Chapter 4

Aqueous solutions

Types of reactions

Parts of Solutions

- Solution- homogeneous mixture.
- Solute- what gets dissolved.
- Solvent- what does the dissolving.
- Soluble- Can be dissolved.
- Miscible- liquids dissolve in each other.

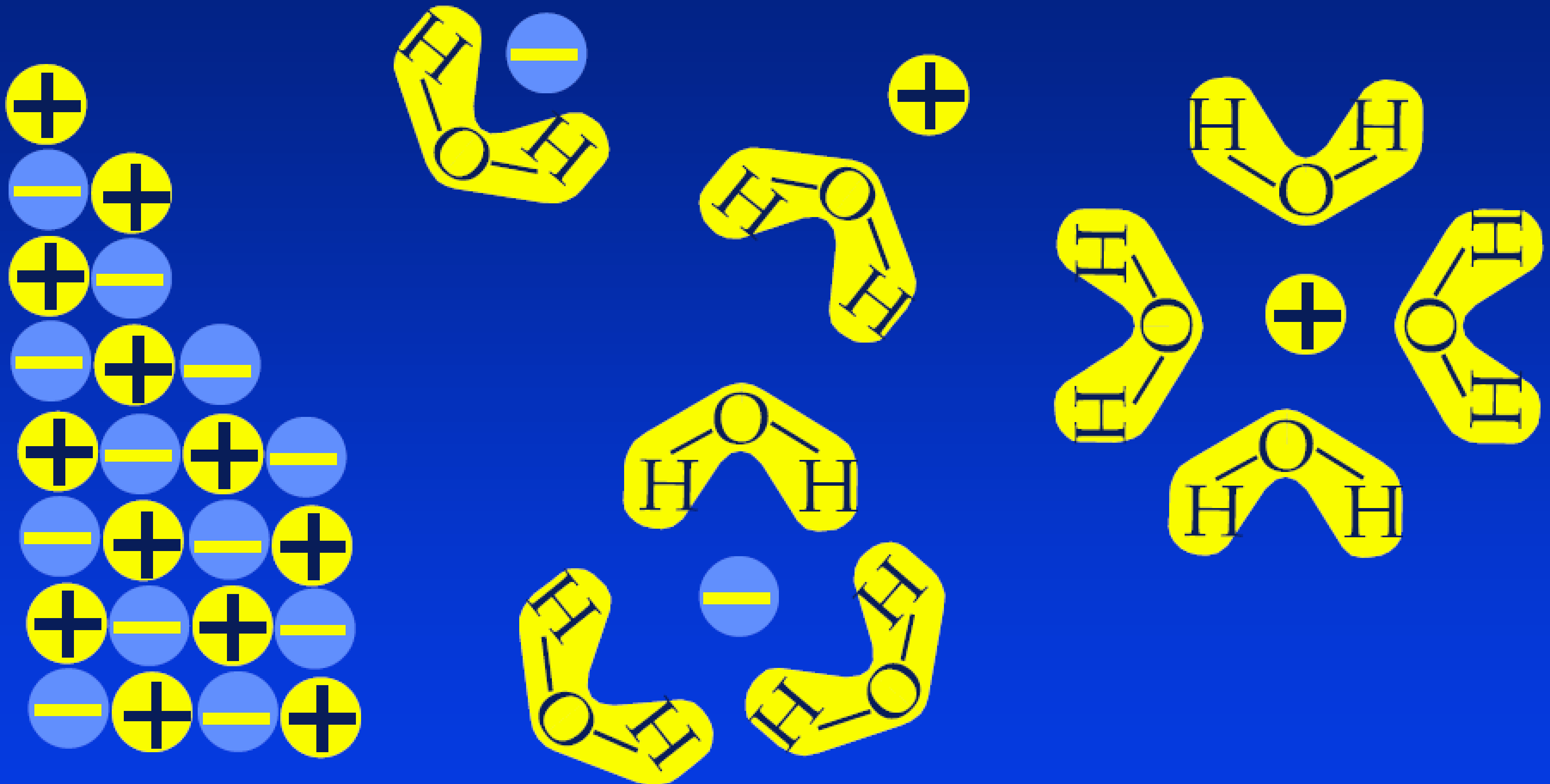
Aqueous solutions

- Dissolved in water.
- Water is a good solvent because the molecules are polar.
- The oxygen atoms have a partial negative charge.
- The hydrogen atoms have a partial positive charge.
- The angle is 105ºC.

