Introduction

A. Ionic vs Molecular Solutions

1. Ionic Compounds form Ionic Solutions

a) **Ionic** compounds (metal + non-metal) dissolved in water to form **Ionic Solutions**

**eg1:** "dissociation" equation

\[
\text{AlCl}_3(s) \rightarrow \text{Al}^{3+}(aq) + 3\text{Cl}^-(aq)
\]

**Ionic compound**  
**free ions / electrolytes**

**Electrolyte:** A substance which dissolves to give an electrically conducting solution containing ions. (e.g. AlCl\(_3\) above)

b) Compounds containing polyatomic ions form Ionic solutions.

**eg:** KMnO\(_4\)

\[
\text{KMnO}_4(s) \rightarrow \text{K}^+(aq) + \text{MnO}_4^-(aq)
\]

2. Covalent Compounds form Molecular Solutions

a) **Covalent** Compounds (Non-metal + Non-metal) generally form **Molecular** solutions.

**eg.** \(\text{I}_2(s) \rightarrow \text{I}_2(aq)\) \(\text{C}_2\text{H}_2(g) \rightarrow \text{C}_2\text{H}_2(aq)\)

- do NOT break up into ions when dissolve:
- are called non **electrolytes**

b) **Organic compounds** form molecular solutions.

**eg:** \(\text{C}_{12}\text{H}_{22}\text{O}_{11}(s) \rightarrow \text{C}_{12}\text{H}_{22}\text{O}_{11}(aq)\)

- exception: organic acids (they end in the "COOH")

What Causes Electrical Conductivity?

- A liquid must have **IONS** in it to conduct electrical current
- Dissolved ions move **freely** and carry a **charge**, therefore can **transfer** electrical current
SUMMARY

1. Compounds made up of a **metal** and a **non-metal** will form **ionic** solutions.
2. Compounds containing **polyatomic ions** will form **ionic** solutions.
3. Compounds containing **only non-metals** (covalent compounds) will form **molecular** solutions.
4. **Organic** compounds (other than those ending in the "COOH" group) will form **molecular** solutions.
5. **Organic acids** (organic compounds ending in the "COOH" group) will form **partially ionic** solutions. (These are ionic but because there are fewer ions formed, these solutions are weak conductors or weak electrolytes.)

### B. Solubility and Solubility Equilibrium

E.g. Put solid ionic solid (crystal) in water:

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Initially, no ions in the water, so rate of dissolving is high.

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no free ions will go back onto the crystal, so rate of precipitation is zero.

1. **Ionic Solid → Aqueous Solution**  (**dissolving**)  
   - At first, only the **forward reaction** is taking place:
   - Eventually, solution fills with free ions and forward rate begins to decrease.

2. **Ionic Solid ↔ Aqueous Solution**  (**precipitation** or **crystallization**)  
   - The number of free ions that collide and stick to the crystal increases.
   - **This is the reverse reaction:**
   - Eventually, reverse rate matches forward rate, and we reach **Solubility Equilibrium**

\[
\text{CaCO}_3(\text{s}) \rightleftharpoons \text{Ca}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq})
\]

- **Solubility Equilibrium**: The rate of **dissolving** = The rate of **precipitation**
- A saturated solution is a solution in which there exists a dissolved substance in equilibrium with the undissolved substance.
- **Solubility** is the **equilibrium concentration** of a substance in a **solution at a given temperature**
- The **Solubility** of a substance is the maximum amount that can dissolve in a given amount of solution at a given temperature. (the "ability" to dissolve)
- **Solubility** increases with **temperature**
• **Molar solubility:** when concentration is expressed in \textit{moles/L} or g/L

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**WRITING FORMULA, COMPLETE AND NET IONIC EQUATIONS**

eg. Mix solution of AgNO\(_3\) with a solution of Na\(_2\)CO\(_3\).

