Chemistry 12

Unit III - Solubility
I) Ionic & Covalent Solutions

What is a solution? 
- A homogeneous mixture

What are some examples of different types solutions?
- Solid-solid: metal alloys like brass, steel
- Liquid-solid: salt water
- Liquid-liquid: alcohol + water
- Liquid-gas: soft drinks
- Gas-gas: air

Solutions are made up of solute(s) and a solvent. Define each of these terms:
- Solute: the component(s) in a solution in lesser amount
- Solvent: the component is a solution in greatest amount

What is the difference between an ionic solution and a covalent solution?
- Ionic solution: an ionic salt in water (table salt in water)
- Covalent solution: a covalent substance dissolved in water (sugar in water)

What is a salt and what composes a salt?
- A salt is an ionic compound composed of metal cations and non-metal anions in a crystal lattice

How does a solid salt such as NaCl(s) dissolve in water?

http://highered.mcpw-hill.com/sites/0072512644/student_view0/chapter4/animations_center.html (hydration)
http://preparatorychemistry.com/Bishop_Solubility_frames.htm

- Water molecules collide with the crystal lattice and knock Na⁺ cations and Cl⁻ anions into solution
- Water molecules surround the cations and anions
When a salt dissolves in water, it can be described as dissolving, but a more accurate term is dissociation. Why? 
NaCl does not dissolve in water as an NaCl unit. Instead, it dissociates (breaks apart) into Na⁺ cations and Cl⁻ anions.

Sometimes, NaCl\(_{\text{aq}}\) is used to depict a salt in solution (after dissociation). How come this is an inaccurate way of writing it?
When NaCl is in solution, the sodium and chloride are no longer together. A more accurate description is Na\(^+\)\(_{\text{aq}}\) + Cl\(^-\)\(_{\text{aq}}\).

Dissociation Equation for a salt: NaCl\(_{\text{(s)}}\) \(\rightarrow\) Na\(^+\)\(_{\text{aq}}\) + Cl\(^-\)\(_{\text{aq}}\)

*In the solubility unit, always write states when writing equations
Compounds that contain polyatomic ions are obviously ionic in nature (salts) and would dissociate in water to form ions.
Write the dissociation equation for K\(_2\)CO\(_3\)\(_{\text{(s)}}\) dissolving in water:

\[ \text{K}_2\text{CO}_3\text{(s)} \rightarrow 2\text{K}^+\text{aq} + \text{CO}_3^{2-}\text{aq} \]

Ions in solution are called electrolytes. Ions, or electrolytes are what allow an ionic solution to conduct electricity.

What is a molecular or covalent compound composed of?
Non-metals i.e. a sugar molecule C\(_6\)H\(_{12}\)O\(_6\) (C, H, O all non-metals)

What is different about molecular compounds when they dissolve in water compared to ionic compounds?
Molecular compounds dissolve as whole molecules. They don't dissociate into ions.
Diagram of sugar ($C_6H_{12}O_6$) dissolving in water:

Example of a dissolving equation for a molecular compound in water:

$$C_6H_{12}O_6(aq) \rightarrow C_6H_{12}O_6(aq)$$

Write the dissolving equation for $CH_3OH$ (methanol) in water:

$$CH_3OH(l) \rightarrow CH_3OH(aq)$$

Do molecular solutions conduct? Why or why not?

No. They do not create ions in solution


What do we call dissolved molecular compounds?

Non-electrolytes

II) Solubility

What is a saturated solution?

A solution that has the maximum amount of solute dissolved.

How can you recognize a saturated solution visibly?

You will see undissolved solute at the bottom of the container

*If the excess undissolved solute is filtered off, you can still have a saturated solution with no visible undissolved solute on the bottom.

Is a saturated solution static or dynamic?

dynamic → it looks static, but particles are continually dissolving as well as re-crystallizing.

What kind of system is a saturated solution?

Equilibrium

The rate of dissolving equals the rate of crystallization
Therefore, does the amount of undissolved salt at the bottom of a saturated solution change?

No, however much is dissolving, that same amount is simultaneously re-crystallizing

How can a saturated solution equation be written for a saturated solution of KBr?

\[ \text{KBr}_\text{(s)} \overset{\text{dissolving}}{\rightarrow} \text{K}^{\text{aq}} + \text{Br}^{-\text{aq}} \]

Explain the difference between the rate of dissolving and rate of precipitation for an unsaturated solution:

rate of dissolving > rate of precipitation : eventually all solute dissolves

Write an unsaturated solution equation for KBr:

\[ \text{KBr}_\text{(s)} \rightarrow \text{K}^{\text{aq}} + \text{Br}^{-\text{aq}} \]

What’s another name for this equation?

dissociation equation

How do the saturated and unsaturated equations differ?

Saturated \( \Rightarrow \) double arrow
unsaturated \( \Rightarrow \) single arrow

Solubility: the solubility of a substance is a quantitative value.

What is it? the molarity at saturation.

Therefore, solubility can only be measured for what types of solutions?

Saturated solutions

What is ‘concentration’ in chemistry terms?

how much solute is in a certain volume of solution

Concentration is most commonly measured using

Molarity (M) = \( \frac{\text{moles of solute (mol)}}{\text{litres of solution (L)}} \)

\[ \text{Mol} \]

\[ \text{L} \]
Concentration can also be measured using mass of solute/volume (g/L or g/mL) of solution, as you may see in some examples.

Every salt dissociates to some extent in water, but every salt has a different solubility in water. What does this mean?
- Some ionic salts will make a solution in water. However, some become saturated very quickly (at a low molarity), whereas other salts can dissolve quite a lot in water before making a saturated solution (at a high molarity).
Some salts can dissociate to a large extent before becoming saturated, while other salts dissociate very little before becoming saturated.

Solubility is also temperature dependent. A general rule for solid-liquid solutions is that solubility increases with increasing temperature.

**Assignment 1:** Read pages 73-76 (check out the comic on p.76) in Hebden and do Questions #1-5

**III) Calculating Solubility and Ion Concentrations**

Use stoichiometry to calculate the concentration of ions given the concentration of solute.