Hydration

- The process of breaking the ions of salts apart.
- Ions have charges and attract the opposite charges on the water molecules.

Hydration

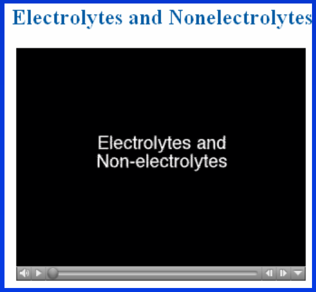


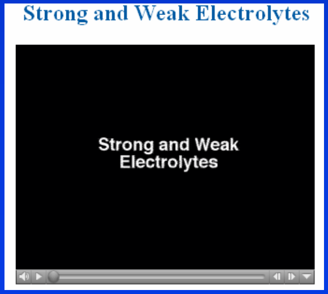
Solubility

- How much of a substance will dissolve in a given amount of water.
- Usually g/100 mL
- Varies greatly, but if they do dissolve the ions are separated,
- and they can move around.
- Water can also dissolve non-ionic compounds if they have polar bonds.

Electrolytes

- Electricity is moving charges.
- The ions that are dissolved can move.
- Solutions of ionic compounds can conduct electricity.
- Electrolytes.
- Solutions are classified three ways.





Types of solutions

- Strong electrolytes- completely dissociate (fall apart into ions).
- Many ions- Conduct well.
- Weak electrolytes- Partially fall apart into ions.
- Few ions -Conduct electricity slightly.
- Non-electrolytes- Don’t fall apart.
- No ions- Don’t conduct.

Water Soluble Compounds

Strong Electrolytes

Strong acids
And
Strong Bases

Example:
HNO_3
H_2SO_4
$NaOH$
$Ba(OH)_2$

Ionic
Compounds

Example:
$NaCl$
KNO_3
KF

Weak Electrolytes

Weak acids
And
Weak Bases

Example:
$HC_2H_3O_2$
HF
NH_3

Non Electrolytes

Molecular
Compounds

Example:
CH_3OH
C_2H_5OH (ethanol)
$C_6H_{12}O_6$ (sugar)

Types of Reactions

1 Precipitation reactions

- When aqueous solutions of ionic compounds are poured together a solid forms.
- A solid that forms from mixed solutions is a precipitate
- If you're not a part of the solution, your part of the precipitate

Precipitation reactions

- $NaOH(aq) + FeCl_3(aq) \rightarrow$
  $NaCl(aq) + Fe(OH)_3(s)$
- is really
- $Na^+(aq) + OH^-(aq) + Fe^{+3} + Cl^-(aq) \rightarrow$
  $Na^+(aq) + Cl^-(aq) + Fe(OH)_3(s)$
- So all that really happens is
- $OH^-(aq) + Fe^{+3} \rightarrow Fe(OH)_3(s)$
- Double replacement reaction

Precipitation reaction

- We can predict the products
- Can only be certain by experimenting
- The anion and cation switch partners
- $AgNO_3(aq) + KCl(aq) \rightarrow$
- $Zn(NO_3)_2(aq) + BaCr_2O_7(aq) \rightarrow$
- $CdCl_2(aq) + Na_2S(aq) \rightarrow$

Precipitations Reactions

- Only happen if one of the products is insoluble
- Otherwise all the ions stay in solution- nothing has happened.
- Need to memorize the rules for solubility (pg 145)

Solubility Rules

1. All nitrates are soluble
2. Alkali metals ions and NH_4^+ ions are soluble
3. Halides are soluble except Ag^+, Pb^{+2}, and Hg_2^{+2}
4. Most sulfates are soluble, except Pb^{+2}, Ba^{+2}, Hg^{+2},and Ca^{+2}

