

Chemistry 11
Solution Chemistry II

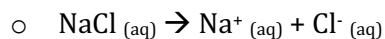
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| <ol style="list-style-type: none">1. Ions in Solutions2. Solubility Table3. Separating Ions |
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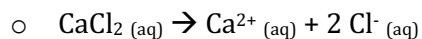
Ions in Solutions

Ionization Equation

- Represents the salt breaking apart into ions.



- If the salt were CaCl_2 .



Practice:

1. KOH

2. SrCl_2

3. NH_4NO_3

4. Ag_2CrO_4

5. $\text{Pb}(\text{IO}_3)_2$

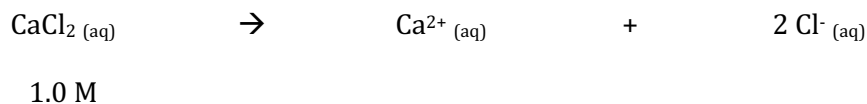
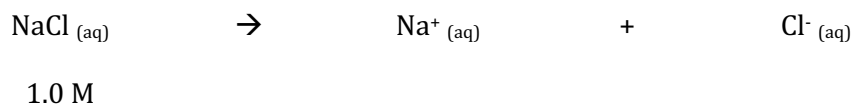
6. $\text{Mg}(\text{OH})_2$

7. BaCO_3

8. $(\text{NH}_4)_2\text{SO}_3$

Calculating Concentration

- Mole ratios represent the relative amounts of ions in solution.



Practice:

1. 0.25 M KOH

2. 0.75 M SrCl₂3. 0.20 M NH₄NO₃4. 0.45 M Ag₂CrO₄5. 0.60 M Pb(IO₃)₂6. 0.011 M Mg(OH)₂7. 0.45 M BaCO₃8. 0.50 M (NH₄)₂SO₃**Ionization + Dilution**

- ✓ A 15.0 mL sample of 3.0 M HCl was added to 10.0 mL of 2.0 M CaCl₂. Calculate the concentration of each ion in the solution. Assume no reaction occurs ($[H^+] = 1.8 M$ $[Cl^-] = 3.4 M$ $[Ca^{2+}] = 0.80M$)

Practice 1.

What are the concentrations of both ions in a 3.00 M solution of H_3PO_4 ? ($[\text{H}^+] = 9.00\text{M}$ $[\text{PO}_4^{3-}] = 3.00\text{M}$)

Practice 2.

What is the sodium ion concentration when 250.0 mL of water is added to 125.5 mL of a 3.21 M solution of sodium phosphate? ($[\text{Na}^+] = 3.21\text{M}$)

Practice 3.

Determine the concentration of each ion when 45.0 mL of 7.20 M magnesium sulphate is mixed with 35.0 mL of 0.900 M magnesium hydroxide. ($[\text{Mg}^{2+}] = 4.44\text{M}$ $[\text{SO}_4^{2-}] = 4.05\text{M}$. $[\text{OH}^-] = 0.788\text{M}$)

Practice 4.

What is the molar concentration of each ion in solution resulting from mixing 55.0 mL of 0.15 M potassium hydroxide with 75.0 mL of 0.25 M potassium sulphate? ($[\text{K}^+] = 0.34\text{M}$ $[\text{OH}^-] = 0.063\text{M}$ $[\text{SO}_4^{2-}] = 0.14\text{M}$)

Solubility Table

When some ions are combined, they create a solid →
they are NOT soluble (will form a _____)

SOLUBLE → Dissolves in water; aqueous
LOW SOLUBILITY → Does not dissolve in water; solid (ppt)

SOLUBILITY OF COMMON COMPOUNDS IN WATER

The term soluble here means > 0.1 mol/L at 25°C.