Calculate [Na\(^+\)] and [SO\(_4^{2-}\)] in a 0.20M solution of Na\(_2\)SO\(_4\)(aq)

Dissoc eqn: \( \text{Na}_2\text{SO}_4(s) \rightarrow 2\text{Na}^{+}(aq) + \text{SO}_4^{2-}(aq) \)

\[
\begin{align*}
0.20M & \quad 0.40M \\
\text{X2} & \quad \text{X1}
\end{align*}
\]

Calculate [Fe\(^{3+}\)] and [SO\(_4^{2-}\)] in a 0.50M solution of iron III sulphate: \( \text{Fe}_2(\text{SO}_4)_3(s) \rightarrow 2\text{Fe}^{3+}(aq) + 3\text{SO}_4^{2-}(aq) \)

\[
\begin{align*}
0.50M & \quad 1.0M \\
\text{X1} & \quad 1.5M
\end{align*}
\]

A saturated solution of PbCl\(_2\) is found to contain 9.90g of PbCl\(_2\) per litre of solution. Find [Pb\(^{2+}\)]. *change grams to moles, then to M. Then write dissociation eqn and use stoich:

\[
\frac{9.90g \text{PbCl}_2}{278.2g} = 0.035586 \text{ mol PbCl}_2
\]

\[
\begin{align*}
\text{PbCl}_2(s) & \rightarrow \text{Pb}^{2+}(aq) + 2\text{Cl}^{-}(aq) \\
0.035586M & \quad 0.0356M
\end{align*}
\]

\[
\frac{0.035586\text{mol}}{1.00L} = 0.035586\text{M}
\]

\[
\begin{align*}
\text{[Pb}^{2+}] &= 0.0356\text{M}
\end{align*}
\]
Write an equation showing the equilibrium in a saturated solution of silver carbonate. The solubility of \( \text{Ag}_2\text{CO}_3 \) is \( 1.29 \times 10^{-4} \) M. Find \([\text{Ag}^+]\) in grams per 100.0mL:

\[
\begin{align*}
\text{Ag}_2\text{CO}_3(s) & \rightleftharpoons 2\text{Ag}^+(aq) + \text{CO}_3^{2-}(aq) \\
1.29 \times 10^{-4} \text{ M} & \quad \times 2 \\
2.58 \times 10^{-4} \text{ mol} & \quad \div 1 \text{ mol} \\
\frac{2.58 \times 10^{-4} \text{ mol} \times \text{Ag}^+}{2} & = 0.0278 \text{ \% Ag}^+ \\
0.0278 \frac{\text{mol \ Ag}^+}{\text{L}} \div 10 & = 0.00278 \frac{\text{g \ Ag}^+}{100.0 \text{ mL}}
\end{align*}
\]

A 558mL solution of \( \text{Al}_2(\text{SO}_4)_3 \) has a concentration of 0.0342M. You then add 325mL of water to it. What is the new concentration?

\[
\text{dilution} : \quad M_iV_i = M_fV_f \\
M_i = 0.0342 \text{ M} \\
V_i = 0.558 \text{ L} \\
V_f = 558 + 325 = 883 \text{ mL} = 0.883 \text{ L} \\
(0.0342 \text{ M})(0.883 \text{ L}) = M_f(0.883 \text{ L}) \\
M_f = \frac{(0.0342 \text{ M})(0.883 \text{ L})}{0.883 \text{ L}} = 0.0216 \text{ M}
\]

\[
\left[ \text{Al}_2(\text{SO}_4)_3 \right]_f = 0.0216 \text{ M}
\]

250mL of 0.30M \( \text{K}_2\text{SO}_4 \) and 250mL of 0.80M \( \text{MgCl}_2 \) are mixed and no precipitate (solid) forms. Calculate the concentration of each ion in the final solution.

\[
\text{dilution} : \\
\left[ \text{K}^+ \right]_f = \frac{(0.30 \text{ M})(0.25 \text{ L})}{0.50 \text{ L}} = 0.15 \text{ M} \\
\left[ \text{K}^+ \right]_f \rightarrow 2\text{K}^+(\text{aq}) + 3\text{SO}_4^{2-}(\text{aq}) \\
0.15 \text{ M} \quad 0.30 \text{ M} \quad 0.15 \text{ M} \\
\left[ \text{Mg}^{2+} \right]_f \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) \\
0.40 \text{ M} \quad 0.40 \text{ M} \quad 0.80 \text{ M} \\
\left[ \text{K}^+ \right] = 0.30 \text{ M} \quad \left[ \text{Mg}^{2+} \right] = 0.40 \text{ M} \\
\left[ \text{SO}_4^{2-} \right] = 0.15 \text{ M} \quad \left[ \text{Cl}^- \right] = 0.80 \text{ M}
\]

Assignment 2: Hebden p.77 # 8, 9 & p.78 #13 & p.81 #18a-d, 20abfg
IV) Using the Solubility Table

The solubility table in the data booklet is used to predict whether a salt is soluble in water or low solubility in water. Soluble salts can dissociate and produce a solution of at least 0.1M before becoming saturated. Low Solubility salts dissociate very little as they become saturated before reaching 0.1M.

Identify three soluble salts, and three low solubility salts using the table:

**SOLUBLE**
- NaCl
- K₂SO₄
- Sr(OH)₂

**LOW SOLUBILITY**
- AgCl
- CaSO₄
- BaCO₃

If two ions that are low solubility with one another are put into the same solution and the resulting concentration exceeds the solubility for that salt (molarity of saturation for that salt), then some of those cations & anions will form together and precipitate (form a solid) out of solution until the ions have reached their saturation molarity.

**Precipitate**: a solid that forms from a solution due to the mixing of two ions that originated in different solutions, but when mixed together, became ‘oversaturated’.

\[ K₂CrO₄(aq) + Ba(NO₃)₂(aq) \rightarrow \text{Products} \]

\[ \text{BaCrO}_4 \]

http://www.absorbblearning.com/media/item.action?quick=uu
http://preparatorychemistry.com/Bishop_Solubility_frames.htm
Is it possible that a soluble salt could form a precipitate?
Yes. Any salt will precipitate if its solubility (molarity of saturation) is exceeded.

E.g. solubility of NaCl is 6.15M @ 25°C which is about 360g in 1L. What if you put 500g into 1L?
So, saturated solutions of a salt (for example, PbCl₂) can be created in two different ways. What are they?

