

CHE 1301

Basic Principles of Modern Chemistry I

Week 11

Hi! Thanks for checking out the weekly resources for Chemistry 1301! This resource covers topics typically taught by professors during the 11th week of classes.

Visit our website, <https://baylor.edu/tutoring>, to sign up for appointments and check out additional resources for your course! You'll find helpful links with the following titles:

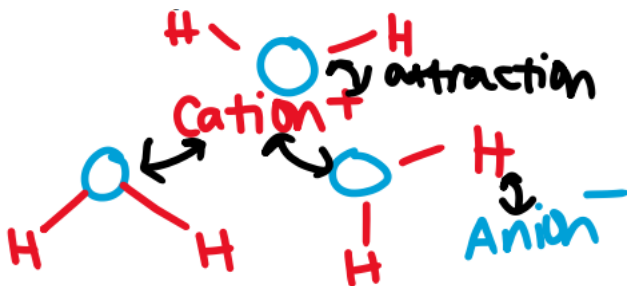
- "Online Study Guide Resources" – The pace of your course may vary slightly from what's shown in this document. If you don't see the topics you're learning right now, use this link to find the weekly resources for the rest of the semester.
- "How to Participate in Group Tutoring" - See if there is a Chemistry 1301 group tutoring session being hosted this semester. These are weekly question/answer sessions taught by our master tutors!
- "View tutoring times for your course" or "Schedule a private 30-minute appointment!"

You can also give us a call at (254)710-4135, or drop in! Our hours are Monday-Thursday 9 am – 8 pm on class days.

KEY WORDS: Precipitation Reaction, Neutralization Reaction, Titration, Oxidation & Reduction

TOPIC OF THE WEEK: Writing Ionic Equations

When ions are put into solution, the properties of water cause them to dissolve. Water has partial positive and negative charges, and since ions are charged, they attract to the oppositely charged poles of the water molecules. **Anions attract to different poles than cations, which causes them to separate:**



This is what the subscripted (aq) means in a chemical formula: The ions have been dissolved in water to form an [aqueous] solution.

So what if you mix two aqueous solutions together? The ions are still dissolved, but now, there are other cations and other anions in solution that could potentially interact with the cations and anions already

there. Sometimes, water is strong enough to hold them apart, just like it was strong enough to hold the original two apart.

But sometimes, an old cation and a new anion (or vice versa) will have a strong enough force of attraction that water cannot overcome it. This is called a **precipitation reaction**. How do you know which cases these are? See the next page for a table created by Flinn Scientific:

<https://www.flinnsci.ca/solubility-rules-chart-notebook-size-pad-of-30/ap7226/>

**Solubility Table
Common Ionic Compounds**

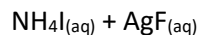
	Group 1				Group 2			Transition Metals					
	NH ₄ ⁺	Li ⁺	Na ⁺	K ⁺	Mg ²⁺	Ca ²⁺	Ba ²⁺	Al ³⁺	Fe ³⁺	Cu ²⁺	Ag ⁺	Zn ²⁺	Pb ²⁺
F ⁻	sol	sol	sol	sol	insol	insol	sl sol	sol	sl sol	sol	sol	sol	insol
Cl ⁻	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	insol	sol	sol
Br ⁻	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	insol	sol	sol
I ⁻	sol	sol	sol	sol	sol	sol	sol	sol			insol	sol	insol
OH ⁻	sol	sol	sol	sol	insol	sl sol	sol	insol	insol	insol		insol	insol
S ²⁻	sol	sol	sol	sol		sol			insol	insol	insol	insol	insol
SO ₄ ²⁻	sol	sol	sol	sol	sol	sl sol	insol	sol	sol	sol	sl sol	sol	insol
CO ₃ ²⁻	sol	sol	sol	sol	insol	insol	insol				insol	insol	insol
NO ₃ ⁻	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol
PO ₄ ³⁻	sol	sol	sol	sol	insol	insol	insol	insol	insol	insol	insol	insol	insol
CrO ₄ ²⁻	sol	sol	sol	sol	sol	sol	insol		insol	insol	insol	insol	insol
CH ₃ CO ₂ ⁻	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol	sol

sol — soluble
sl sol — slightly soluble
insol — insoluble
(blank) — compound does not exist

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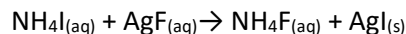
Writing an Ionic Equation:

When you're given reactants and told to write the precipitation reaction, here are some steps you can follow:



