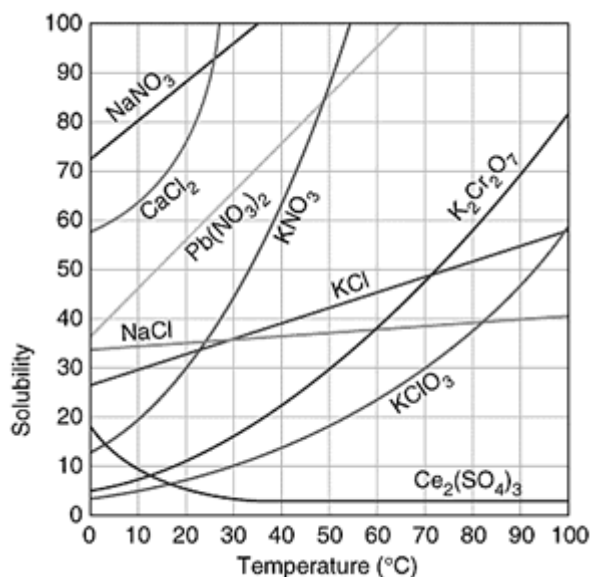


UNIT 8

SOLUBILITY AND PRECIPITATION REACTIONS

Student Version



Source: www.dynamicscience.com.au

Contents

- Solubility and Solubility Curves
- Solubility and Precipitation Reactions
- Qualitative Analysis Part 3
- Hard and Soft Water

Key words: aqueous solution, solvent, solute, concentration, molarity, strong electrolyte, weak electrolyte, soluble, solubility, saturated solution, sparingly soluble, insoluble, solubility curve, crystallisation, recrystallisation, precipitate, precipitation, hard water, soft water, limescale, sum, lather, ion exchange

Units which must be completed before this unit can be attempted:

- Unit 1 – Atoms and the Periodic Table
- Unit 2 – Particles, Bonding and Structures
- Unit 3 – Amount of Substance and Measurement
- Unit 4 – Introduction to Physical Chemistry
- Unit 5 – Acids, Bases and Salts
- Unit 6 – Redox Reactions

Estimated Teaching Time: 8 hours

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS

UNIT 8 SUMMARY AND SYLLABUS REFERENCE

Lesson	Title and Syllabus Reference
1	<p>Solubility and Saturated Solutions CA8a solubility – general principles (saturated and unsaturated solutions, saturated solution as an equilibrium system, solubility expressed in mol dm^{-3}, solubility of sparingly soluble salts - complete dissociation of the portion that dissolves (qualitative treatment only)); CA13i General Skills and Principles - measurement of mass and volume; CA13aiii General Skills and Principles – filtration; CA13biii acid-base titrations - the use of standard solutions of acids and alkalis and the indicators methyl orange and phenolphthalein to determine the solubility of acids and bases</p>
2	<p>Crystallisation and Solubility Curves CA8a solubility – general principles (solubility curves and their uses, relationship between solubility and crystallisation, crystallisation as a method of purification); CA13aiii General Skills and Principles – filtration and recrystallisation</p>
3	<p>Precipitation CA4aiii chemical equations (precipitation); CA8b practical application of solubility (generalisations of solubility of salts)</p>
4	<p>Preparation of Insoluble Salts CA13aiii filtration; CC4a methods of preparation of simple salts; ISA7.4 methods of preparation of salts (preparation of salts using the following methods: precipitation)</p>
5	<p>Qualitative Analysis (Cations) – Precipitation Reactions CA8b practical application of solubility (generalisations of solubility of salts and their applications in qualitative analyses)); CA13ci characteristic tests of the following cations with dilute NaOH(aq): Ca^{2+}, Pb^{2+}, Cu^{2+}, Fe^{2+}, Fe^{3+}, Al^{3+} and Zn^{2+}, confirmatory tests for these cations</p>
6	<p>Qualitative Analysis (Anions) – Precipitation Reactions CA8b practical application of solubility (generalisations of solubility of salts and their applications in qualitative analyses)); CA13cii confirmatory tests for the following ions: Cl^-, SO_3^{2-}, CO_3^{2-}, SO_4^{2-}; CA13civ characteristic test for the following gases: CO_2 and SO_2</p>
7	<p>Hard and Soft Water CC7biii water and Solution - hardness of water: causes and methods of removing it (advantages and disadvantages of hard water and soft water, experiments to compare the degree of hardness in different samples of water); ISA9.2 hardness and softness of water (advantages and disadvantages of hard and soft water, causes of hardness of water (Ca^{2+}, Mg^{2+}, Fe^{2+} ions), softening hard water (addition of washing soda, ion exchange, boiling and distillation))</p>
8	<p>Unit 8 Revision and Summary</p>