- **Formula equation** - balanced eqn with all reactants and products:
  \[
  2\text{AgNO}_3 (aq) + \text{Na}_2\text{CO}_3 (aq) \rightarrow \text{Ag}_2\text{CO}_3(s) + 2\text{NaNO}_3 (aq)
  \]

- **Complete ionic equation** - balanced eqn showing all ions:
  \[
  2\text{Ag}^+(aq) + 2\text{NO}_3^-(aq) + 2\text{Na}^+(aq) + \text{CO}_3^{2-}(aq) \rightarrow 2\text{Ag}^+(s) + \text{CO}_3^{2-}(s) + 2\text{Na}^+(aq) + 2\text{NO}_3^-(aq)
  \]

- Only Ag and CO\(_3\) are reacting and precipitating. How do we know which ions will precipitate? (must look to Ksp value),
- the other ions are "spectator ions"

- **Net ionic equation** - balanced eqn showing only species involved in the reaction:
  \[
  2\text{Ag}^+(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{Ag}_2\text{CO}_3(s)
  \]

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**PREDICTING THE SOLUBILITY OF SALTS**

1. **Definitions**

"Low Solubility" = a precipitate \textit{will form}.
"Soluble" = a precipitate \textit{won't form}.

A substance has \textit{low solubility} if a saturated solution of the substance has a \textit{concentration less than} 0.1 M

2. **Finding the solubility of a compound**

To find this, we use a Solubility Table (Hebden p.332)

1. Find anion. (column 1)
2. Look for matching cation (column 2)
3. Answer (column 2)

<table>
<thead>
<tr>
<th>Anion</th>
<th>Soluble</th>
<th>Low Solubility</th>
</tr>
</thead>
<tbody>
<tr>
<td>FeCO(_3)</td>
<td></td>
<td>√</td>
</tr>
<tr>
<td>CuI</td>
<td></td>
<td>√</td>
</tr>
<tr>
<td>PbBr(_2)</td>
<td></td>
<td>√</td>
</tr>
</tbody>
</table>

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**SEPARATING MIXTURES OF IONS BY PRECIPITATION METHODS**

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}\text{Hebden}\ #1-7

}\text{Hebden}\ #25

}\text{Hebden}\ #21-25
eg1: A solution contains both Ag⁺ and Ba²⁺. How do you separate these cations?

• Use Solubility Table to find a negative ion which will form a precipitate with one … but not both!

∴ if you add Cl⁻ or Br⁻ or I⁻ the precipitate should contain Ag⁺ and the solution should still contained dissolved Ba²⁺.

ALSO… SO₄²⁻, would be a bad choice for the above example because both would form a precipitate.

E.g. 2) A solution may contain Ag⁺, Ba²⁺, and Ni²⁺. What ions could be added, and in what order to determine which cations are present?

• make a precipitate table:

<table>
<thead>
<tr>
<th></th>
<th>Cl⁻</th>
<th>SO₄²⁻</th>
<th>S²⁻</th>
<th>OH⁻</th>
<th>PO₄³⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ag⁺</td>
<td>precipitate</td>
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<td>precipitate</td>
<td>precipitate</td>
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<tr>
<td>Ba²⁺</td>
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<td>precipitate</td>
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<td>precipitate</td>
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</tr>
<tr>
<td>Ni²⁺</td>
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<td>precipitate</td>
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</tr>
</tbody>
</table>

Can not use OH⁻ or PO₄³⁻ b/c they will form precipitate w all three cations

**Step 1.**

• Add Cl⁻ first because it forms precipitate only with Ag⁺
• then filter out the AgCl
  - Ag⁺ is gone, so now only consider Ba²⁺ and S²⁻

**Step 2.** (Two Possibilities)

• Can add either SO₄²⁻ which will form a precipitate with only Ba²⁺
  - then filter out the BaSO₄
  - only S²⁻ remains in the solution
• or S²⁻ (forms precipitate w only Ni²⁺)
  - then filter out the NiS
  - only SO₄²⁻ remains in the solution

**Note:** Anions cannot be added by themselves ∴ you should add a spectator cation (Li⁺, Na⁺, K⁺) at the same time

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**REVIEW: DILUTION AND UNITS OF CONCENTRATION**
A. Dilution
dilution formula: \( C_1V_1 = C_2V_2 \)

eg1) Mix 5.0mL of 0.020 Cl₂ solution with 15.0mL of 0.012 M Br⁻ solution. What is the concentration(C₂) of the Br⁻ ion in the new solution?