- Separate** the compounds into ions
 - $\text{NH}_4^+ + \text{I}^- + \text{Ag}^+ + \text{F}^-$
- Check the solubility** of all cations with all anions.
 - Original combinations (NH₄I & AgF) we know to be soluble
 - NH₄F: soluble
 - AgI: insoluble
- If there are any combinations that are listed as "insoluble," you can assume that these two ions will find each other in solution and form a **precipitate**. Write a subscript (s) to denote "solid." Any other combinations of ions can be listed together as being in the aqueous state.
 - (If there are no combinations of ions that will form a precipitate, write "no reaction" on the right side of the equation.)
 - $\text{NH}_4\text{I}_{(\text{aq})} + \text{AgF}_{(\text{aq})} \rightarrow \text{NH}_4\text{F}_{(\text{aq})} + \text{AgI}_{(\text{s})}$
- Balance** the equation if necessary.

Writing a Net Ionic Equation:



1. Separate all compounds into their ions (solids will no longer separate into ions!)
 - a. $\text{NH}_4^+ + \text{I}^- + \text{Ag}^+ + \text{F}^- \rightarrow \text{NH}_4^+ + \text{AgF}_{(\text{s})} + \text{I}^-$
2. Some of these ions are on both sides of the equation—in other words, they didn't react. Rewrite the equation **without these spectator ions**.
 - a. $\text{Ag}^+ + \text{F}^- \rightarrow \text{AgF}_{(\text{s})}$

Highlight 1: Acids and Bases – Neutralization Reactions

Arrhenius proposed that acids, or the group of compounds with distinct characteristics such as a sour taste and ability to dissolve metals, are those that increase the concentration of H^+ when dissolved in water. Similarly, he proposed that bases were compounds that increase the concentration of OH^- when dissolved in water.

However, some acidic and basic solutions were formed by compounds that didn't contain H^+ and OH^- ! So, Bronsted and Lowry came up with new requirements:

Acid: Bronsted-Lowry definition = **proton donor**

Base: Bronsted-Lowry definition = **proton acceptor**

When a solution of strong acid and a solution of strong base are mixed, the acid donates all of its protons to the base, which is called a neutralization reaction.

When given the reactants for a neutralization reaction, write the products as follows:

1. **Identify the acid and the base.** The acid has an available proton. Sometimes, bases have protons too! You'll get more into this in organic chem (<https://www.khanacademy.org/science/organic-chemistry/organic-structures/acid-base-review/v/acid-base-lewis-bronsted-lowry>) but I'd say the best way for now is to memorize the acids and bases that come up most often! See table below

TABLE 8.2 Some Common Acids and Bases

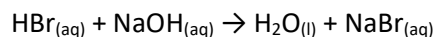
Name of Acid	Formula	Name of Base	Formula
Hydrochloric acid	HCl	Sodium hydroxide	NaOH
Hydrobromic acid	HBr	Lithium hydroxide	LiOH
Hydroiodic acid	HI	Potassium hydroxide	KOH
Nitric acid	HNO_3	Calcium hydroxide	$\text{Ca}(\text{OH})_2$
Sulfuric acid	H_2SO_4	Barium hydroxide	$\text{Ba}(\text{OH})_2$
Perchloric acid	HClO_4	Ammonia*	NH_3 (weak base)
Acetic acid	$\text{HC}_2\text{H}_3\text{O}_2$ (weak acid)		
Hydrofluoric acid	HF (weak acid)		

2. A couple things happen at once:

- a. H^+ is lost. Imagine a hydrogen atom, one proton and one electron. Its electron is being used in a bond to the rest of the molecule... this electron will stay, but the proton will be taken. Since the molecule has **lost a positive proton**, its charge **decreases by 1**.
 - b. The other molecule picks up H^+ . It has **gained the positive proton**, which makes its charge **increase by 1**.
3. The results are “neutral” – the products don’t act like acids or like bases!
- a. Important to note: **if you start with an acid and a metal hydroxide, products will be 1) water and 2) salt.**

Here’s an example: $HBr_{(aq)} + NaOH_{(aq)} \rightarrow ?$

Think: We’re dealing with a metal hydroxide: the products will be water and a salt. So, take away the ingredients for water— H and OH —and make water with them. The remaining ions will combine as a salt.