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS

The Periodic Table of the Elements

1	2	3	4	5	6	7	0											
(1) 6.9 Li lithium 3	(2) 9.0 Be beryllium 4	(3) 39.1 K potassium 19	(4) 47.9 Ti titanium 22	(5) 50.9 V vanadium 23	(6) 52.0 Cr chromium 24	(7) 54.9 Mn manganese 25	(8) 55.8 Fe iron 26	(9) 58.9 Co cobalt 27	(10) 58.7 Ni nickel 28	(11) 63.5 Cu copper 29	(12) 65.4 Zn zinc 30	(13) 10.8 B boron 5	(14) 12.0 C carbon 6	(15) 14.0 N nitrogen 7	(16) 16.0 O oxygen 8	(17) 19.0 F fluorine 9	(18) 4.0 He helium 2	
23.0 Na sodium 11	24.3 Mg magnesium 12	40.1 Ca calcium 20	87.6 Sr strontium 38	88.9 Y yttrium 39	88.9 Sc scandium 21	91.2 Zr zirconium 40	91.2 Ru ruthenium 44	102.9 Rh rhodium 45	106.4 Pd palladium 46	107.9 Ag silver 47	112.4 Cd cadmium 48	27.0 Al aluminium 13	28.1 Si silicon 14	31.0 P phosphorus 15	32.1 S sulfur 16	35.5 Cl chlorine 17	39.9 Ar argon 18	
85.5 Rb rubidium 37	137.3 Ba barium 56	138.9 La * lanthanum 57	178.5 Hf hafnium 72	180.9 Ta tantalum 73	183.8 W tungsten 74	186.2 Re rhenium 75	190.2 Os osmium 76	192.2 Ir iridium 77	195.1 Pt platinum 78	197.0 Au gold 79	200.6 Hg mercury 80	69.7 Ga gallium 31	72.6 Ge germanium 32	74.9 As arsenic 33	79.0 Se selenium 34	79.9 Br bromine 35	83.8 Kr krypton 36	
[223] Fr francium 87	[226] Ra radium 88	[227] Ac † actinium 89	[267] Rf rutherfordium 104	[268] Db dubnium 105	[271] Sg seaborgium 106	[272] Bh bohrium 107	[270] Hs hassium 108	[276] Mt meitnerium 109	[281] Ds darmstadtium 110	[280] Rg roentgenium 111	Elements with atomic numbers 112-116 have been reported but not fully authenticated	114.8 In indium 49	118.7 Sn tin 50	121.8 Sb antimony 51	127.6 Te tellurium 52	126.9 I iodine 53	131.3 Xe xenon 54	
												204.4 Pb lead 82	207.2 Po polonium 84	209.0 Bi bismuth 83	[209] Po polonium 84	[210] At astatine 85	[222] Rn radon 86	
												162.5 Dy dysprosium 66	164.9 Ho holmium 67	167.3 Er erbium 68	168.9 Tm thulium 69	173.1 Yb ytterbium 70	175.0 Lu lutetium 71	
												158.9 Tb terbium 65	157.3 Gd gadolinium 64	157.3 Er erbium 68	158.9 Tm thulium 69	173.1 Yb ytterbium 70	175.0 Lu lutetium 71	
												[251] Cf californium 98	[252] Es einsteinium 99	[257] Fm fermium 100	[258] Md mendelevium 101	[259] No nobelium 102	[262] Lr lawrencium 103	
												150.4 Sm samarium 62	152.0 Eu europium 63	157.3 Gd gadolinium 64	162.5 Dy dysprosium 66	164.9 Ho holmium 67	167.3 Er erbium 68	173.1 Yb ytterbium 70
												[244] Pu plutonium 94	[243] Am americium 95	[247] Cm curium 96	[251] Cf californium 98	[252] Es einsteinium 99	[257] Fm fermium 100	[262] Lr lawrencium 103
												[145] Pm promethium 61	150.4 Sm samarium 62	152.0 Eu europium 63	157.3 Gd gadolinium 64	162.5 Dy dysprosium 66	164.9 Ho holmium 67	167.3 Er erbium 68
												[237] Np neptunium 93	[244] Pu plutonium 94	[247] Cm curium 96	[251] Cf californium 98	[252] Es einsteinium 99	[257] Fm fermium 100	[262] Lr lawrencium 103
												144.2 Nd neodymium 60	150.4 Sm samarium 62	152.0 Eu europium 63	157.3 Gd gadolinium 64	162.5 Dy dysprosium 66	164.9 Ho holmium 67	167.3 Er erbium 68
												232.0 Pa protactinium 91	[244] Pu plutonium 94	[247] Cm curium 96	[251] Cf californium 98	[252] Es einsteinium 99	[257] Fm fermium 100	[262] Lr lawrencium 103
												140.1 Ce cerium 58	140.9 Pr praseodymium 59	144.2 Nd neodymium 60	150.4 Sm samarium 62	162.5 Dy dysprosium 66	164.9 Ho holmium 67	167.3 Er erbium 68
												90 Th thorium 90	232.0 Pa protactinium 91	[244] Pu plutonium 94	[247] Cm curium 96	[251] Cf californium 98	[252] Es einsteinium 99	[257] Fm fermium 100