\[
C_2 = \frac{C_1V_1}{V_2} = \frac{0.012M \times 15.0mL}{15.0mL + 5.0mL} = .0090 \text{ M}
\]

eg2) There are \( 7.0 \times 10^{-4} \) mol [Na₂SO₄] in 100.0 mL of solution. What is the [Na⁺]?

\[
[\text{Na}^+] = \frac{7.0 \times 10^4 \text{ mol Na}_2\text{SO}_4}{0.1000\text{L}} \times \frac{2\text{Na}^+}{1\text{Na}_2\text{SO}_4} = 1.4 \times 10^{-2} \text{ M Na}^+
\]

B. Units of Concentration

eg1) 1L of saturated solution contains 1.96g of AgBrO₃. What is the molar solubility of AgBrO₃?

\[
[\text{AgBrO}_3] = \frac{1.96\text{g AgBrO}_3}{\text{L}} \times \frac{1\text{mol}}{235.8g} = 8.31 \times 10^{-3} \text{ M}
\]

eg2) The molar solubility of BaF₂ is 3.58mol/L. How many grams of BaF₂ exist in 726mL of saturated solution?

\[
\text{Grams of BaF}_2 = \frac{3.58\text{mol BaF}_2}{\text{L}} \times \frac{175.3\text{g BaF}_2}{\text{mol BaF}_2} \times 0.726\text{L} = 456\text{g BaF}_2
\]

THE SOLUBILITY PRODUCT (\( K_{sp} \))

\( K_{sp} \) = the equilibrium constant (\( K_{eq} \)) for an ionic substance dissolving in water.

eg. \( \text{CaCO}_3(s) \) dissolves in water

net-ionic eqn: \( \text{CaCO}_3(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) \).

\[
K_{sp} = [\text{Ca}^{2+}] [\text{CO}_3^{2-}]
\]

([\text{CaCO}_3] \text{ not included -- is a solid})

[\text{Ca}^{2+}] \text{ and } [\text{CO}_3^{2-}] \text{ at equilib are their “solubilities”}

Writing \( K_{sp} \) expression from the Net-Ionic Equation.
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**Solubility Unit Notes**

**Steps**

1. Write out equilibrium equation showing the salt dissolving
2. Write out the equilibrium expression (leave out the solid)

\[
\text{eg1: } \text{PbCl}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq) \\
K_{sp} = [\text{Pb}^{2+}] [\text{Cl}^-]^2
\]

\[
\text{eg2: } \text{Ag}_2\text{SO}_4(s) \rightleftharpoons 2\text{Ag}^+(aq) + \text{SO}_4^{2-}(aq) \\
K_{sp} = [\text{Ag}^+]^2 [\text{SO}_4^{2-}]
\]

**Definitions:**

Solubility: the amount of a substance required to make a saturated solution

Molar solubility: the molar concentration of a saturated solution

Solubility product: the Ksp value obtained when the concentrations of the ions are multiplied together

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**CALCULATING SOLUBILITY & ION CONCENTRATION**

E.g. 1: What is the molar solubility (mol/L) of a saturated AgBrO$_3$(aq) solution if 1 L contains 1.96 g of AgBrO$_3$?

\[
[\text{AgBrO}_3(aq)] = \frac{1.96 \text{ g AgBrO}_3}{1 \text{ L AgBrO}_3} \times \frac{1 \text{ mol AgBrO}_3}{235.8 \text{ g AgBrO}_3} = 8.31 \times 10^{-3} \text{ M}
\]

E.g. 2: Express the molar solubility of PbI$_2$ in g/L.
Molar solubility = 1.37 x 10$^{-3}$

Solubility in g/L = \[
\frac{1.37 \times 10^{-3} \text{ mol PbI}_2}{1 \text{ L PbI}_2} \times \frac{461.0 \text{ g PbI}_2}{1 \text{ mol PbI}_2} = 0.632 \text{ g/mol}
\]

Read examples on p. 80

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**CALCULATIONS WITH K$_{sp}$**

**Calculating Solubility given K$_{sp}$**

E.g. Calculate molar solubility of AgCl. K$_{sp}$ for AgCl = 1.8 x 10$^{-10}$

**Step 1.** Pick a variable and use mole ratios to find other concs:

\[
\text{AgCl(s)} \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq)
\]

