

III.1 SOLUBILITY CONCEPT REVIEW

Read Hebden p. 73 – 76 and review basic solubility definitions.

Soluble means _____

Insoluble means _____

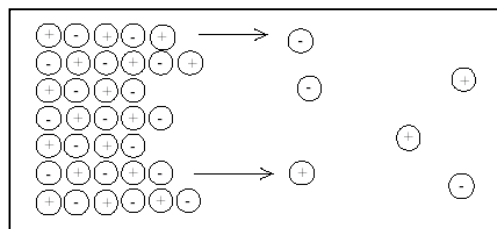
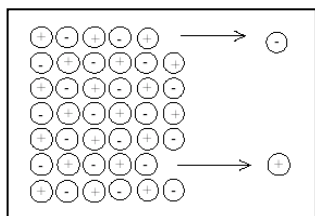
The Dissolving Process

IONIC Solutions	MOLECULAR Solutions (Covalent compounds)
made up of	made up of
They are electrolytes.	They are non-electrolytes
e.g.	e.g.
Ionic compounds DISSOCIATE when placed in water:	Molecular compounds do not dissociate:
NaCl _(s)	I _{2(s)}
KMnO _{4(s)}	C ₁₂ H ₂₂ O _{11(s)}
CH ₃ COOH _(l)	
organic acids (-COOH) are molecular, but are PARTIALLY IONIC when dissolved in water.	

Example 1. For each solution, write I for ionic or partially ionic and M for molecular.

- | | | | |
|--|--|---|--------------------------------------|
| a) NiCl ₂ | b) CH ₃ OH | c) CH ₃ CH ₂ COOH | d) Fe(NO ₃) ₃ |
| e) K ₂ Cr ₂ O ₇ | f) C ₆ H ₁₂ O ₆ | g) PCl ₃ | h) CsBr |

Solubility Equilibrium



1. When an ionic solid is first put into water, the rate of dissolving is (high/low) and the rate of precipitation is (high/low).
2. As dissolving continues, the rate of precipitation (increases/decreases).
3. At some time, rate of dissolving = rate of precipitation = **SOLUBILITY EQUILIBRIA**.

A **SATURATED SOLUTION** is _____

The **SOLUBILITY** of a substance is _____

MOLAR SOLUBILITY is the molarity of a saturated solution

III.2 SOLUBILITY CALCULATIONS

Molar Solubility

Example 2. 0.49 g of AgBrO_3 is required to saturate 250 mL of water. What is the molar solubility of AgBrO_3 ?

Step 1: convert mass to moles

Step 2: substitute values into

$$C = n/V$$

Dilution Calculations

Use the formula:

$$C_1V_1 = C_2V_2$$

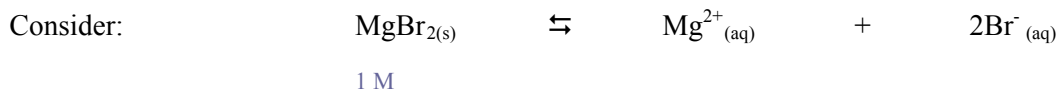
$$(\text{initial conc.})(\text{initial volume}) = (\text{final conc.})(\text{final volume})$$

Example 3. Suppose you have 300.0 mL of 1.2 M HCl. How much water do you need to add to dilute it to 0.50 M HCl?

Step 1: Write dilution formula

Step 2: Substitute values into formula and solve for missing variable

Calculating Concentration of Individual Ions



Example 4. 8.5 g of MgCl_2 is dissolved in 2.0L. What is $[\text{Mg}^{2+}]$ and $[\text{Cl}^-]$?

Step 1: Convert $[\text{MgCl}_2]$ in mol/L

Step 2: Write dissociation equation

Step 3: Use mole ratio to calculate individual conc.

Example 5. 500 mL of 0.8 M Li_2CO_3 is mixed with 500 mL of 0.5 M BeF_2 . Find the concentration of all ions.

Step 1: Calculate new conc. of substances after mixing ($C_1V_1 = C_2V_2$)

Step 2: Write dissociation equations

Step 3: Use mole ratio to calculate individual conc.