Solubility Rules

5 Most hydroxides are slightly soluble (insoluble) except NaOH and KOH

6 Sulfides, carbonates, chromates, and phosphates are insoluble

* Lower number rules supersede so Na_2S is soluble

Three Types of Equations

- **Molecular Equation**- written as whole formulas, not the ions.
- $K_2CrO_4(aq) + Ba(NO_3)_2(aq) \rightarrow$
- **Complete Ionic equation** show dissolved electrolytes as the ions.
- $2K^+ + CrO_4^{-2} + Ba^{+2} + 2\ NO_3^- \rightarrow$

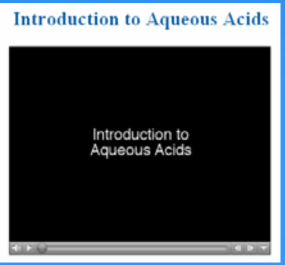
  $BaCrO_4(s) + 2K^+ + 2\ NO_3^-$
- **Spectator ions** are those that don't react.

Three Type of Equations

- Net Ionic equations show only those ions that react, not the spectator ions
- $Ba^{+2} + CrO_4^{-2} \rightarrow BaCrO_4(s)$
- Write the three types of equations for the reactions when these solutions are mixed.
- iron (III) sulfate and potassium sulfide
  Lead (II) nitrate and sulfuric acid.

Acids and Bases

- An acid is a molecular substance that ionizes to form a hydrogen ion (H^+) and increases the concentration of aqueous H^+ ions when it is dissolved in water.



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

Strong Acid	Formula
Hydrochloric	HCl
Hydrobromic	HBr
Hydroiodic	HI
Nitric	HNO_3
Chloric	$HClO_3$
Perchloric	$HClO_4$
Sulfuric *	$\boxed{H_2SO_4}$

The ionization of sulfuric acid occurs in two steps, only the first of which happens completely.

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

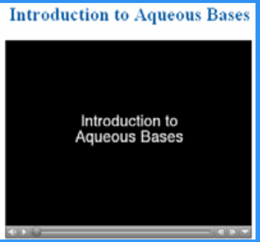
2.) $HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$

*Sulfuric acid is the only strong acid that is diprotic, meaning it has two protons to donate.

**All other acids are considered weak

Acids and Bases

- A **base** is a substance that increases the concentration of aqueous OH^- ions when it is dissolved in water. Bases can be either ionic or molecular substances
- Bases accept H^+ ions



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

<u>Strong Bases</u>

LiOH	$Ca(OH)_2$
NaOH	$Sr(OH)_2$
KOH	$Ba(OH)_2$
RbOH	
CsOH	

Weak Bases

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

Acids and Bases

An acid and a base can react with one another to form a molecular compound and a **salt**. The combination of hydrochloric acid and sodium hydroxide is a familiar **neutralization reaction**.

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

Three equations are written for Acid/Base reactions

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

Complete Ionic: $H^+ + Cl^- + Na^+ + OH^- \rightarrow H_2O(l) + Na^+ + Cl^-$

Net Ionic $H^+ + Cl^- \rightarrow H_2O(l)$

**All in aqueous form. If present as a solid, the entire reactant must be written.

Acids and Bases

- Not all acid/base reactions are easy to identify
- Not all produce salt and water
- The molecular compound produced by the reaction of an acid and a base can also be a gas.