(Net ionic equation: $H^+ + OH^- \rightarrow H_2O$)

***Make sure that the equation is balanced!**

Highlight 2: Acids and Bases - Titration

Molarity = (# mol solute)/(volume of solution)

See Chem 2: Week 10 for more on these calculations

Titration is a way to determine unknown concentrations. You start out knowing a couple of things:

1. The reaction is a neutralization reaction.
 - a. The unknown is either an acid or a base, and the solution with a known concentration is the opposite.
 - b. Let’s say we’re adding an acid to a base. When there’s exactly enough acid to react with base, the result is a neutral solution—this is called the **equivalence point**.
 - c. After we’ve added enough acid to react with the base, the acid will go on to react with the indicator, and the indicator will change color! This is called the **endpoint**. This is just barely past the equivalence point, so now we can estimate it.
2. We also know the **concentration of one solution**—either acidic or basic.
 - a. This can be used to calculate the number of **moles in the known solution that we started with**.
3. The **number of moles in the known solution that we started with** is the **same as the number of moles that had to be added to reach the equivalence point**. We know how many moles were added, and we know the **volume that we added**, so we know **concentration**.

Highlight 3: Identifying Oxidation and Reduction

In a redox reaction, one compound is reduced, and one is oxidized. But what does this mean?

Reduction is an atom's gain of electrons; oxidation is the loss of electrons.

Oxidation number is a way to calculate this. If an atom increases in oxidation number, it has been oxidized. If an atom decreases in oxidation number, it has been reduced.

Follow the rules in this order—if two of them conflict with each other, choose to follow the one that is listed first.

When the atoms are on their own:

Free atom	0
Monoatomic ions	Charge of ion

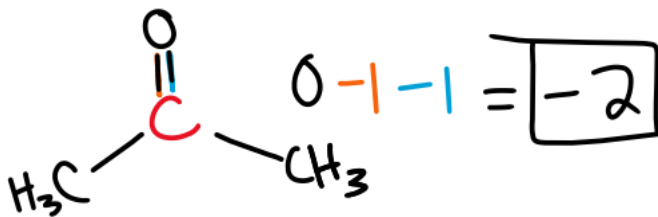
When the atoms are bonded to other atoms:

Sum of the oxidation numbers in a compound = charge of compound

Group IA and IIA metals, aluminum	+1, +2, +3
Hydrogen	+1 when bonded to nonmetals, -1 when bonded to metals
Oxygen	-2 unless H_2O_2 (-1) or O_2^{2-} (-1)
Fluorine	-1
Other halogens	-1

Another way to think about this:

Start with zero. The oxidation number decreases by 1 for every bond to a more electronegative element. It increases by 1 for every bond to a less electronegative element. If it has a bond to the same element, it will stay the same. Here's an example!



In a redox reaction, one atom gets oxidized and the other gets reduced. Here's an example: $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$

1. Calculate oxidation numbers of each atom.
 - a. Left side
 - i. K is always +1.
 - ii. O is always -2.

- iii. The sum of all the oxidation numbers in the molecule is zero. There are 3 oxygen atoms, so the total oxidation number from oxygen is -6. $-6 + 1 + \text{chlorine oxidation number} = 0$. Chlorine oxidation number = +5.
 - b. Right side
 - i. K is always +1
 - ii. Cl is -1
 - iii. Check that the sum of these two equals zero—it does!
 - iv. Sum of all the oxidation numbers in the oxygen molecule is 0. So, each of them must have an oxidation number of 0.
 - c. **Compare each atom's oxidation-number-before to its oxidation-number-after.**
 - i. K is the same
 - ii. O increased \rightarrow O was oxidized
 - iii. Cl decreased \rightarrow Cl was reduced
-

Check Your Understanding

1. Given reactants HCl and $\text{Ba}(\text{OH})_2$, write a balanced ionic equation.
 2. You used 30 mL of acid solution to reach equivalence point, and you know that you started with 100 mL of 2 M basic solution. What is the molarity of the acidic solution?
 3. Identify reactant that is oxidized and reactant that is reduced in:
 - a. $\text{Cl}_2 + \text{SnCl}_2 \rightarrow \text{SnCl}_4$
-

Things You May Struggle With

1. When writing a precipitation reaction, if neither reactant precipitates, no reaction occurs. Writing each cation with the opposite anion is equivalent to what you had before! Just write "no reaction" after the reaction arrow.
2. Endpoint and equivalence point have 2 different meanings. Endpoint happens after equivalence point, when the solution being added has already reacted with all of the acid/base present.

That's all this week! Please reach out if you have any questions and don't forget to visit the Tutoring Center website for further information at www.baylor.edu/tutoring. Answers to Check Your Learning are below.

1. $2\text{HCl} + \text{Ba}(\text{OH})_2 \rightarrow \text{BaCl}_2 + 2\text{H}_2\text{O}$
2. 0.2 mol (.1 L * 2 mol/L) of basic solution must have reacted with 0.2 mol of acidic solution. 0.2 mol acidic solution per .03 L (30 mL) of acidic solution means 6.7 M
3. Sn is oxidized. Some of the Cl is reduced.