Do Hebden Set 14 #1, 2, 6, 9, 10, 16, 17, 18 (a, b, c) (p.74–81)

III.3 SOLUBILITY OF IONIC COMPOUNDS

PREDICTING SOLUBILITY

Use the “Solubility of Common Compounds in Water” Table in your DATA BOOKLET p. 4.

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
Chloride, Cl^- or Bromide, Br^- or Iodide, I^-	All others <hr style="border-top: 1px dashed black;"/> Ag^+ , Pb^{2+} , Cu^+	Soluble <hr style="border-top: 1px dashed black;"/> Low Solubility

Example 6. Is $\text{FeCl}_3(\text{s})$, $\text{MgSO}_4(\text{s})$, $\text{Na}_2\text{S}(\text{s})$ soluble in water?

Step 1: Look in the Negative ion column for negative ion and see if positive ion is soluble

PREDICTING A PRECIPITATE

A **PRECIPITATE** is a solid formed when two liquids or aqueous solutions react.

Example 7. Will a precipitate form from mixing solutions of $\text{BaS}_{(\text{aq})}$ and $\text{AgNO}_3(\text{aq})$?

Step 1: Do double replacement, don't worry about balancing equation

Step 2: Check their solubility. Low solubility = solid, soluble = aqueous

III.4 WRITING EQUATIONS

Example 8. Write the formula, complete ionic, and net ionic equations when solutions of calcium chloride and sodium sulfate are mixed.

Molecular/Formula Equation

= balanced chemical equation

Step 1: Double replace reactants.

Step 2: Check solubility of products. Must indicate (s), (l), (g) states.

Complete Ionic Equation

shows all soluble ionic compounds broken into ions.

Step 1: From formula equation, break soluble species (aq) into ions. Leave solids alone.

Net Ionic Equation

shows ONLY the ions actively involved in forming the precipitate.

Step 1: From complete ionic equation, OMIT spectator ions.

OR Write equation as follows

Ions forming ppt \rightleftharpoons formula of ppt

Do Hebden Set 15 #21, 22, 24, 25 (a, c, e, h, j, l) (p.83–87)

III.5 USING SOLUBILITY AND PRECIPITATION TO IDENTIFY OR SEPARATE IONS

QUALITATIVE ANALYSIS means using experimental procedures to _____

Process

Example 9. You suspect a solution contains Ba^{2+} or Cu^{+} . Explain the process for identifying the unknown ion.

Step 1: Look on Solubility Table for anions (first column) that can precipitate with cations.

Step 2: Choose an anion that will only form a precipitate with ONE of the cations.

Step 3: Conclude which cation is present.

Ba^{2+}				
Cu^{+}				

Experimental Procedure

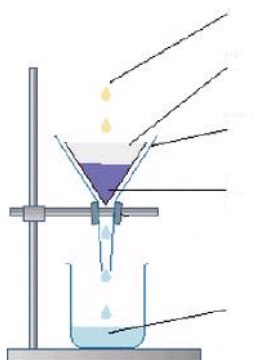
Example 10. Suggest an experimental procedure to identify the unknown ions for the previous example.

In reality, you can't just add ions to a solution because they don't exist by themselves in a solution. All solutions are neutral, positive ions must be with a negative ion.

When adding ions, always choose _____

For **POSITIVE IONS**, combine with:

For **NEGATIVE IONS** combine with:



Procedure:

1. Add some 1 M _____ solution until no more precipitate forms.
2. Filter the solution. The precipitate _____ will remain in the filter paper.
3. To the filtrate, add
- 4.

Example 11. You have 3 unlabelled test tubes containing I^- , Cu^{2+} , and Ca^{2+} . What procedures could you use to test these and find out which is which?

Step 1: Look on Solubility Table for anions (first column) that can precipitate with cations.

Step 2: Choose an anion that will only form a precipitate with ONE of the cations.

Step 3: Conclude which cation is present.

Step 4: Write procedure, remember to add soluble compounds not ions alone

I^-				
Cu^{2+}				
Ca^{2+}				

Read through examples on p. 88 – 89. Do Hebden Set 16 #26 – 29, 31, 33, 35, 37 (p.90-91)

III.6 THE SOLUBILITY PRODUCT

SOLUBILITY PRODUCT =

Equilibrium expression = SOLUBILITY PRODUCT EXPRESSION.

