Chem 103, Section F0F
Unit V - Chemical Reactions and Chemical Properties
Lecture 13

- Mixtures
- Water as the solvent in a solution mixture
- Reactions of ionic compounds in solution
- Reactions which form precipitates

Lecture 13 - Reactions

Reading in Silberberg

- Chapter 2, Section 9
  - *Mixtures: Classification and Separation*
- Chapter 4, Section 1
  - *The Role of Water as a Solvent*
- Chapter 4, Section 2
  - *Writing Equations for Aqueous Ionic Reactions*
- Chapter 4, Section 3
  - *Precipitation Reactions*

Lecture 13 - Introduction

Chemical change involves chemical reactions

- According to Dalton's postulates
  - *Atoms of one element **cannot** be converted to atoms of another element. In chemical reactions, the atoms of the original substances recombine to form different substances.*
  - *Compounds result from the chemical combination of a specific ratio of atoms of different element.*

Lecture 13 - Mixtures

Most matter in nature exists as mixture.

- You are a mixture!

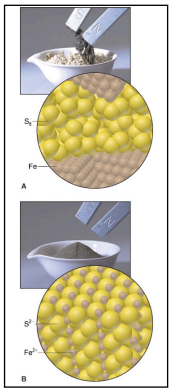
Classes of mixtures

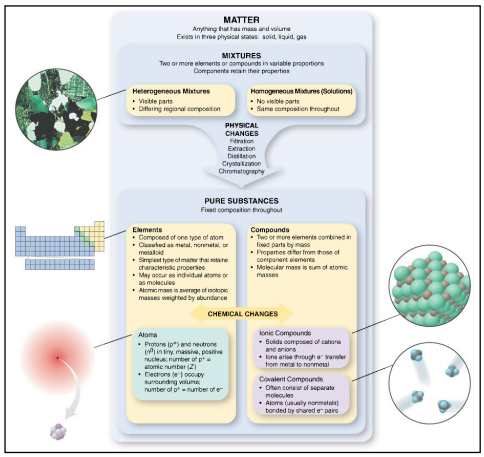
- Heterogenous mixtures
  - Composition is not uniform
- Homogeneous mixtures
  - Composition is uniform down to the molecular level
  - Solutions are homogeneous mixtures

Lecture 13 - Mixtures

Some properties of mixtures:

- Unlike pure substances, the composition can vary.
- The properties of the individual components can still be observed.
- The components of the mixture can be separated using physical means, as opposed to chemical means.



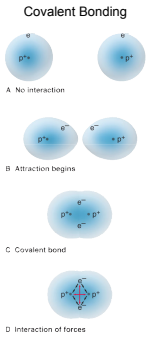


Lecture 13 - Water

We will examine three different types of reactions, which occur in water:

- Therefore, we need to first understand the nature of water.

Water is a polar, covalent molecule.



Lecture 13 - Water

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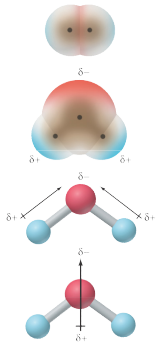
- Therefore, we need to first understand the nature of water.

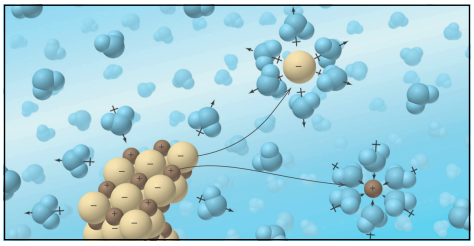
Water is a polar, covalent molecule.

- Even though water is not ionic, it has a positive (δ+) and a negative side (δ-).
  - This is what makes it polar.



Lecture 13 - Water

Ionic compounds dissolve in water because the polar water molecules are able to disrupt the ion ionic interactions.



Lecture 13 - Water

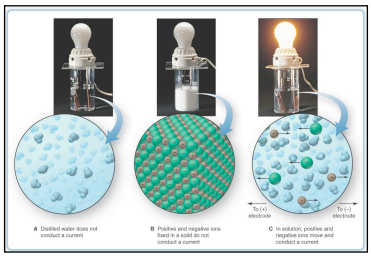
Ionic compounds dissolve in water because the polar water molecules are able to disrupt the ion ionic interactions.

- Not all ionic substances dissolve well in water
  - Solubility of NaCl in H_2O at 20°C = 365 g/L
  - Solubility of AgCl in H_2O at 20°C = 0.009 g/L

Lecture 13 - Water

Ionic compounds dissolve in water because the polar water molecules are able to disrupt the ion ionic interactions.

- The presence of ions in solution can be determined by measuring a solution's conductivity.



A substance that is able to increase the conductivity of water is called an **electrolyte.**

Lecture 13 - Clicker Questions 1

What is the molar concentration of sulfate ions in a 0.25 M solution of ammonium sulfate?

A) 0.0 M
B) 0.25 M
C) 0.50 M
D) 0.75 M

Lecture 13 - Clicker Questions 2

What is the molar concentration of ammonium ions in a 0.25 M solution of ammonium sulfate?

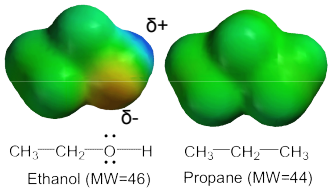
A) 0.0 M
B) 0.25 M
C) 0.50 M
D) 0.75 M

Lecture 13 - Water

Some covalent (molecular) compounds are also able to dissolve in water.