Example #1

$$2HCl(aq) + Na_2S(aq) \rightarrow H_2S(g) + 2NaCl(aq)$$

Net Ionic: $2H^+ + S^{2-} \rightarrow H_2S(g)$

Example #2

$$HCl + NaHCO_3 \rightarrow NaCl + H_2CO_3$$

H_2CO_3 decomposes into:

$$H_2CO_3 \rightarrow H_2O(l) + CO_2$$

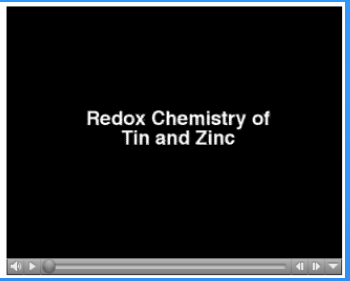
Complete Ionic: $HCl + NaHCO_3 \rightarrow NaCl + H_2CO_3 + H_2O$

Net Ionic: $H^+ + HCO_3^- \rightarrow H_2O + CO_2$

REDOX

- In addition to precipitation and neutralization reactions, aqueous ions can participate in **oxidation-reduction reactions**. Oxidation-reduction reactions involve the transfer of electrons from one chemical species to another. A piece of calcium metal, for example, dissolves in aqueous acid.

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



Rule #1

- The oxidation number of an individual atom in the free state of the element = 0
- Seven elements exist as diatomic molecules in the free state:

Br I N Cl H O F "The diatomic 7"

Examples: Mg He Kr I_2 Cl_2

Rule #2

The oxidation number of a monatomic ion is equal to the charge of the ion.

- Example Mg^{2+} = $^{2+}$

  F^{-} = 1-

Rule #3

- The sum of the oxidation numbers for a polyatomic ion = the charge of the ion
- Example SO_4^{2-} The total must equal 2^-
- Therefore S is a 6+ and O is 2-
- S = 6+ and 4 O each at 2- totals 8-.

Adding 6+ and 8- total 2-.

Rule #4

The sum of the oxidation numbers for a compound must equal 0

Example Cu_2S

The total from the Cu and S must total 0

Therefore the Cu is 1+ and S is 2-

Rules to Remember

Group 1 elements make 1+

Group 2 elements make 2+

Group 17 elements make 1-

Oxygen is 99% always 2-

Metals are listed before nonmentals

Metals are +++ and nonmetals are -----

REDOX - Example #1

$$\overset{0}{Zn_{(s)}} + \overset{+2}{Cu^{2+}_{(aq)}} \rightarrow \overset{+2}{Zn^{2+}_{(aq)}} + \overset{0}{Cu_{(s)}}$$

Which is oxidized? $Zn_{(s)}$: $0 \rightarrow +2$ LEO! Reducing Agent

Which is reduced? Cu^{+2}: $+2 \rightarrow 0$ GER! Oxidizing Agent

REDOX – Example #2

-4 total, -2 each

-2 total, -1 each

+2 total, +1 each

+4 +1 -1 +2 0 -2

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

Which is oxidized? Cl^-: $-1 \rightarrow 0$ LEO Reducing Agent

Which is reduced? Mn^{+4}: $+4 \rightarrow +2$ GER Oxidizing Agent

Writing Net Ionic Equations

- We can write also write a net equation for this type of reaction

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

Net Equation

$$Mn^{+4} + 4Cl^{-} \rightarrow Mn^{+2} + Cl_2$$

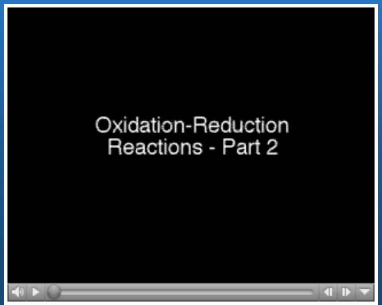
How Can We Tell if a Reaction Will Occur?

- The activity series!
- Elements higher up will react with the ion of the metal below it
- For Example:
  - Pb(s) will displace Cu^{+} in solution
  - However, Cu(s) will not replace Pb^{+2} in solution