$K_{eq} = K_{sp}$

Example 12. Write the SOLUBILITY PRODUCT EXPRESSION for a saturated solution of MgF_2

Step 1: Write net ionic equation.

Step 2: Write the K_{eq} expression as before. Omit solids and remember to change coefficients to exponents.

Meaning of K_{sp}

High K_{sp} value =

Low K_{sp} value =

K_{sp} is CONSTANT; See The DATA BOOKLET p.5.

K_{sp} CALCULATIONS

Type 1: Find K_{sp} from solubility

Example 13. What is the K_{sp} for Ag₂CrO₄ if the solubility is 1.31×10⁻⁴ M?

Step 1: Write net ionic equation.

Step 2: Use mole ratios to find ion conc.

Step 3: Write K_{sp} expression

Step 4: Substitute conc. into K_{sp} expression. Solve. Check s.f., omit units.

Example 14. If 1.64 × 10⁻⁶ g of Zn(OH)₂ can dissolve in 1.0 mL of water, what is the K_{sp}?

Step 1: Convert solubility from g/L to mol/L. (use molar mass)

Step 2: Write net ionic equation.

Step 3: Use mole ratios to find ion conc.

Step 4: Write K_{sp} expression. omit (s)

Step 5: Substitute conc. into K_{sp} expression. Solve. Check s.f., omit units.

Type 2: Find solubility from K_{sp}

Example 15. What is the [Ca²⁺] and [CO₃²⁻] if the K_{sp} for CaCO₃ is 4.8 × 10⁻⁹?

Step 1: Look up K_{sp}.
Write net ionic equation

Step 2: Let unknown concentration of salt = "s".
Use mole ratios to represent other concentrations.

Step 3: Write K_{sp}

expression.

Step 4: Substitute “s” into K_{sp} expression and solve.

Example 16. The K_{sp} for MgF_2 is 6.4×10^{-9} M. What is the $[Mg^{2+}]$ and $[F^-]$?

Step 1: Look up K_{sp} .
Write net ionic equation

Step 2: Let unknown concentration of salt = “s”.
Use mole ratios to represent other concentrations.

Step 3: Write K_{sp} expression

Step 4: Substitute “s” into K_{sp} expression and solve.

Read through III.6 in Hebden and Do Hebden Set 17 #40, 41, 42, 45, 47, 48, 52, 55 (p.91-95)

III.7 K_{SP} AND PRECIPITATE FORMATION

So far, we are able to use the Solubility Table to predict whether a precipitate will form, BUT the Solubility Table is only good when solubility is > 0.1 mol/L.

If we mix ions in LOW ENOUGH CONCENTRATION, a precipitate will not form.

For example, the K_{sp} of $CaCO_3 = [Ca^{2+}][CO_3^{2-}] = 5.0 \times 10^{-9}$ = saturated solution on the verge of precipitating.

This is the SATURATION POINT.

The Trial Ion Product/Trial K_{sp} / Q

TRIAL ION PRODUCT (TIP), Trial K, Q = $\frac{[\text{products}]}{[\text{reactants}]}$

TIP < K_{sp}

PPT will NOT form

TIP > K_{sp}

PPT will form

Predicting Precipitate Calculations

Type 3: Will ions precipitate when mixed together

1. Write net ionic equation
2. Write K_{sp} expression and Look up K_{sp} value (in Data Booklet p. 5)
3. Find Ion Concentration and plug them in to calculate TIP
4. Compare TIP with K_{sp}

Example 17. Will a precipitate form when 4.5 mL of 5.0×10^{-4} M Cu^+ is mixed with 2.0 mL of 6.0×10^{-5} M of I^- ?

Step 1: Write net ionic equation.

Step 2: Write K_{sp} expression and look up K_{sp} value.

Step 3: Find ion concentrations. Use $C_1V_1 = C_2V_2$ or dilution factor.

Step 4: Write expression for TIP and calculate TIP (using initial concentrations)

Step 5: Compare TIP with K_{sp} .

Type 4: Finding MAX. possible conc. = MIN. conc. needed to ppt.