**Step 2.** Write out solubility expression.
Chemistry 12

**Step 3.** substitute in variables
Since: \( K_{sp} = [Ag^+] [Cl^-] = [s][s] = s^2 \)

**Step 4.** Solve for \( s \)

\[
K_{sp} = s^2 = 1.8 \times 10^{-10} M
\]

\[
s = \sqrt{1.8 \times 10^{-10}} = 1.34164 \times 10^{-5} M
\]

Answer: the molar solubility of AgCl is \( 1.3 \times 10^{-5} M \)

**Calculating \( K_{sp} \), Given Solubility**

eg. solubility of \( Ag_2CrO_4 \) is \( 1.31 \times 10^{-4} \) moles/L.

Calculate \( K_{sp} \).

**Step 1.** write net-ionic eqn

\[
Ag_2CrO_4 (s) \rightleftharpoons 2Ag^+(aq) + CrO_4^{2-}(aq)
\]

**Step 2.** use mole ratios to find the other concs:

\[
[Ag^+] = 2 \times 1.31 \times 10^{-4} M = 2.62 \times 10^{-4} M
\]

\[
[CrO_4^{2-}] = 1.31 \times 10^{-4} M
\]

**Step 3.** substitute values into \( K_{sp} \) equation:

\[
K_{sp} = [Ag^+]^2 [CrO_4^{2-}] = (2.62 \times 10^{-4})^2 (1.31 \times 10^{-4}) = 8.99 \times 10^{-12}
\]

**Calculating [ion]**

Eg. What is the [\( Mg^{2+} \)] in a saturated solution of Mg(OH)\(_2\)?

\[
Mg(OH)_2 (s) \rightleftharpoons Mg^{2+} (aq) + 2OH^- (aq) \quad K_{sp} = [Mg^{2+}][OH^-]^2
\]

\[
K_{sp} = [Mg^{2+}][OH^-]^2 = (x)(2x)^2 = 4x^3 = 5.6 \times 10^{-12}
\]

\[
x = 1.1 \times 10^{-4}
\]

Therefore, [\( Mg^{2+} \)] = \( 1.1 \times 10^{-4} M \)

**Predicting Precipitates and Maximum Ion Concentration**
A. Predicting Precipitates When Two Solutions Are Mixed

1. Review

The Ksp for CaCO₃ is $5.0 \times 10^{-9}$ 

Ksp = $[Ca^{2+}] [CO_3^{2-}]$

- The Ksp value from above comes from a saturated solution
  - if you try to add more than $5.0 \times 10^{-9} \text{ Ca}^{2+}$ or $\text{CO}_3^{2-}$ a precipitate will form
  - because the solution can’t dissolve any more!

2. Mixing two solutions (ions are from different sources)

E.g. Mixing solutions of Ca(OH)₂ and Na₂CO₃ resulted in the following concentrations:

[Ca²⁺] = $2.3 \times 10^{-4}$ M and [CO₃²⁻] = $8.8 \times 10^{-2}$ M.

Will a precipitate form?

Table shows CaCO₃ Ksp = $5.0 \times 10^{-9}$.

Step 1. Calculate "trial Ksp"

\[
\text{Trial Ksp} = \frac{[Ca^{2+}][CO_3^{2-}]} = \frac{(2.3 \times 10^{-4})(8.8 \times 10^{-2})}{2.024 \times 10^{-5}}
\]

Step 2. Compare the Trial Ksp with the real value for Ksp

- Trial Ksp = $2.024 \times 10^{-5}$
- Ksp = $5.0 \times 10^{-9}$
- Trial Ksp $\geq$ Ksp $\therefore$ precipitate will form

B. Maximum Possible Concentration of an Ion in Solution

\[
\text{AgCl(s)} \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq)
\]

e.g. 1: Cl⁻ is added to a solution having only Ag⁺ ([Ag⁺] = is $1.0 \times 10^{-5}$ M). What is the maximum [Cl⁻] before precipitate forms?