1. Add solid PbCl₂ to water until no more PbCl₂ dissolves and just sinks to the bottom.
2. You can mix two solutions - one with Pb²⁺ ions, such as Pb(NO₃)₂, and another with Cl⁻ ions, such as NaCl(aq). If the Pb²⁺ and Cl⁻ ions are oversaturated when mixed together, a precipitate of PbCl₂ will form, leaving a saturated solution of PbCl₂ (as well as having Na⁺ and NO₃⁻ ions present).

Using your table, which ions are soluble with all other ions?

**Alkali ions**: Li⁺, Na⁺, K⁺, Cs⁺, Rb⁺, Fr⁺ soluble with all anions
H⁺ is soluble with all anions
NH₄⁺ " " " " " "
NO₃⁻ is soluble " " all cations.

Use your table to predict whether the following salts are soluble, or low solubility (meaning they are likely to form a ppt):

1. Sodium hydroxide S
2. Ammonium acetate S
3. Calcium sulphate LS
4. Lead II chloride LS
5. Potassium chloride S
6. Calcium bromide S
7. Potassium carbonate S
8. Aluminum sulphate S
9. Lead sulphide LS
10. Copper II sulphate S
11. Iron II sulphide LS
12. Barium hydroxide LS
13. Silver bromide LS
14. Magnesium carbonate LS
15. Copper II chloride S
16. Copper I chloride LS

http://hiphered.mcgraw-hill.com/sites/0072512644/student_view0/chapter4/animations_center.html
i) Will a precipitate form if equal volumes of the following 0.2M solutions are mixed? *do the double replacement and use your table

\[ \text{Pb(NO}_3\text{)}_2(aq) + 2\text{KI(aq)} \Rightarrow \text{PbI}_2(s) + 2\text{KNO}_3(aq) \]

- low solubility on table, so will form a ppt
- soluble on table - stays aqueous

Yes, PbI\(_2\).

ii) Which 2.0M soluble solutions could be mixed in order to produce a precipitate of Mg(OH)\(_2\)?

- Need to mix two solutions: one with Mg\(^{2+}\) and one with OH\(^-\)

\[ \text{Mg(NO}_3\text{)}_2(aq) \text{ and NaOH}(aq) \]

- soluble with all cations
- soluble with all anions.

Assignment 3: Hebden p.83 #21, 22 & p. 84 #24

V) Formula, Complete, and Net Ionic Equations

**Formula Equation** (double replacement equation):

\[ \text{CaCl}_2(aq) + 2\text{AgNO}_3(aq) \Rightarrow \text{Ca(NO}_3\text{)}_2(aq) + 2\text{AgCl(s)} \]

- low solubility on table, so if ppt.

**Complete Ionic Equation:**

In reality, aqueous salts are actually dissociated into ions. Precipitates (solids) are not dissociated, so are written in compound form.

Build and balance the complete ionic equation using the formula equation.

\[ \text{Ca}^{2+}(aq) + 2\text{Cl}^-(aq) + 2\text{Ag}^{+}(aq) + 2\text{NO}_3^-(aq) \Rightarrow \text{Ca}^{2+}(aq) + 2\text{NO}_3^-(aq) + 2\text{AgCl(s)} \]

**Net Ionic Equation:**

Only ions which become a part of the precipitate are written as reactants. Build the net ionic eqn using the complete ionic eqn as assistance.

\[ 2\text{Ag}^+(aq) + 2\text{Cl}^-(aq) \rightarrow 2\text{AgCl(s)} \]

\[ \text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl(s)} \]
Assignment 4: For each of the following, write formula, complete, and net ionic equations:
1. Aqueous magnesium chloride is mixed with aqueous potassium hydroxide.
2. Aqueous lithium carbonate is mixed with aqueous iron (II) sulphate.
3. Aqueous copper I sulphate is mixed with aqueous strontium bromide.
4. Hebden p. 87 #25e,i

VI) Separating Mixtures of Ions by Precipitation
Describe a method to separate Ba\(^{2+}\) and Pb\(^{2+}\) ions which are in solution together: \(Cl^-\) will precipitate Pb\(^{2+}\), but not Ba\(^{2+}\)

1. Add \(Cl^-\) as NaCl\(_{aq}\) to ppt Pb\(^{2+}\) as PbCl\(_2\)(s)
2. Filter off the solid PbCl\(_2\).
3. Add SO\(_4^{2-}\) as Na\(_2\)SO\(_4\)(aq) to ppt Ba\(^{2+}\) as BaSO\(_4\)(s)

Assignment 5: Ion Separation Exercises
1. Describe a method to separate Cl\(^-\) and OH\(^-\) that are in solution together. Whatever cation you choose in each step, add it as component of a salt with nitrate (since nitrate is soluble with any cation).
2. A solution is known to possibly contain one or more of the following ions: Mg\(^{2+}\), Ca\(^{2+}\), Pb\(^{2+}\)
   What steps can you carry out to find the composition of the solution?
3. You have a solution known to contain any or all of Cu\(^+\), Ca\(^{2+}\), Fe\(^{3+}\), and Sr\(^{2+}\). You have the following 'test' solutions available: 1M Na\(_2\)CO\(_3\), 1M Na\(_2\)SO\(_4\), 1M NaOH, 1M NaCl, 1M Na\(_2\)S. In what order would you add each of the above test solutions to test for the presence of each cation?
4. Hebden p. 90 #28, 31, 32

VII) Quantitative Solubility
A saturated solution is in equilibrium because the rate of \underline{dissolving} equals the rate of \underline{precipitation}.
Assignment 4

1. \[ \text{F: } \text{MgCl}_2(\text{aq}) + 2\text{KOH}(\text{aq}) \rightarrow \text{Mg(OH)}_2(\text{s}) + 2\text{KCl}(\text{aq}) \]
   \[ \text{C: } \text{Mg}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) + 2\text{K}^+(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Mg(OH)}_2(\text{s}) + 2\text{K}^+(\text{aq}) + 2\text{Cl}^-(\text{aq}) \]
   \[ \text{N: } \text{Mg}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Mg(OH)}_2(\text{s}) \]

2. \[ \text{F: } \text{Li}_2\text{CO}_3(\text{aq}) + \text{FeSO}_4(\text{aq}) \rightarrow \text{Li}_2\text{SO}_4(\text{aq}) + \text{FeCO}_3(\text{s}) \]
   \[ \text{C: } 2\text{Li}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow 2\text{Li}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{FeCO}_3(\text{s}) \]
   \[ \text{N: } \text{Fe}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{FeCO}_3(\text{s}) \]

3. \[ \text{F: } \text{Cu}_2\text{SO}_4(\text{aq}) + \text{SrBr}_2(\text{aq}) \rightarrow 2\text{CuBr}(\text{s}) + \text{SrSO}_4(\text{s}) \]
   \[ \text{C: } 2\text{Cu}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Sr}^{2+}(\text{aq}) + 2\text{Br}^-(\text{aq}) \rightarrow 2\text{CuBr}(\text{s}) + \text{SrSO}_4(\text{s}) \]
   \[ \text{N: } \text{same as complete ionic eqn.} \]
BaSO₄(s) ⇌ Ba²⁺(aq) + SO₄²⁻(aq).