* 58 – 71 Lanthanides

† 90 – 103 Actinides

Lesson 1 – What is solubility and what are saturated solutions?

a) Solubility and Solubility Curves

(i) Solubility and Saturated Solutions



Summary Activity 1.1: What is a solution?

- What do the terms “solution”, “solute” and “solvent” mean?
- What is the most common solvent? What is the name given to solutions containing this solvent?
- Can you give an example of a solution?
- What is molarity?
- What is an electrolyte? What is the difference between a strong and a weak electrolyte?

- In SS1 (Unit 2 – Particles, Bonding and Structure) you learned that solutions consisted of a solute dissolved in a solvent
- **Solubility** is the ability of a solute to dissolve in a solvent; the solubility of a solute in a solvent is defined as the maximum possible **concentration** of the solute in that solvent; in this course we will consider only **aqueous solutions** (ie solutions in which the solvent is water)
- There are different ways to measure concentration, but the one we use in this course is the moles of solute per cubic decimetre of solution (mol dm^{-3}); this is also known as **molarity**
- The solubility of solutes varies with temperature, so when describing the solubility of a solute it is important to state the temperature; generally, the solubility of solids and liquids increases with increasing temperature but the solubility of gases decreases with increasing temperature
 Eg The solubility of sodium chloride (NaCl) in water at 25 °C is 6.2 mol dm^{-3} .
 The solubility of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in water is 5.1 mol dm^{-3} at 25 °C and 14 mol dm^{-3} at 50 °C
 The solubility of ammonia (NH_3) in water is 32 mol dm^{-3} at 20 °C and 17 mol dm^{-3} at 50 °C
- If the maximum possible molarity of a solution is low (ie below 0.2 mol dm^{-3}) the solute is described as slightly soluble; if the solubility is much lower than this the solute might be described as sparingly soluble or insoluble
- A solution which is at its maximum possible concentration at that temperature is said to be **saturated**; it is not possible to dissolve any more solute in a solution which is saturated; the solubility of a substance in water can also therefore be defined as the concentration of a saturated solution; if a solution is not saturated, its concentration is less than the maximum possible and it should be possible to dissolve more of the solute in it

The solubility of a substance in a solvent is the concentration of a saturated solution of that substance in the solvent

A saturated solution is a solution which is at its maximum possible concentration