Ksp = $1.8 \times 10^{-10}$

Ksp = $[\text{Ag}^+] [\text{Cl}^-] = 1.8 \times 10^{-10}$

\[
\therefore [\text{Cl}^-] = \frac{1.8 \times 10^{-10}}{1.0 \times 10^{-5}} = 1.8 \times 10^{-5} \text{M}
\]

e.g. 2: If a 5.0 mL sample of $6.0 \times 10^{-5}$ M Ag⁺ is added to 10.0 mL of $4.2 \times 10^{-6}$ Cl⁻, will a precipitate form?
\[ C_2 = \frac{C_1 V_1}{V_2} \]

\[ [\text{Ag}^+] = \frac{(6.0 \times 10^{-5})(5.0 \text{ mL})}{15.0 \text{ mL}} = 2.0 \times 10^{-5} \text{ M} \]

\[ [\text{Cl}^-] = \frac{(4.2 \times 10^{-6})(5.0 \text{ mL})}{15.0 \text{ mL}} = 2.8 \times 10^{-6} \text{ M} \]

\[ Q = [\text{Ag}^+][\text{Cl}^-] = [2.0 \times 10^{-5} \text{ M}][2.8 \times 10^{-6} \text{ M}] = 5.6 \times 10^{-11} \]

\[ Q = 5.6 \times 10^{-11} < K_{sp} = 1.8 \times 10^{-10} \]

\[ \therefore \text{a precipitate will/will not form} \]

e.g. 3: If 0.1 M KCl is added dropwise to a beaker containing 0.10 M Ag\(^{+}\) and 0.10 M Pb\(^{2+}\), which precipitate would form first? Show all equations and calculations.

We can calculate the minimum \([\text{Cl}^-]\) needed to start the precipitation of AgCl and PbCl\(_2\) by using the Ksp values.

\[
\begin{array}{|c|c|}
\hline
\text{AgCl} & K_{sp} = 1.8 \times 10^{-10} \\
\text{AgCl}(s) & \Leftrightarrow \text{Ag}^+(aq) + \text{Cl}^-(aq) \\
\hline
K_{sp} = [\text{Ag}^+] [\text{Cl}^-] & 1.8 \times 10^{-10} = [0.10 \text{ M}] [\text{Cl}^-] \\
[\text{Cl}^-] = \frac{1.8 \times 10^{-10}}{0.10 \text{ M}} & = 1.8 \times 10^{-9} \text{ M} \\
\hline
\text{PbCl}_2 & K_{sp} = 1.2 \times 10^{-5} \\
\text{PbCl}_2(s) & \Leftrightarrow \text{Pb}^{2+}(aq) + 2 \text{Cl}^-(aq) \\
\hline
K_{sp} = [\text{Pb}^{2+}] [\text{Cl}^-]^2 & 1.2 \times 10^{-5} = [0.10 \text{ M}] [\text{Cl}^-]^2 \\
[\text{Cl}^-] = \sqrt{(1.2 \times 10^{-5})} = 0.011 \text{ M} \\
\hline
\end{array}
\]

As \(1.8 \times 10^{-9} \text{ M} < 0.011 \text{ M}\)

\[ \therefore \text{AgCl will precipitate out first.} \]
Sample Solution - A solution of unknown conc.

Standard Solution - solution of known conc used in a titration.

Equivalence Point – The point where moles of reactant in sample solution have all reacted with moles of the reactant added from standard solution

Titration – the process of adding a standard solution to a sample solution until the equivalence point is reached, in order to determine the [sample solution]

Indicator – substance which does something (i.e. changes colour, form a precipitate) to show equivalence point

The purpose of a titration is to find the conc of an unknown solution.

E.g.: What is the concentration of a Cl\textsuperscript{-} solution which is in the range 0.10 and 5.0 M?

- Use CrO\textsubscript{4}\textsuperscript{2-} as an indicator and force out a precipitate with Ag\textsuperscript{+}.

