions
  - Look again at the reaction of iron with copper(II) sulfate.

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

  - We can write this reaction in terms of two ***half-reactions.***

Types of Chemical Reactions

Oxidation-Reduction Reactions

- Describing Oxidation-Reduction Reactions
  - A ***half-reaction*** is one of the two parts of an oxidation-reduction reaction. One involves the loss of electrons (oxidation) and the other involves the gain of electrons (reduction).

$$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$$ **oxidation half-reaction**

$$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$$ **reduction half-reaction**

Types of Chemical Reactions

Oxidation-Reduction Reactions

- Describing Oxidation-Reduction Reactions
  - An ***oxidizing agent*** is a species that oxidizes another species; ***it is itself reduced***.
  - A ***reducing agent*** is a species that reduces another species; ***it is itself oxidized***.

Loss of 2 e⁻ oxidation

reducing agent

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

oxidizing agent

Gain of 2 e⁻ reduction

Types of Chemical Reactions

Oxidation-Reduction Reactions

- Balancing Simple Oxidation-Reduction Reactions
  - At first glance, the equation representing the reaction of zinc metal with silver(I) ions might appear to be balanced.

$$Zn(s) + Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + Ag(s)$$

  - However, a balanced equation must have a ***charge balance*** as well as a ***mass balance***.

Types of Chemical Reactions

Oxidation-Reduction Reactions

- Balancing Simple Oxidation-Reduction Reactions
  - Since the number of electrons lost in the oxidation half-reaction must equal the number gained in the reduction half-reaction,

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$$ **oxidation half-reaction**

$$2\,Ag^{+}(aq) + 2\,e^{-} \rightarrow 2\,Ag(s)$$ **reduction half-reaction**

we must double the reaction involving the reduction of the silver.

Types of Chemical Reactions

Oxidation-Reduction Reactions

- Balancing Simple Oxidation-Reduction Reactions
  - Adding the two half-reactions together, the electrons cancel,

$$Zn(s) \rightarrow Zn^{2+}(aq) + \cancel{2e^{-}}$$ **oxidation half-reaction**

$$2Ag^{+}(aq) + \cancel{2e^{-}} \rightarrow 2Ag(s)$$ **reduction half-reaction**

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

Figure 4.12: The burning of calcium metal in chlorine. Photo courtesy of James Scherer.

$Ca (s) + Cl_2 (g) \rightarrow CaCl_2 (s)$

$Ca \rightarrow Ca^{2+} + 2 e^-$

$\underline{Cl_2 + 2 e^- \rightarrow 2 Cl^-}$

$Ca (s) + Cl_2 (g) \rightarrow CaCl_2 (s)$

Return to Slide 62

Rules for Assigning Oxidation States

1. Oxidation state of an atom in an element = 0
2. Oxidation state of monatomic element = charge
3. Oxygen = –2 in covalent compounds (except in peroxides where it = –1)
4. H = +1 in covalent compounds
5. Fluorine = –1 in compounds
6. Sum of oxidation states = 0 in compounds
   Sum of oxidation states = charge of the ion

Balancing by Half-Reaction Method

1. Write separate reduction, oxidation reactions.
2. For each half-reaction:
   - Balance elements (except H, O)
   - Balance O using H_2O
   - Balance H using H^+
   - Balance charge using electrons

Half-Reaction Method - Balancing in Base

1. Balance as in acid.
2. Add OH^- that equals H^+ ions (both sides!)
3. Form water by combining H^+, OH^-.
4. Check elements and charges for balance.

Balancing by Half-Reaction Method (continued)

3. If necessary, multiply by integer to equalize electron count.
4. Add half-reactions.
5. Check that elements and charges are balanced.

Stoichiometry Steps for reactions in solution.

Solving Stoichiometry Problems for Reactions in Solution

1. Identify the species present in the combined solution, and determine what reaction occurs.
2. Write the balanced net ionic equation for the reaction.
3. Calculate the moles of reactants.
4. Determine which reactant is limiting.
5. Calculate the moles of product or products, as required.
6. Convert to grams or other units, as required.