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS

- Solubility in water is generally restricted to ionic compounds and molecular substances; metals and giant structures will not dissolve in water unless there is a chemical reaction with the water
 - Ionic compounds, if soluble, are strong electrolytes, which means that they dissociate completely to form ions in solution:
eg $\text{NaCl(s)} \rightarrow \text{Na}^{\text{+}}(\text{aq}) + \text{Cl}^{\text{-}}(\text{aq})$
 - Most molecular substances, if soluble, are usually weak electrolytes or non-electrolytes; weak electrolytes only dissociate slightly to form ions in solution:
eg $\text{NH}_3(\text{g}) \rightarrow \text{NH}_3(\text{g}) \rightleftharpoons \text{NH}_4^{\text{+}}(\text{aq}) + \text{OH}^{\text{-}}(\text{aq})$
 - non-electrolytes do not dissociate at all:
eg $\text{C}_2\text{H}_6\text{O(l)} \rightarrow \text{C}_2\text{H}_6\text{O(aq)}$
 - a very small number of molecular substances are strong electrolytes:
eg $\text{HCl(g)} \rightarrow \text{H}^{\text{+}}(\text{aq}) + \text{Cl}^{\text{-}}(\text{aq})$
- The process of dissolving a solute in a solvent is reversible; it is possible for the solute to separate itself from the solution and form the pure substance (usually as a solid, in which case it settles at the bottom of the container, or as a gas, in which case it escapes); a saturated solution is in a state of **dynamic equilibrium** (the solute is dissolving in the solvent at the same rate as it is separating it out from the solvent)



Test your knowledge 1.2: Understanding Solubility and Saturated Solutions

At 50 °C, it is possible to dissolve up to 244 g of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in 100 cm^3 of water.

- What is the solubility of glucose in water (in mol dm^{-3}) at 50 °C?
- What is the maximum mass of glucose which can dissolve in 25 cm^3 of water at 50 °C?
- A student attempts to dissolve 135 g of glucose in 60 cm^3 of water at 50 °C. Will she succeed? Explain your answer.
- Suggest two ways in which glucose could be crystallised from a solution.

A saturated solution of calcium hydroxide (Ca(OH)_2) is known as limewater. At 25 °C, limewater has a molarity of 0.020 mol dm^{-3} .

- Write an equation, using state symbols for the equilibrium which exists in limewater
- Calculate the minimum mass of calcium hydroxide required to prepare 250 cm^3 of limewater
- A student attempts to dissolve 2.0 g of calcium hydroxide in 100 cm^3 of water at 25 °C. Will he succeed? Explain your answer
- Which is more soluble in water? Glucose or calcium hydroxide? Explain.
- Are glucose and calcium hydroxide electrolytes? Explain your answer.

(ii) Measuring the Solubility of Acids and Bases

- It is possible to measure the solubility of an acid or base in water by carrying out an **acid-base titration**; firstly a saturated solution of the acid or base must be prepared; the concentration of the saturated solution, and hence the solubility of the acid/base, can be determined using an acid-base titration:

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS



Practical 1.3: Determine the solubility of calcium hydroxide in water by titration

- 1) Use a weighing boat to weigh out 1 – 2 g of calcium hydroxide
- 2) Pour the calcium hydroxide into a 250 cm³ volumetric flask
- 3) Add distilled water, shaking continuously, until the base of the meniscus rests on the mark
- 4) Leave for a few days, shaking occasionally
- 5) Filter around 100 cm³ of the solution into a 250 cm³ beaker
- 6) Pipette two 25.0 cm³ portions of the filtrate into a conical flask, so a total of 50.0 cm³ of the Ca(OH)₂ solution is in the conical flask
- 7) Add a few drops of phenolphthalein indicator to the conical flask
- 8) Prepare a burette for first use with 0.05 mol dm⁻³ HCl and titrate the Ca(OH)₂ solution against the HCl solution until the phenolphthalein decolorises
- 9) Repeat the titration until two concordant results have been obtained; note the average titre volume

Write an equation for the reaction between Ca(OH)₂ and HCl

Calculate the molarity of the Ca(OH)₂ solution and hence the molar solubility of Ca(OH)₂ in water

Lesson 2 – What is crystallisation and what are solubility curves?