TABLE 4.5 Activity Series of Metals in Aqueous Solution

Metal	Oxidation Reaction
Lithium	$Li(s) \longrightarrow Li^{+}(aq) + e^{-}$
Potassium	$K(s) \longrightarrow K^{+}(aq) + e^{-}$
Barium	$Ba(s) \longrightarrow Ba^{2+}(aq) + 2e^{-}$
Calcium	$Ca(s) \longrightarrow Ca^{2+}(aq) + 2e^{-}$
Sodium	$Na(s) \longrightarrow Na^{+}(aq) + e^{-}$
Magnesium	$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^{-}$
Aluminum	$Al(s) \longrightarrow Al^{3+}(aq) + 3e^{-}$
Manganese	$Mn(s) \longrightarrow Mn^{2+}(aq) + 2e^{-}$
Zinc	$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^{-}$
Chromium	$Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^{-}$
Iron	$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^{-}$
Cobalt	$Co(s) \longrightarrow Co^{2+}(aq) + 2e^{-}$
Nickel	$Ni(s) \longrightarrow Ni^{2+}(aq) + 2e^{-}$
Tin	$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^{-}$
Lead	$Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^{-}$
Hydrogen	$H_2(g) \longrightarrow 2H^{+}(aq) + 2e^{-}$
Copper	$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^{-}$
Silver	$Ag(s) \longrightarrow Ag^{+}(aq) + e^{-}$
Mercury	$Hg(l) \longrightarrow Hg^{2+}(aq) + 2e^{-}$
Platinum	$Pt(s) \longrightarrow Pt^{2+}(aq) + 2e^{-}$
Gold	$Au(s) \longrightarrow Au^{3+}(aq) + 3e^{-}$

Ease of oxidation increases



Predict if the Following Reactions will occur

$Mg(s) + Al(NO_3)_{3(aq)} \rightarrow$

$Zn(s) + Na_2SO_{4(aq)} \rightarrow$

$Sn(s) + HCl_{(aq)} \rightarrow$

$Au(s) + HCl_{(aq)} \rightarrow$

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Calcium	$Ca(s) \rightarrow Ca^{2+}(aq) + 2e^{-}$
Sodium	$Na(s) \rightarrow Na^{+}(aq) + e^{-}$
Magnesium	$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^{-}$
Aluminum	$Al(s) \rightarrow Al^{3+}(aq) + 3e^{-}$
Manganese	$Mn(s) \rightarrow Mn^{2+}(aq) + 2e^{-}$
Zinc	$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$
Chromium	$Cr(s) \rightarrow Cr^{3+}(aq) + 3e^{-}$
Iron	$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$
Cobalt	$Co(s) \rightarrow Co^{2+}(aq) + 2e^{-}$
Nickel	$Ni(s) \rightarrow Ni^{2+}(aq) + 2e^{-}$
Tin	$Sn(s) \rightarrow Sn^{2+}(aq) + 2e^{-}$
Lead	$Pb(s) \rightarrow Pb^{2+}(aq) + 2e^{-}$
Hydrogen	$H_2(g) \rightarrow 2H^{+}(aq) + 2e^{-}$
Copper	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-}$
Silver	$Ag(s) \rightarrow Ag^{+}(aq) + e^{-}$
Mercury	$Hg(l) \rightarrow Hg^{2+}(aq) + 2e^{-}$
Platinum	$Pt(s) \rightarrow Pt^{2+}(aq) + 2e^{-}$
Gold	$Au(s) \rightarrow Au^{3+}(aq) + 3e^{-}$

Ease of oxidation increases

Molarity

- Quantifies concentration of a solution
- Basic Equation

$$\text{Molarity} = \frac{\text{moles solute}}{\text{volume of solution in liters}}$$

Molarity – *practice problem*

- How many grams of Na_2SO_4 are there in 15 mL of 0.50 M Na_2SO_4?

- How many mL of 0.50 M Na_2SO_4 solution are required to supply 0.038 mol of this salt?

Dilution

- Dillution is used to reduce the concentration of stock solutions
- Add water to concentrated stock solutions to obtain a solution of lower concentration
- Key idea: the number of moles of solute remains unchanged..so

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

Dilution – practice problem

- How many milliliters of 5.0 M $K_2Cr_2O_7$ solution must be diluted in order to prepare 250 mL of 0.10 M solution?

Section 4.6 Titrations

- We will skip this for now and cover titration later in the year