1. Write net ionic equation
2. Write K_{sp} expression and Look up K_{sp} value (in Data Booklet p. 5)
3. Substitute given values into TIP and Solve for unknown

How much of one ion can be present in solution if there is a certain amount of another ion?

Consider a SATURATED solution of PbI_2 . K_{sp} for PbI_2 is $8.5 \times 10^{-9} = [\text{Pb}^{2+}][\text{I}^-]^2$

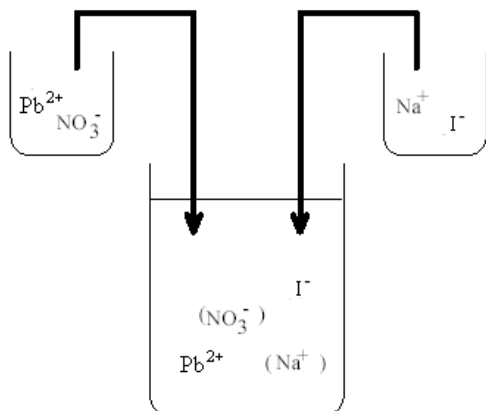
For every Pb^{2+} or I^- that dissolves, PbI_2 forms back



- only a limited number of Pb^{2+} and I^- can exist dissolved in solution

- this limit is determined by $K_{sp} = [Pb^{2+}][I^-]^2 = 8.5 \times 10^{-9}$.
- So $[Pb^{2+}] \times [I^-]^2$ CANNOT EXCEED 8.5×10^{-9}

The Pb^{2+} and I^- ions can come from different sources.



When the ions come from different sources:

eg. Pb^{2+} from 0.1M $PbNO_3$ and I^- from 0.01M NaI

$[Pb^{2+}]$ does not have to equal $[I^-]$

BUT when $[Pb^{2+}] \times [I^-]^2$ exceeds K_{sp} (8.5×10^{-9}) then

PPT will form.

Example 18. If we have a 3.0×10^{-4} M solution of Pb^{2+} , what $[I^-]$ is required to just start the precipitation of PbI_2 from the solution?

Step 1: Write net ionic equation.

Step 2: Write K_{sp} expression and look up K_{sp} value.

Step 3: Substitute values into TIP and solve for unknown concentration.

Looking at the results:

If $[I^-]$ is less than _____, the solution will be _____

If $[I^-]$ is just equal to _____, the solution will be _____

If more I^- is added, it will _____

Do Hebden Set 18 #56, 59, 64, 68 [Type 1], 58, 62, 66, 67 [Type 2] (p.98-99)

III.8 APPLICATIONS OF SOLUBILITY - TITRATION

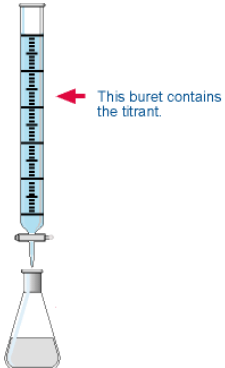
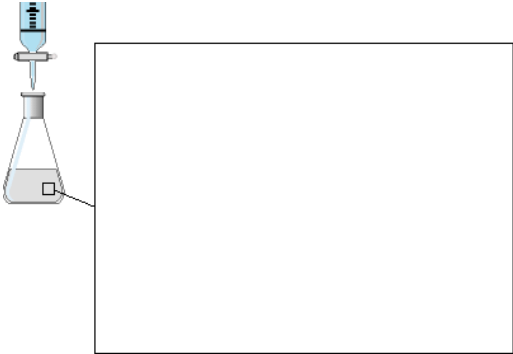
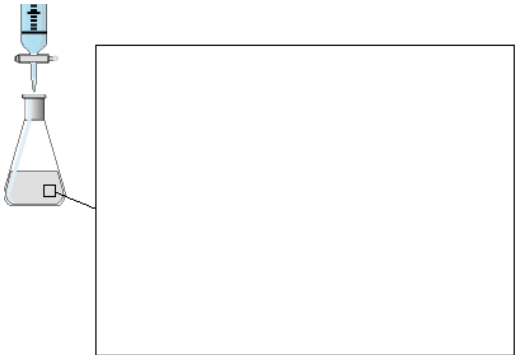