Note: the flask should be swirled throughout the procedure and the titrant should be added slowly

These 2 equilibria exist in the flask

\[
\text{AgCl(s)} \leftrightarrow \text{Ag}^+ + \text{Cl}^-
\]

Ksp = [Ag\textsuperscript{+}][Cl\textsuperscript{-}] = 1.8 \times 10^{-10}

\[
\text{Ag}_2\text{CrO}_4(s) \leftrightarrow 2\text{Ag}^+ + \text{CrO}_4^{2-}
\]

Ksp = [Ag\textsuperscript{+}]^2[CrO\textsubscript{4}^{2-}] = 1.1 \times 10^{-12}

- Recall Ksp tells us when a precipitate will form.
- Therefore we can use the Ksp values to determine what concentrations of Ag\textsuperscript{+} will form a precipitate first
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Ksp = [Ag⁺][Cl⁻] → [Ag⁺] = Ksp
     [Cl⁻]

[Ag⁺] needed for a precipitate

<table>
<thead>
<tr>
<th></th>
<th>0.10 M Cl⁻</th>
<th>5.0 M Cl⁻</th>
<th>0.010 M CrO₄²⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>[Ag⁺] = Ksp [Cl⁻]</td>
<td>= 1.8 x 10⁻¹⁰ = 1.8 x 10⁻⁹ M Ag⁺</td>
<td>= 1.8 x 10⁻¹⁰ = 3.6 x 10⁻¹¹ M Ag⁺</td>
<td>= 1.1 x 10⁻¹² = 1.0 x 10⁻⁵ M Ag⁺</td>
</tr>
</tbody>
</table>

Therefore 5.0 M Cl⁻ would precipitate out first followed by 0.10 M Cl⁻ and finally 0.010 M CrO₄²⁻.

Note: if CrO₄²⁻ did not precipitate last then it could not be used as an indicator.

Visual Observations (while slowly adding Ag⁺ with swirling)

1. WHITE precipitate forms = AgCl(s) → take biuret reading A
2. RED precipitate forms = Ag₂CrO₄(s) → take biuret reading B
3. when red precipitate has formed almost all of Cl⁻ has reacted

∴ A – B = mL of 0.10 Ag⁺ needed to precipitate out Cl⁻

∴ Calculate moles Ag⁺
And since moles Ag⁺ = moles Cl⁻ and we know the initial volume of Cl⁻ solution

Finally we can calculate [Cl⁻]

E.g. 2: What is the concentration of [Cl⁻] in a 25.0 mL sample of sea water? To get to the equivalence point 26.8 mL of 0.500 M AgNO₃ was added.

Ag⁺ + Cl⁻ → AgCl(s)

Moles Cl⁻ = \( \frac{0.0268 \text{ L AgNO}_3 \times 0.0500 \text{ mol AgNO}_3}{1 \text{ L AgNO}_3} \times \frac{1 \text{ mol AgCl}}{1 \text{ mol AgNO}_3} \times \frac{1 \text{ mol Cl⁻}}{1 \text{ mol AgCl}} \)

= 0.0134 mol Cl⁻

∴ [Cl⁻] = 0.0134 mol Cl⁻ / 0.0250 L Cl⁻ = 0.536 M Cl⁻

Removing Pollution and Hardness from Water by Precipitation Methods

PROBLEM: Waste-water from a mine has [Sr²⁺] = 0.005M, but according to law, [Sr²⁺] cannot exceed 1.0x10⁻⁵M.
**SOLUTION:**
- add CO$_3^{2-}$ to waste water to remove Sr$^{2+}$ as SrCO$_3$.

SrCO$_3(s)$ $\leftrightarrow$ Sr$^{2+}$ + CO$_3^{2-}$

**Question:** What [CO$_3^{2-}$] is needed in the waste-water to lower [Sr$^{2+}$] to $1.0 \times 10^{-5}$M?

K$_{sp}$ of SrCO$_3$ = [Sr$^{2+}$][CO$_3^{2-}$] = $5.6 \times 10^{-10}$ (from Ksp table)

Want [Sr$^{2+}$] = $1.0 \times 10^{-5}$M

Therefore, [CO$_3^{2-}$] = $\frac{5.6 \times 10^{-10}}{1.0 \times 10^{-5}}$ = $5.6 \times 10^{-5}$ M

**Hard Water**

Water is “hard” if too much Ca$^{2+}$ or Mg$^{2+}$ is present.