- Like water, these compounds tend to be polar

Ethanol,which is polar, is soluble in water



$CH_3—CH_2—\ddot{O}—H$ Ethanol (MW=46)

$CH_3—CH_2—CH_3$ Propane (MW=44)

Propane,which is nonpolar, is not soluble in water

Lecture 13 - Water

Some covalent (molecular) compounds are also able to dissolve in water.

- Like water, these compounds tend to be polar

Because they are not charged, they do not conduct electricity.

- Such molecules are called **nonelectrolytes.**

Lecture 13 - Water

Later we will see that there is a subgroup of covalent compounds that dissolve and interact strongly enough with water, to lose a hydrogen ion to become form ions.

- These substances are electrolytes.
- They are called **acids.**

hydrochloric acid: $HCl + H_2O \longrightarrow H_3O^+ + Cl^-$

acetic acid: $CH_3COOH + H_2O \longrightarrow H_3O^+ + CH_3COO^-$

Lecture 13 - Equations for Ionic Reactions

Writing equations for ionic reactions is something that we have covered extensively in lab.

- There are three types of equations that are written:
  - Molecular equation
  - Total ionic equation
  - Net ionic equation
- We have seen that the net ionic equation allows us to observe the actual reaction, if any, that is taking place.

Lect tions

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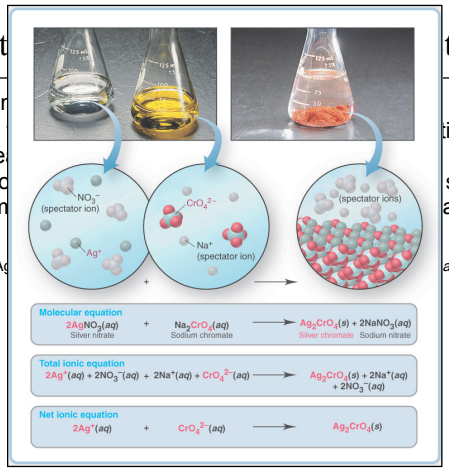
- Write tions for the re aquec silver chrom aNO3).

Ag aq)



Lecture 13 - Equations for Ionic Reactions

For example,

- Write the molecular, total ionic, and net ionic equations for the reaction of aqueous silver nitrate ($AgNO_3$) with aqueous sodium chromate (Na_2CrO_4) to form solid silver chromate (Ag_2CrO_4) and aqueous sodium nitrate ($NaNO_3$).

Molecular

$$2\,AgNO_3(aq) + Na_2CrO_4(aq) \longrightarrow Ag_2CrO_4(s) + 2\,NaNO_3(aq)$$

Total Ionic

$$2Ag^+(aq) + 2\,NO_3^-(aq) + 2\,Na^+(aq) + CrO_4^{-2}(aq) \longrightarrow Ag_2CrO_4(s) + 2\,Na^+(aq) + 2\,NO_3^-(aq)$$

Net Ionic

$$2Ag^+(aq) + CrO_4^{-2}(aq) \longrightarrow Ag_2CrO_4(s)$$

- The ions that do not appear in the net ionic equation are called the **spectator ions**

Lecture 13 - Equations for Ionic Reactions

If all the species turn out the spectator ions, then there is no reaction.

In these reations, the ions are switching partners

- This type of reaction is called an **exchange**, **double-displacement** or **metathesis** reaction.
- For a reaction to occur, one of the new ion pairs must be insoluble.
  - Solubility rules can be used to predict whether a precipitate will form.

Lecture 13 - Equations for Ionic Reactions

Solubility rules can be used to predict whether a precipitate will form.

Table 4.1 Solubility Rules for Ionic Compounds in Water

Soluble Ionic Compounds	Insoluble Ionic Compounds
1. All common compounds of Group 1A(1) ions (Li^+, Na^+, K^+, etc.) and ammonium ion (NH_4^+) are soluble.	1. All common metal hydroxides are insoluble, *except* those of Group 1A(1) and the larger members of Group 2A(2) (beginning with Ca^{2+}).
2. All common nitrates (NO_3^-), acetates (CH_3COO^- or $C_2H_3O_2^-$), and most perchlorates (ClO_4^-) are soluble.	2. All common carbonates (CO_3^{2-}) and phosphates (PO_4^{3-}) are insoluble, *except* those of Group 1A(1) and NH_4^+.
3. All common chlorides (Cl^-), bromides (Br^-), and iodides (I^-) are soluble, *except* those of Ag^+, Pb^{2+}, Cu^+, and Hg_2^{2+}. All common fluorides (F^-) are soluble, *except* those of Pb^{2+} and Group 2A(2).	3. All common sulfides are insoluble *except* those of Group 1A(1), Group 2A(2), and NH_4^+.
4. All common sulfates (SO_4^{2-}) are soluble, *except* those of Ca^{2+}, Sr^{2+}, Ba^{2+}, Ag^+, and Pb^{2+}.	

Lecture 13 - Questions 3

Predict whether a reaction occurs, and write a balanced total and net ionic equation:

B) Sodium hydroxide(*aq*) + iron(III) sulfate(*aq*) →

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Unit IV - Up Next

Lecture 14 - Reactions con'd

- Acid-base reactions
- Oxidation-reduction reactions
- Elements in redox reactions
- Reversibility of reactions

The End