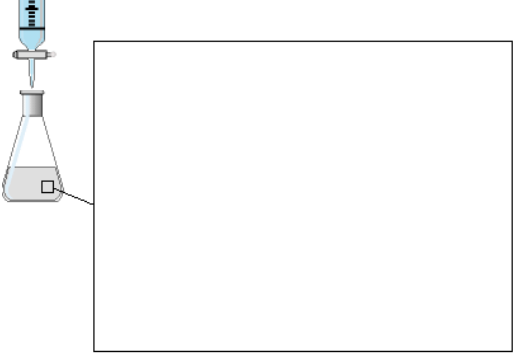
TITRATION is a procedure used to _____

STANDARD SOLUTION (titrant) is a solution that has an _____ concentration

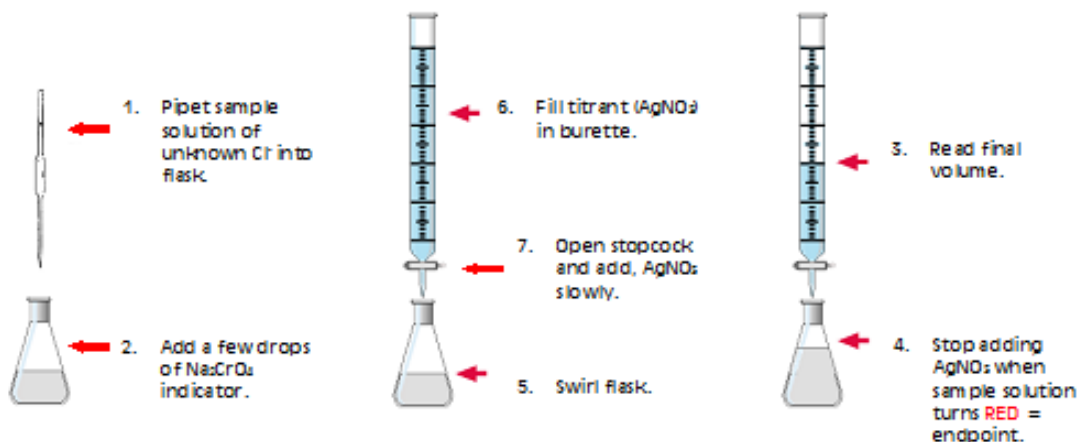
SAMPLE SOLUTION is the solution with the _____ concentration

INDICATOR is a substance that will change colour to show that the titration is complete.

Principle of Method

	
<p>1. Titration of unknown Cl^- ions with AgNO_3</p>	<p>2. Ag^+ added will react with Cl^- and form a precipitate $\text{AgCl}_{(s)}$. $\text{Ag}^+ + \text{Cl}^- \rightleftharpoons \text{AgCl}_{(s)}$</p>
	
<p>3. Stoichiometric/equivalence Point</p>	<p>4. $\text{AgCl}_{(s)}$ will stop forming</p>
	
<p>5. CrO_4^{2-} is the indicator. Note: AgCl is less soluble than Ag_2CrO_4, so it will ppt before Ag_2CrO_4</p>	<p>6. When all the Cl^- ions have reacted, the next drop of Ag^+ will react with CrO_4^{2-} and turn red.</p>

Procedure



Example 19. A 50.0 mL solution with an unknown $[Cl^-]$ was titrated with 0.250 M $AgNO_3$. A precipitate started to form after 32.8 mL of $AgNO_3$ was added.

- a) What is a precipitate that definitely forms at the equivalence point? _____
- b) What is the $[Cl^-]$ in the 50.0 mL solution?

Step 1: Calculate moles of standard solution used to reach equivalence point. (mol = CV)
Step 2: Write net ionic equation for precipitate.
Step 3: Find moles of Cl^- . At equivalence point: moles Ag^+ = moles Cl^-
Step 3: Calculate $[Cl^-]$. $C = mol/V$

Example 20. Titration of Ag Ion Concentration.