Negative Ions (Anions)	Positive Ions (Cations)	Solubility of Compounds
All	Alkali ions: Li ⁺ , Na ⁺ , K ⁺ , Rb ⁺ , Cs ⁺ , Fr ⁺	Soluble
All	Hydrogen ion: H ⁺	Soluble
All	Ammonium ion: NH ₄ ⁺	Soluble
Nitrate, NO ₃ ⁻	All	Soluble
Chloride, Cl ⁻ or Bromide, Br ⁻ or Iodide, I ⁻	All others	Soluble
	Ag ⁺ , Pb ²⁺ , Cu ⁺	Low Solubility
Sulphate, SO ₄ ²⁻	All others	Soluble
	Ag ⁺ , Ca ²⁺ , Sr ²⁺ , Ba ²⁺ , Pb ²⁺	Low Solubility
Sulphide, S ²⁻	Alkali ions, H ⁺ , NH ₄ ⁺ , Be ²⁺ , Mg ²⁺ , Ca ²⁺ , Sr ²⁺ , Ba ²⁺	Soluble
	All others	Low Solubility
Hydroxide, OH ⁻	Alkali ions, H ⁺ , NH ₄ ⁺ , Sr ²⁺	Soluble
	All others	Low Solubility
Phosphate, PO ₄ ³⁻ or Carbonate, CO ₃ ²⁻ or Sulphite, SO ₃ ²⁻	Alkali ions, H ⁺ , NH ₄ ⁺	Soluble
	All others	Low Solubility

Practice:

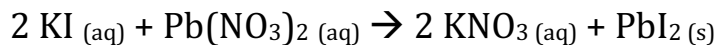
1. Classify the following salts as being soluble or having low solubility in water:
 - a. Sodium phosphate
 - b. Aluminum hydroxide
 - c. * Copper (II) chloride
 - d. Calcium sulphate
 - e. Iron (II) sulphide
 - f. Strontium hydroxide
 - g. Zinc bromide
 - h. Cesium sulphite
 - i. Potassium chromate

2. Write the formula for the following:

- a. A salt containing carbonate that is soluble
- b. A salt containing sulphate with low solubility
- c. A cation that forms a salt with low solubility with both chloride and sulphate ions
- d. An anion that forms soluble salts with all cations

Types of chemical equations:

Formula Equation: shows the chemical formulas of the compounds and their states.



Complete Ionic Equation: shows the soluble salts represented in their dissociated form.

Net Ionic Equation: shows only the ions that take part in the reaction. Ions that are the same on both sides of the equation are called **spectator ions**.

Practice:

1. Write the formula for the precipitate that forms when the following solutions are mixed:

a. BaS and MgSO₄

b. NH₄OH and FeBr₂

c. H₃PO₄ and ZnCl₂

d. K₂CO₃ and CrSO₄

e. MnI₂ and Sr(OH)₂

2. Write a formula equation, complete ionic equation and net ionic equation for the following reactions:

a. Strontium hydroxide and silver nitrate

b. Magnesium sulphide and zinc chloride

c. Sodium carbonate and barium sulphide

d. $(\text{NH}_4)_2\text{S}_{(\text{aq})} + \text{FeSO}_4_{(\text{aq})} \rightarrow$

e. $\text{H}_2\text{SO}_3_{(\text{aq})} + \text{CaCl}_2_{(\text{aq})} \rightarrow$

f. Copper (II) sulphate + $\text{H}_2\text{SO}_3 \rightarrow$

Separating Ions

Example:

A solution may contain the ions Ca^{2+} , Sr^{2+} and Zn^{2+} . How would you precipitate the ions out of solution individually? Describe your answer using a flow chart.

- All are cations - therefore an addition of an anion will precipitate out these cations.
- “Low solubility” – means will precipitate out.

1. Which anion will precipitate just one of the ions out first?

a. Which ions are left?

2. Which anion will precipitate just one of the two ions left?

3. Which anion will precipitate out the last ion left?

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All	Ammonium ion: NH_4^+	Soluble
Nitrate, NO_3^-	All	Soluble
Chloride, Cl^- or Bromide, Br^- or Iodide, I^-	All others	Soluble
	Ag^+ , Pb^{2+} , Cu^+	Low Solubility
Sulphate, SO_4^{2-}	All others	Soluble
	Ag^+ , Ca^{2+} , Sr^{2+} , Ba^{2+} , Pb^{2+}	Low Solubility
Sulphide, S^{2-}	Alkali ions, H^+ , NH_4^+ , Be^{2+} , Mg^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+}	Soluble
	All others	Low Solubility
Hydroxide, OH^-	Alkali ions, H^+ , NH_4^+ , Sr^{2+}	Soluble
	All others	Low Solubility
Phosphate, PO_4^{3-} or Carbonate, CO_3^{2-} or Sulphite, SO_3^{2-}	Alkali ions, H^+ , NH_4^+	Soluble
	All others	Low Solubility

Practice:

1. For each of the following solutions, describe a process to individually remove each ion. Be sure to list the compounds that you add in order, and the method of removing the precipitate. You may wish to use a flow chart.