Because equilibrium exists, there can be an equilibrium constant. The equilibrium constant for a saturated solution is called a $K_{sp}$.
Write the $K_{sp}$ equation for the saturated solution above:

$$K_{sp} = [Ba^{2+}][SO_{4}^{2-}]$$

The 'sp' stands for solubility product because...

product is the result of a multiplication. A $K_{sp}$ is always simply a multiplication (never a denominator) because the substance on the left side (reactant side) of the reaction equation is always a solid.

Write the balanced saturated solution equations, and then $K_{sp}$ equations for saturated solutions of Ca(OH)₂ and Ag₃PO₄.

*When writing saturated solution equations that will be used to make a $K_{sp}$ equation, always put the solid on the left, and the ions on the right.

$$\text{Ca(OH)}_2(s) \rightleftharpoons Ca^{2+}(aq) + 2OH^-(aq)$$
$$K_{sp} = [Ca^{2+}][OH^-]^2$$

$$\text{Ag}_3\text{PO}_4(s) \rightleftharpoons 3Ag^+(aq) + PO_{4}^{3-}(aq)$$
$$K_{sp} = [Ag^+]^3[PO_{4}^{3-}]$$

Would soluble salts have large or small $K_{sp}$ values? Why?
Large. Soluble salts have high ion concentrations in solution. This would result in large $K_{sp}$ values.

What about low solubility salts?
Small. Low solub. salts have very small ion concs. in solution.

Notice that the $K_{sp}$ table in the data booklet (p.5) only lists $K_{sp}$ constants for low solubility salts (salts that easily form precipitates).
Notice they are all small values (much less than 1), as they should be.
VIII) Type A Problems: Calculating $K_{sp}$

$K_{sp}$ constants can be calculated for soluble or low solubility salts, as long as there is information available about the respective saturated solution. Examples:

1. If the solubility of Ag$_2$S in water is $1.3 \times 10^{-17}$M, calculate the $K_{sp}$.

*What is solubility again? molarity at saturation.*

*Write an equilibrium equation for Ag$_2$S and use stoichiometry to find the concentrations of the ions in the saturated solution.*

\[ \text{Ag}_2\text{S}(s) \rightleftharpoons 2\text{Ag}^+(aq) + \text{S}^{2-}(aq) \]

\[ 1.3 \times 10^{-17} \text{M} \quad 2.6 \times 10^{-17} \text{M} \quad 1.3 \times 10^{-17} \text{M} \]

\[ K_{sp} = [\text{Ag}^+]^2[\text{S}^{2-}] \]

\[ K_{sp} = (2.6 \times 10^{-17})^2(1.3 \times 10^{-17}) = 8.8 \times 10^{-51} \]

$\text{Ag}_2\text{S}$ is a low solubility substance as $K_{sp}$ very small.

2. Calculate the $K_{sp}$ of MgS if a maximum of 300.0g can be dissolved in 2.00L of water:

\[ \frac{300.0\text{g}}{56.4\text{g MgS}} = 5.319 \text{ mol MgS} \]

\[ [\text{MgS}] = \frac{5.319 \text{ mol}}{2.00\text{L}} = 2.66 \text{ M} \]

\[ \text{MgS}(s) \rightleftharpoons \text{Mg}^{2+}(aq) + \text{S}^{2-}(aq) \]

\[ 2.66 \text{ M} \quad 2.66 \text{ M} \quad 2.66 \text{ M} \]

\[ K_{sp} = [\text{Mg}^{2+}][\text{S}^{2-}] = (2.66)(2.66) = 7.07 \leftarrow \text{large } K_{sp} \text{ so MgS soluble.} \]
3. A 25.00mL sample of a ZnF₂ saturated solution (with no excess solid on
the bottom) was evaporated to dryness. The mass of the residue was
0.508g. Calculate the solubility product constant of ZnF₂.
The 0.508g that remained after evaporation was the amount of ZnF₂ that
was originally dissolved to make a saturated solution.

\[
\frac{0.508g\text{ ZnF}_2}{1\text{ mol ZnF}_2} \times \frac{103.45g\text{ ZnF}_2}{1} = 4.913 \times 10^{-3} \text{ mol} 
\]

\[
[ZnF_2] = \frac{4.913 \times 10^{-3} \text{ mol}}{0.02500 \text{ L}} = 0.1965 \text{ M} 
\]

\[
\text{ZnF}_2(s) \rightleftharpoons \text{Zn}^{2+}(aq) + 2\text{F}^{-}(aq)
\]

\[
0.1965 \text{ M} \quad 0.1965 \text{ M} \quad 0.393 \text{ M}
\]

\[
K_{sp} = [\text{Zn}^{2+}][\text{F}^{-}]^2
\]

\[
= (0.1965)(0.393)^2 
\]

\[
= 3.04 \times 10^{-2}
\]

**Assignment 6: Type A Exercises**

1. At 25°C, only 0.00245g of BaSO₄ can be dissolved in 1.0L of H₂O.
   Calculate the Kₚ for BaSO₄. \(1 \times 10^{-10}\)

2. At 25°C, the solubility of Ag₃PO₄ is \(1.8 \times 10^{-5}\)M. Calculate the Kₚ for
   Ag₃PO₄. \(2.8 \times 10^{-16}\)

3. An experiment showed that a maximum of 1.49g of AgBrO₃ can dissolve
   in 1.00L of water at 25°C. What is the Kₚ for AgBrO₃ at this
temperature? \(3.99 \times 10^{-5}\)

4. A saturated solution of CaF₂ contains \(4.15 \times 10^{-4}\)mol of CaF₂ in 2.0L of
   solution. What is the Kₚ for CaF₂? \(3.6 \times 10^{-11}\)

5. A solution in equilibrium with solid Ag₂S on the bottom of the beaker
   was found to contain \(1.6 \times 10^{-16}\)M S²⁻ and \(2.5 \times 10^{-18}\)M Ag⁺. Calculate the
   solubility product constant of Ag₂S. \(1.0 \times 10^{-51}\)
IX) Type B Problems: Calculating Solubility

In Type B problems, you must calculate the solubility of a salt using the $K_{sp}$ constant from p.5 of the data booklet. Remember that solubility means the molarity of the salt at equilibrium (when it is saturated).