(iii) Crystallisation



Summary Activity 2.1: Preparing salts

- When you prepared salts in SS2 (Unit 5 – Acids, Bases and Salts and Unit 6 – Redox Reactions), how did you separate the pure salt from the solution? How does this technique work?
- Salts and other solid solutes can usually be separated from their solutions by **crystallisation**:
 - first heat the solution to evaporate off some of the water; as the water is removed the concentration of the solution increases; when the concentration becomes equal to the solubility of the solute (ie the solution becomes saturated) the solute will start to crystallise
 - then allow the solution to cool; the water will continue to evaporate, causing more solute to crystallise; also as the solution cools down the solubility decreases (most solids become less soluble at lower temperatures) so more solute crystallises
 - don't allow all the solvent to evaporate during heating; this will cause the temperature to rise rapidly and this may cause the solute to burn or decompose
- The technique of crystallisation (or recrystallisation) can also be used as a **purification** technique:
 - the pure solid should dissolve in hot water but crystallise in cold water
 - some impurities will not dissolve in water and so they can be filtered out after the solid dissolves in hot water
 - other impurities dissolve even in cold water and remain in solution after the pure solid crystallises



Practical 2.2: Purify a sample of copper sulphate by recrystallisation

- 1) Pour 30 cm³ of distilled water into a 250 cm³ beaker
- 2) Add impure copper sulphate to the beaker using a spatula; stir the mixture gently so that the solid dissolves; keep adding the impure copper sulphate until it stops dissolving (you might need 10 g of copper sulphate or even slightly more)
- 3) Add 3 cm³ of 1 mol dm⁻³ H₂SO₄ and stir the mixture gently
- 4) Place some filter paper in a funnel; place the funnel over an evaporating dish and pour the copper sulphate solution into the funnel
- 5) Now place the evaporating dish containing the filtrate onto a gauze on a tripod above a Bunsen burner (or onto a sand bath); heat the solution until just over half of the liquid has evaporated
- 6) Allow the mixture to cool (it's best to leave it for a few days in a fridge)
- 7) Carefully decant the excess solution and use a clean spatula to scrape any crystals onto a piece of filter paper
- 8) Transfer the crystals onto another piece of filter paper; fold the filter paper over the crystals and squeeze them gently

At what point during this procedure are the insoluble impurities removed?

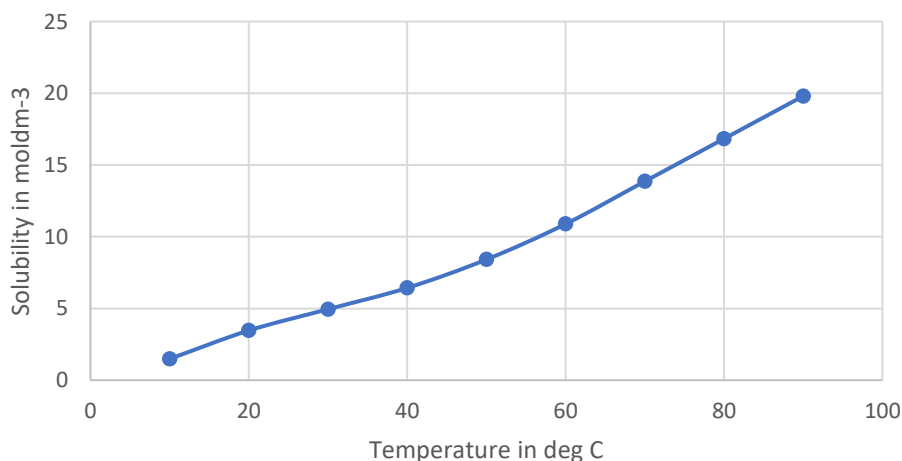
At what point during this procedure are the soluble impurities removed?

Can't do this experiment? Watch this video instead: <https://www.youtube.com/watch?v=SAU-gptAFe0>

b) Solubility Curves

- Most solids are more soluble in water at high temperatures than at low temperatures; it is useful to know the solubility of different solids at different temperatures and **solubility curves** are used for this purpose; the solubility curve for KNO₃ in water is shown below:

Solubility curve for KNO₃ in water





Test your knowledge 2.3: Using solubility curves

- What is the solubility of KNO_3 at $55\text{ }^\circ\text{C}$?
- What is the molarity of a saturated solution of KNO_3 at $15\text{ }^\circ\text{C}$?
- What mass of KNO_3 can be dissolved in 50 cm^3 of water at $25\text{ }^\circ\text{C}$?
- What volume of water is required to dissolve 20 g of KNO_3 at $70\text{ }^\circ\text{C}$?
- 10 g of KNO_3 is dissolved in 10 cm^3 of boiling water at $100\text{ }^\circ\text{C}$ and the solution is allowed to cool. At what temperature will crystals start to form?
- What is the minimum temperature required to dissolve 15 g of KNO_3 in 20 cm^3 of water?
- Is it possible to dissolve completely 30 g of KNO_3 in 25 cm^3 of water at $60\text{ }^\circ\text{C}$?