**Acid rain** dissolves natural limestone (CaCO$_3$), and ions flow into water sources.

CaCO$_3(s)$ + 2H$^+(aq)$ $\leftrightarrow$ Ca$^{2+}(aq)$ + H$_2$O(l) + CO$_2(g)$ + heat

(Same happens to MgCO$_3(s)$)

In hot pipes and hot pots:
- the equilib is forced left
- evaporation leaves the ions and solid behind

∴ pots and pipes get that white solid stuck on them.

**How to soften the water:**
- Add an ion that will form a precipitate with Ca$^{2+}$ and Mg$^{2+}$ (eg. C$_{17}$H$_{22}$COO$^-$)
- Boil the water and force equilib left (precipitate out the ions)
- Filter out the precipitate

**The Common Ion Effect and Altering Solubility**

recall Le Chetaliers Principle:
When a system at equilibrium is disturbed, the equilibrium will shift so as to counteract the disturbance.

E.g. 1-3 Saturated solution: \[ \text{CaCO}_3(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) \]

1. The Common Ion Effect

<table>
<thead>
<tr>
<th>Disturbance</th>
<th>Dissociation</th>
<th>Ion effect</th>
<th>Shift</th>
<th>Amount of solid</th>
</tr>
</thead>
<tbody>
<tr>
<td>add CaCl₂</td>
<td>( \text{CaCl}_2(s) \rightarrow \text{Ca}^{2+}(aq) + 2\text{Cl}^- (aq) )</td>
<td>↑[Ca²⁺]</td>
<td>left</td>
<td>increases</td>
</tr>
<tr>
<td>add K₂CO₃</td>
<td>( \text{K}_2\text{CO}_3(s) \rightarrow 2\text{K}^+(aq) + \text{CO}_3^{2-}(aq) )</td>
<td>↑[CO₃²⁻]</td>
<td>left</td>
<td>increases</td>
</tr>
</tbody>
</table>

2. Decreasing Solubility

<table>
<thead>
<tr>
<th>Added compound</th>
<th>Ions</th>
<th>Effect on Solubility of CaCO₃(s)</th>
<th>Reason for effect</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ca(NO₃)₂</td>
<td>Ca²⁺ NO₃⁻</td>
<td>decrease</td>
<td>equilib shifts left to decrease [Ca²⁺]</td>
</tr>
<tr>
<td>KNO₃</td>
<td>K⁺ NO₃⁻</td>
<td>none</td>
<td>neither ion effects equilib</td>
</tr>
<tr>
<td>CaCO₃</td>
<td>Ca²⁺ CO₃²⁻</td>
<td>none</td>
<td>Same substance</td>
</tr>
</tbody>
</table>

Notice: the equilib shifts when a substance that has a common ion is added. This is the common ion effect.

The added compound should not have low solubility because their dissolved ions are needed for the above reaction.

3. Increasing Solubility

a) By Adding Acid

\[ \text{CaCO}_3(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) \]
• decrease $[Ca^{2+}]$ or $[CO_3^{2-}]$ to **increase** solubility of $CaCO_3(s)$

• how to decrease? Add acid ($H^+$)

$$CaCO_3(s) \leftrightarrow Ca^{2+}(aq) + CO_3^{2-}(aq)$$

$$\downarrow H^+ \text{ (from acid)}$$

$$H_2CO_3(aq) \text{ decomposes}$$

$$H_2O(l) + CO_2(g)$$

• Overall effect: $\downarrow[CO_3^{2-}]$ and equilib shifts **right**
• Solubility of $CaCO_3(s)$ is **increased**.

b) **By Forming Another Precipitate**

$$AgCl(s) \leftrightarrow Ag^+(aq) + Cl^-(aq)$$

Add something that forms a **precipitate** with $Ag^+$ or $Cl^-$

Sulphide forms a precipitate with $Ag^+$:

$$AgCl(s) \leftrightarrow Ag^+(aq) + Cl^-(aq)$$

$$+ S^{2-} \text{ (S}^2\text{- is added)}$$

$$Ag_2S(s)$$

Overall effect:
• $\downarrow[Ag^+]$ and equilib shifts **right**
• Solubility of $AgCl(s)$ is **increased**.
• Could also use $Pb^{2+}$ to precipitate $Cl^-$ ions

**Solubility Test Coming Soon**