Examples:

1. Calculate the solubility of $\text{CaCO}_3$ in water at $25^\circ\text{C}$.

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

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

   $K_{sp}$ from table = $5.0 \times 10^{-9}$

   Let $x = [\text{Ca}^{2+}] = [\text{CO}_3^{2-}]$

   \[
   5.0 \times 10^{-9} = (x)(x)
   \]

   \[
   x^2 = 5.0 \times 10^{-9}
   \]

   \[
   x = \sqrt{5.0 \times 10^{-9}}
   \]

   \[
   x = 7.07 \times 10^{-5}
   \]

   So $[\text{Ca}^{2+}] = [\text{CO}_3^{2-}] = 7.07 \times 10^{-5}$

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

   \[
   7.07 \times 10^{-5} \quad 7.07 \times 10^{-5} \quad 7.07 \times 10^{-5}
   \]

   So solubility of $\text{CaCO}_3 = 7.1 \times 10^{-3} \text{M}$

2. Calculate the solubility of $\text{PbI}_2$ in g/L.

   \[
   \text{PbI}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2\text{I}^-(aq)
   \]

   \[
   K_{sp} = [\text{Pb}^{2+}][\text{I}^-]^2
   \]

   $K_{sp}$ from table = $8.5 \times 10^{-9}$

   Let $x = [\text{Pb}^{2+}]$

   then $2x = [\text{I}^-]$

   \[
   8.5 \times 10^{-9} = (x)(2x)
   \]

   \[
   8.5 \times 10^{-9} = 4x^3
   \]

   \[
   x = \sqrt[3]{\frac{8.5 \times 10^{-9}}{4}}
   \]

   \[
   x = 1.286 \times 10^{-3}
   \]

   So solubility of $\text{PbI}_2 = 1.286 \times 10^{-3} \text{M}$

   Change moles to grams:

   \[
   1.286 \times 10^{-3} \text{mol} \quad 461.09 \text{g} \text{PbI}_2
   \]

   \[
   = 0.59 \%
   \]
3. How many grams of PbBr₂ can be dissolved in 250.0mL of water at 25°C? 

\[ \text{PbBr}_2 (s) \rightleftharpoons \text{Pb}^{2+} (aq) + 2\text{Br}^- (aq) \]

Find solubility of PbBr₂ first:

\[ K_{sp} = \left[ \text{Pb}^{2+} \right] \left[ \text{Br}^- \right]^2 \]

Kₚₛ from table = 6.6 x 10⁻⁶

Let \( x = \left[ \text{Pb}^{2+} \right] \)

then \( 2x = \left[ \text{Br}^- \right] \)

\[
\begin{align*}
6.6 \times 10^{-6} &= 4x^3 \\
x &= \sqrt[3]{\frac{6.6 \times 10^{-6}}{4}} \\
x &= 0.01817 \\
\left[ \text{Pb}^{2+} \right] &= 0.01817 \text{M}
\end{align*}
\]

Solubility = 0.01817 M of PbBr₂

\[ \text{now must change } \frac{\text{mol}}{\text{L}} \text{ to } \frac{\text{g}}{250\text{mL}} \]

\[ \frac{0.01817 \text{ moles PbBr}_2}{1 \text{ mol PbBr}_2} \times \frac{367.0 \text{ g PbBr}_2}{1 \text{ mol PbBr}_2} = \frac{4.3368 \text{ g}}{L} \]

\[ \text{change L to 250 mL} \Rightarrow \text{divide by 4} \]

\[ 4.3368 \div 4 = 1.1 \text{ g PbBr}_2 \]

Assignment 7: Type B Exercises

1. The Kₚₛ at a certain temperature for Ni(OH)₂ is 1.6 x 10⁻¹⁶. Calculate the solubility of Ni(OH)₂. \( 3.4 \times 10^{-4} \text{M} \)

2. Find the solubility of CaSO₄ in g/L. \( 1.1 \% \)

3. Which saturated solutions at 25°C will have a greater [Ag⁺], AgCl or Ag₂CO₃? \( \frac{\text{AgCl} \; \sqrt{[\text{Ag}^+]}}{\text{Ag}_2\text{CO}_3} = 2.6 \times 10^{-4} \text{M} \) vs. \( \frac{\text{AgCl} \; \sqrt{[\text{Ag}^+]}}{\text{Ag}_2\text{CO}_3} = 1.3 \times 10^{-5} \text{M} \)

4. Calculate the mass of MgCO₃ which could be dissolved in 3.0L of water at 25°C. \( 0.66 \text{ g} \)

5. What mass of PbSO₄ will dissolve in 5.0L of water at 25°C? \( 2.0 \times 10^{-3} \text{ g} \)
X) Type C Problems: Trial $K_{sp}$
When two soluble solutions are mixed, a cation from one solution is introduced to an anion from the other solution, and *vis a versa*. One or both of the new ion combinations (cation and anion) could have low solubility. If this is the case, and there are too many of the low solubility ions present (more than what is necessary for a saturated solution), a precipitate will result. If there are not enough of the low solubility ions present to reach the saturation concentration, a precipitate will not form and the ions will stay dissolved in solution.

A Trial $K_{sp}$ (same concept as a trial $K_{eq}$) can be calculated for these mixtures to deduce whether a precipitate will form or not when the two solutions are mixed.

If the trial $K_{sp}$ is larger than the actual $K_{sp}$, the concentrations of low solubility ions mixed together is greater than saturation, so a precipitate will result. If the trial $K_{sp}$ is smaller than the actual $K_{sp}$, the concentrations of low solubility ions mixed together has not yet reached saturation, therefore a precipitate will not form.