Indicator = $Fe(NO_3)_3$

Standard solution = $Ag^+_{(aq)} + SCN^-_{(aq)} \rightleftharpoons AgSCN_{(s)}$
 colourless
 white precipitate

Endpoint: When all the Ag^+ has reacted, SCN^- will react with the indicator Fe^{2+} to form $FeSCN^{2+}$

Unknown $[Ag^+]$ $Fe^{2+}_{(aq)} + SCN^-_{(aq)} \rightleftharpoons FeSCN^{2+}_{(aq)}$
 pale yellow dark red

A solution containing silver ions (Ag^+) is titrated with 0.200 M KSCN solution to find the $[\text{Ag}^+]$ in the sample. The indicator $\text{Fe}(\text{NO}_3)_3(\text{aq})$ is used to signal when the stoichiometric point is reached. It is found that 15.6 mL of 0.200 M KSCN is needed to titrate a 25.0 mL sample of Ag^+ solution. Determine the $[\text{Ag}^+]$ in the sample. (answer: $[\text{Ag}^+] = 0.125 \text{ M}$)

Step 1: Calculate moles of standard solution used to reach equivalence point. ($\text{mol} = CV$)

Step 2: Write net ionic equation for precipitate.

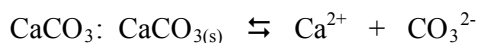
Step 3: Find moles of the unknown using the mole ratio from the net ionic equation.

Step 3: Calculate $[\text{Ag}^+]$. $C = \text{mol}/V$

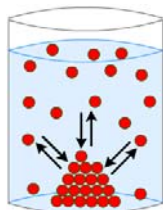
Read III.8 in Hebden and Do Set 19 #70-72 [Cl^- Titrations], 76-80 [Hardness of water] (p.98-99)

III.10 COMMON ION EFFECT – SOLUBILITY AND LE CHATELIER

The solubility of a salt can be changed by applying Le Chatelier's Principle. Consider a saturated solution of



Let's disturb the equilibrium by adding CaCl_2 .



CaCl_2 dissolves to form _____ + _____ ions.

($\text{Ca}^{2+} / \text{Cl}^-$) is a spectator ion, but ($\text{Ca}^{2+} / \text{Cl}^-$) will affect the equilibrium.

So adding ($\text{Ca}^{2+} / \text{Cl}^-$) causes the equilibrium to shift (right/left), forming (more/less) of $\text{CaCO}_{3(s)}$.

COMMON ION EFFECT: _____

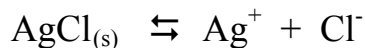
Example 21. Predict which compounds would decrease the solubility of $\text{CaCO}_{3(s)}$.



Stress	Ions	Effect on solubility of CaCO_3 (increase/decrease/no effect)	Reason for effect
$\text{Ca}(\text{NO}_3)_2$			

KNO ₃				
K ₂ CO ₃				
CaCO ₃				

We can also use Le Chatelier's Principle to INCREASE SOLUBILITY. Consider a saturated solution of AgCl.



To increase solubility (make it dissolve more), we need the equilibrium to shift (left/right). The reaction will shift this way if we can decrease the concentration of [Ag⁺] or [Cl⁻].

Look at the Solubility Table and find ions the will react with Ag⁺

Cl⁻

Remember you have to add ions as a SOLUBLE compound. Add positive ions with _____

Add negative ions with _____

Example 22. How can we increase the solubility CuI_(s)? Suggest 2 different compounds that can be added.

Show equilibrium equations to show how it works.

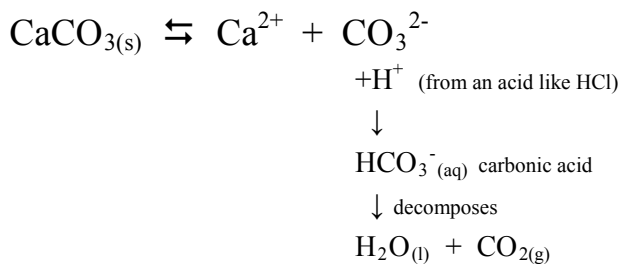
Step 1: Write net ionic equation

Step 2: Remove common ion by adding an ion that will ppt out the common ion. Check Solubility Table.

Step 4: Add ion as soluble salt. Combine cation with NO₃⁻ and anion with Na⁺.

Adding an ACID to an anion that is a base will INCREASE SOLUBILITY of the compound.

For example:



Read Examples in III.10 (p. 106 – 107) and Do Set 20 #81-86 (p.108)