Examples:

1. Will a precipitate form if 40.0mL of 8.0 $\times$ 10$^{-3}$M Mg(NO$_3$)$_2$ is mixed with 60.0mL of 1.0 $\times$ 10$^{-2}$M K$_2$CO$_3$?

   $\text{Mg(NO}_3\text{)}_2(\text{aq}) + \text{K}_2\text{CO}_3(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{MgCO}_3(\text{s})$

   Since two solutions are being mixed, dilution occurring:
   
   $$[\text{Mg(NO}_3\text{)}_2]_f = \frac{(8.0 \times 10^{-3}\text{M})(0.0400\text{L})}{0.100\text{L}}$$
   
   $$= 3.2 \times 10^{-3}\text{M}$$

   $\text{Mg(NO}_3\text{)}_2(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{NO}_3^- (\text{aq})$

   $\text{K}_2\text{CO}_3(\text{aq}) \rightarrow 2\text{K}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$

   $$\text{K}_2\text{CO}_3(\text{aq}) \rightarrow 2\text{K}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$$

   $$6.0 \times 10^{-3}\text{M}$$

   Trial $K_{sp}$ = $3.2 \times 10^{-3}(6.0 \times 10^{-3})$

   $$= 1.9 \times 10^{-5}$$

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

   $K_{sp}$ from table = $6.8 \times 10^{-6}$

   : a ppt of MgCO$_3$ will form.

   Trial $K_{sp}$ > $K_{sp}$
2. Will a precipitate form if 50.0mL of 0.00100M CaCl$_2$ is added to 50.0mL of 0.0100M Na$_2$SO$_4$? 

\[ \text{CaCl}_2(aq) + \text{Na}_2\text{SO}_4(aq) \rightarrow 2\text{NaCl}(aq) + \text{CaSO}_4(s) \]

Possible precipitate

\[ [\text{CaCl}_2]_f = \frac{(0.00100)(0.0500L)}{0.1000L} = 5.00 \times 10^{-4}M \]

\[ \text{CaCl}_2(aq) \rightarrow \frac{[\text{Ca}^{2+}(aq)] + 2[\text{Cl}^{-}(aq)]}{5.00 \times 10^{-4}M} \]

\[ [\text{Na}_2\text{SO}_4]_f = \frac{(0.0100)(0.0500L)}{0.1000L} = 5.00 \times 10^{-3}M \]

\[ \text{Na}_2\text{SO}_4(aq) \rightarrow 2[\text{Na}^+(aq)] + \frac{[\text{SO}_4^{2-}(aq)]}{5.00 \times 10^{-3}M} \]

Possible ppt is CaSO$_4$:

\[ \text{CaSO}_4(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{SO}_4^{2-}(aq) \]

K$_{sp}$ = [Ca$^{2+}$][SO$_4^{2-}$]

K$_{sp}$ from table = 7.1 x 10$^{-5}$

Trial K$_{sp}$ or $Q$ = (5.00 x 10$^{-4}$)(5.00 x 10$^{-3}$)

\[ Q = 2.5 \times 10^{-6} \]

Since $Q < K_{sp}$

A ppt of CaSO$_4$ will not form.

Assignment 8: Type C Exercises

1. Will a precipitate form when 1.0L of 3.0 x 10$^{-10}$M Zn(NO$_3$)$_2$ is added to 1.0L of 2.0 x 10$^{-11}$M Na$_2$S? $Q = 1.5 \times 10^{-21}$ a ppt will form

2. Will a precipitate form when 2.0L of 6.7 x 10$^{-3}$M SrS is added to 1.0L of 4.3 x 10$^{-4}$M K$_2$SO$_4$? $Q = 6.4 \times 10^{-7}$ a ppt will form

3. Will a precipitate form when 1.0L of 5.0 x 10$^{-4}$M MgCl$_2$ is added to 1.0L of 2.0 x 10$^{-5}$M Na$_2$CO$_3$? $Q = 2.5 \times 10^{-6}$ a ppt will not form

4. 2.5 x 10$^{-7}$M of Fe(NO$_3$)$_2$ is mixed with 1.0 x 10$^{-5}$M NaOH. Will a precipitate be observed? $Q = 3.1 \times 10^{-10}$ a ppt will not be observed if these two solutions are mixed
XI) Type D Problems: Determining the Concentration of a Specific Ion in a Saturated Solution

Recall that there are two ways to make a saturated solution. What are they (use Pb(OH)\(_2\) as an example)?

1. add solid Pb(OH)\(_2\) to water until no more will dissolve.

2. Mix two soluble solutions together: one that contains \(\text{Pb}^{2+}\) such as \(\text{Pb(NO}_3\text{)}_2\) and one that contains \(\text{OH}^-\) such as \(\text{NaOH}\). When mixed, if \([\text{Pb}^{2+}]\) and \([\text{OH}^-]\) exceed the saturation molarity (their solubility), some will precipitate out as \(\text{Pb(OH)}_2\) and you'll be left with a saturated \(\text{Pb(OH)}_2\) solution (with \(\text{NO}_3\) and \(\text{Na}\) ions also present).

It is important to recognize that ion stoichiometry for a saturated solution is only a reality when the saturated solution is made from dissolving the salt in water (dissolving Pb(OH)\(_2\) solid in water until you have a saturated solution – method #1 above). What would be the proportion of \([\text{OH}^-]\) compared to \([\text{Pb}^{2+}]\) in this case? \[\text{Pb(OH)}_2(s) \rightleftharpoons \text{Pb}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})\]

If a saturated solution of a salt (such as Pb(OH)\(_2\)) is made by combining two different solutions, such as \(\text{Pb(NO}_3\text{)}_2\) and \(\text{NaOH}\), the \([\text{Pb}^{2+}]\) relative to the \([\text{OH}^-]\) could be anything, depending on the proportions of the solutions mixed. So, you can't rely on 2 OH\(^-\) to 1 Pb\(^{2+}\) stoichiometry because the ions came from different sources. However, the \(K_{sp}\) equation can still be relied upon in these situations to do calculations for saturated solutions, as you'll see in Type D problems.

Examples:

1. What is the maximum \([\text{Sr}^{2+}]\) that can be dissolved in a 0.020M solution of \(\text{K}_2\text{SO}_4\) without a precipitate of \(\text{SrSO}_4\) forming?

\[\text{K}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{K}^{+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})\]

\[0.020\text{M}\]

\[\text{SrSO}_4(s) \rightleftharpoons \text{Sr}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})\]

\[0.020\text{M}\]

\[K_{sp} = [\text{Sr}^{2+}][\text{SO}_4^{2-}]\]

\[K_{sp}\text{ from table} = 3.4 \times 10^{-7}\]

\[3.4 \times 10^{-7} = [\text{Sr}^{2+}](0.020\text{M})\]

\[[\text{Sr}^{2+}] = 1.7 \times 10^{-5}\text{M}\]
2. If Cu\(^{2+}\) is slowly added to a solution of 0.020M KIO\(_3\), at what [Cu\(^{2+}\)] does a precipitate of Cu(IO\(_3\))\(_2\) just start to form?

\[
\begin{align*}
K_{sp} & = [Cu^{2+}][IO_3^-]^2 \\
\text{Ksp from table} & = 6.9 \times 10^{-8} \\
6.9 \times 10^{-8} & = [Cu^{2+}](0.020M)^2 \\
[Cu^{2+}] & = 1.7 \times 10^{-4} M
\end{align*}
\]

3. Fe(NO\(_3\))\(_2\) is slowly added to a 2.0L solution of 0.010M Na\(_2\)S. What mass of Fe(NO\(_3\))\(_2\) would be required to just start precipitation?

What is ppt?  \( \text{Fe(NO}_3\text{)}_2 + \text{Na}_2\text{S} \rightarrow 2\text{NaNO}_3(\text{aq}) + \text{FeS(s)} \)

\[
\begin{align*}
\text{Na}_2\text{S(\text{aq})} & \rightarrow 2\text{Na}(\text{aq}) + \text{S}^{2-}(\text{aq}) \\
\text{Ksp} & = [\text{Fe}^{2+}][\text{S}^{2-}] \\
\text{Ksp from table} & = 6.0 \times 10^{-19} \\
6.0 \times 10^{-19} & = [\text{Fe}^{2+}](0.010M) \\
[\text{Fe}^{2+}] & = 6.0 \times 10^{-17} M
\end{align*}
\]

\[
\begin{align*}
\text{Fe(NO}_3\text{)}_2(\text{s}) & \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{NO}_3^- \\
\text{mass Fe(NO}_3\text{)}_2 & = \frac{1.2 \times 10^{-4} M \times 179.8 g}{1 \text{ mol \ Fe(NO}_3\text{)}_2} \\
& = 2.2 \times 10^{-4} g
\end{align*}
\]

4. Removing pollution by precipitation:

Waste water from a mining operation was found to have a [Cd\(^{2+}\)] exceeding environmental standards. Before discharging the water into an adjacent river, the [Cd\(^{2+}\)] in the tank had to be lowered to 1.0 \times 10^{-5} M.

What [OH\(^-\)] would be required to bring the [Cd\(^{2+}\)] to an acceptable level?  \( K_{sp} = 2.2 \times 10^{-14} \) for Cd(OH)\(_2\)

\[
\begin{align*}
\text{Cd(OH)}_2(\text{s}) & \rightarrow \text{Cd}^{2+}(\text{aq}) + 2\text{OH}^- (\text{aq}) \\
\text{Ksp} & = [\text{Cd}^{2+}][\text{OH}^-]^2 \\
2.2 \times 10^{-14} & = (1.0 \times 10^{-5}) [\text{OH}^-]^2 \\
[\text{OH}^-] & = \sqrt{\frac{2.2 \times 10^{-14}}{1.0 \times 10^{-5}}} \\
& = 4.7 \times 10^{-5} M
\end{align*}
\]
Assignment 9: Type D Exercises
1. What \([S^{2-}]\) must be present to just start precipitation from a 0.20M solution of \(\text{CuCl}_2\)? 3.0 \(\times\) 10\(^{-3}\) M
2. What \([\text{Cu}^{2+}]\) would be required to just start precipitation of \(\text{Cu(IO}_3)_2\) from a 0.20M solution of \(\text{KIO}_3\)? 1.7 \(\times\) 10\(^{-6}\) M
3. \(\text{AgNO}_3\) is added to a 0.10M solution of \(\text{NaCl}\). What \([\text{Ag}^+]\) must be present in order to observe a precipitate? 1.8 \(\times\) 10\(^{-3}\) M
4. What is the maximum \([\text{SO}_4^{2-}]\) that can be present in 0.010M \(\text{CaBr}_2\) solution without a precipitate forming? 7.1 \(\times\) 10\(^{-3}\) M
5. What \([\text{Ba}^{2+}]\) must be present to just start a precipitation in a 5.0 \(\times\) 10\(^{-4}\)M solution of \(\text{K}_2\text{CrO}_4\)? 2.4 \(\times\) 10\(^{-7}\) M

XII) Hard Water
Hard water is caused by a high concentration of \(\text{Ca}^{2+}\) and/or \(\text{Mg}^{2+}\) ions in water. It is a problem in areas of the world rich in limestone (\(\text{CaCO}_3\)). Water from rivers, lakes, and streams pass over limestone rock, and \(\text{Ca}^{2+}\) ends up in water by the two processes below.
1. Permanent Hard Water is formed by Acidic Water (water containing \(\text{H}^+\) due to Acid Rain. It is deemed ‘permanent’ because the reaction is not reversible:
\[
\text{CaCO}_3(\text{s}) + 2\text{H}^+(_{\text{aq}}) \rightarrow \text{Ca}^{2+}(_{\text{aq}}) + \text{H}_2\text{O(\text{l})} + \text{CO}_2
\]
2. Temporary Hard Water is due to \(\text{CO}_2\) dissolved in regular rainwater. It is deemed ‘temporary’ because it’s a reversible reaction
\[
\text{CaCO}_3(\text{s}) + \text{CO}_2(_{\text{aq}}) + \text{H}_2\text{O(\text{l})} \leftrightarrow \text{Ca}^{2+}(_{\text{aq}}) + 2\text{HCO}_3^-(_{\text{aq}}) + \text{heat}
\]
How do you soften hard water (lower the \([\text{Ca}^{2+}]\) and/or \([\text{Mg}^{2+}]\))?

Permanently hard water:
Use washing soda (\(\text{Na}_2\text{CO}_3\)), or any substance that contains an anion that is low solubility with and will precipitate \(\text{Ca}^{2+}\) and/or \(\text{Mg}^{2+}\)
e.g. \(\text{Ca}^{2+}(_{\text{aq}}) + \text{CO}_3^{2-}(_{\text{aq}}) \leftrightarrow \text{CaCO}_3(\text{s})\)
Temporary hard water:
Use heat (increase temperature) to reverse the direction of the formation of Ca\(^{2+}\)(aq), as the equilibrium will shift left (endo).

\[
\text{CaCO}_3(s) + \text{CO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{Ca}^{2+}(aq) + 2\text{HCO}_3^-(aq) + \text{heat}
\]

What's inconvenient about hard water: [http://www.absorblearning.com/media/item.action?quick=v1](http://www.absorblearning.com/media/item.action?quick=v1)
another way to soften hard water: [http://www.absorblearning.com/media/item.action?quick=v2](http://www.absorblearning.com/media/item.action?quick=v2)

**Assignment 10:** Read Hebden p. 103-104 and do p. 104 #76-80

XIII) Altering the Solubility of a Salt
Saturated solution equations (solid on the left, ions on the right) are simply equilibrium equations. In the equilibrium unit, we studied how concentration changes cause a shift.

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

What if \([\text{Ag}^+]\) was increased? What would happen to the equilibrium?

\[\text{Shift left}\]

How would the AgCl\((s)\) change?

Because the amount of AgCl\((s)\) would increase, we can say that the solubility of AgCl (it's ability to dissolve) has decreased.

By adding more Ag\(^+\) to increase \([\text{Ag}^+]\) or more Cl\(^-\) to increase \([\text{Cl}^-]\), the equilibrium is driven left, resulting in more AgCl\((s)\).

This is called the common ion effect, because an ion that is already part of the equilibrium (common to the equilibrium) is being added. However, laboratories don't have solutions that contain solely Ag\(^+\)(aq) available to add to the equilibrium to increase \([\text{Ag}^+]\), so what can be used as a source of Ag\(^+\)(aq) which can be added to the equilibrium?

AgNO\(_3\)(aq)

What can be used as a source of Cl\(^-\)(aq)?

NaCl\(_3\)(aq)

Summary: To decrease the solubility, add a common ion (common ion effect) to shift the equilibrium left.

What if we want to shift the equilibrium to the right? How can we achieve
this? \[ \text{AgCl}(s) \rightleftharpoons \text{Ag}^{+}(aq) + \text{Cl}^{-}(aq) \]

**Find a way to decrease \([\text{Ag}^+]\) or \([\text{Cl}^-]\)\)**

A shift to the right causes more \(\text{AgCl}(s)\) to _______ dissolve ________, thereby increasing its solubility.

What can be done to the equilibrium to decrease \([\text{Ag}^+]\) or \([\text{Cl}^-]\) to cause a shift to the right? \[ \text{AgCl}(s) \rightleftharpoons \text{Ag}^{+}(aq) + \text{Cl}^{-}(aq) \]

Add an ion that is low solubility with \(\text{Ag}^+\) or \(\text{Cl}^-\), causing one of those two ions to form a solid and ppt out of solution, thereby decreasing the ion conc and causing a shift right.

For example, adding \(\text{Pb(NO}_3\text{)2}\). The \(\text{Pb}^{2+}\) would ppt with the \(\text{Cl}^-\) to form \(\text{PbCl}_2\). \([\text{Cl}^-]\) would decrease, causing a shift right, causing more \(\text{AgCl}(s)\) to dissolve, thereby increasing its solubility.

Summary: By adding an ion that will precipitate out one of the product ions, the equilibrium is driven right resulting in more \(\text{AgCl}(s)\) dissolving, and hence an increase in solubility.

**Examples:**

1. In which of the following 0.10M solutions would \(\text{CaC}_2\text{O}_4\) be least soluble in and most soluble in? *Start by writing a saturated solution equation for \(\text{CaC}_2\text{O}_4\). \[ \text{CaC}_2\text{O}_4(s) \rightleftharpoons \text{Ca}^{2+}(aq) + 2\text{C}_2\text{O}_4^{2-}(aq) \]

   - **NaOH**, \(\uparrow\) low solub of \(\text{Ca}^{2+}\), thus \(\text{Ca(OH)}_2\) ppt \(\uparrow\) \([\text{Ca}^{2+}]\) \(\downarrow\) \(\text{shift right}\) \(\text{inc solubility}\) \(\text{most soluble}\)
   - **KCl**, neither \(\text{K}^+\) or \(\text{Cl}^-\) will affect the equil
   - **\(\text{Ca(NO}_3\text{)2}\)**, \(\text{Ca}^{2+}\) common ion \(\leftarrow [\text{Ca}^{2+}]\uparrow\) \(\text{shift left}\) \(\text{least soluble}\)

2. In which of the following 0.10M solutions would \(\text{PbCl}_2\) be most soluble? Least soluble? \[ \text{PbCl}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2\text{Cl}^-/(aq) \]

   - **HCl**, \(\uparrow\) common ion \(\text{less soluble}\) \(0.10\text{M Cl}^-\) added
   - **MgCl\(_2\)**, \(\uparrow\) common ion \(\text{less soluble}\) \(2\text{Cl}^- = 0.20\text{M Cl}^-\) least soluble
   - **\(\text{AgNO}_3\)**, \(\text{ppt}\) with \(\text{Cl}^-\) to form \(\text{AgCl}\) \(\downarrow\) \([\text{Cl}^-]\) \(\downarrow\) \(\text{more soluble}\) most soluble
   - **\(\text{NH}_4\text{NO}_3\)**, no effect

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Assignment 11:

1. In which of the following 0.10M solutions would Sr(OH)₂ be least soluble in?
   - A. Sr(NO₃)₂
   - B. MgS
   - C. NaCl
   - D. KBr

2. In which of the following 0.10M solutions would NaCl be most soluble in?
   - A. H₂O
   - B. AgNO₃
   - C. NH₄Cl
   - D. HNO₃

3. Hebden p. 108 #81, 82, 84-86