Extension 2.4: Using solubility curves

Completed Exercise 2.3? Use the solubility curves shown on the first page of this topic to set questions for each other to solve!

Lesson 3 – What is precipitation and what is a precipitation reaction?

c) Solubility and Precipitation Reactions

- In SS1 (Unit 2 – Particles, Bonding and Structure) you learned about why most, but not all, ionic compounds were soluble in water; in SS2 (Topic 5 – Acids, Bases and Salts) you learned about the solubility of a few common bases and salts



Summary Activity 3.1: Solubility of Ionic Compounds

- What causes ionic compounds to dissolve in water?
- Why are some ionic compounds not soluble in water?
- Can you name any ionic compounds which are soluble in water?
- Can you name any ionic compounds which are not soluble in water?

(i) Predicting the solubility of ionic compounds

- Most common ionic compounds are soluble in water (ie have solubilities in water of 0.5 mol dm^{-3} or greater at $25\text{ }^\circ\text{C}$); however there are a number of exceptions; by following a few general rules, it is possible to predict the solubility of most ionic compounds in water:
 - Rule 1: All compounds containing the nitrate ion (NO_3^-) are soluble
 - Rule 2: All compounds containing Group 1 ions (eg H^+ , Li^+ , Na^+ , K^+ etc) and ammonium ions (NH_4^+) are soluble
 - Rule 3: Most compounds containing halide ions (Cl^- , Br^- , I^-) are soluble except the halides of lead (eg PbCl_2) and silver (eg AgBr)
 - Rule 4: Most compounds containing sulphate ions (SO_4^{2-}) are soluble except the sulphates of lead, calcium, strontium and barium (PbSO_4 , CaSO_4 , SrSO_4 , BaSO_4)
 - Rule 5: Most compounds containing hydroxide ions (OH^-) and carbonate ions (CO_3^{2-}) are insoluble except those mentioned in Rule 2 (Rule 5); $\text{Ba}(\text{OH})_2$ and $\text{Sr}(\text{OH})_2$ are also soluble

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS



Test your knowledge 3.2: Predicting the Solubility of Ionic Compounds

Give the formula of each of the following ionic compounds, and predict which will be soluble in water:

- | | |
|--------------------------|---------------------------|
| (a) magnesium nitrate | (i) barium carbonate |
| (b) sodium sulphate | (j) potassium carbonate |
| (c) copper (II) chloride | (k) calcium carbonate |
| (d) silver chloride | (l) copper (II) hydroxide |
| (e) lead (II) bromide | (m) lithium hydroxide |
| (f) copper (II) sulphate | (n) barium hydroxide |
| (g) barium sulphate | (o) magnesium hydroxide |
| (h) magnesium sulphate | |



Extension 3.3: Predicting the Solubility of Ionic Compounds

Completed Exercise 3.2? Use the rules to predict the solubility of some more ionic compounds and test each other!

(ii) Precipitation

- If one aqueous solution of an ionic compound is added to another, the cations in the first solution will come in to contact with the anions in the second solution, and the anions in the first solution will come into contact the cations in the second solution; in other words, two new ionic compounds will be created
Eg if NaCl(aq) and $\text{AgNO}_3\text{(aq)}$ are mixed together, the Na^+ will come into contact with the NO_3^- (NaNO_3) and the Ag^+ will come into contact with the Cl^- (AgCl)
- If both of these new compounds are soluble, the ions all mix together freely and no chemical reaction takes place; if either of the new compounds is insoluble, however, the two ions will join together to form a solid; this solid is called a **precipitate**; the formation of an insoluble ionic compound by mixing two different solutions together is called **precipitation**
Eg if NaCl(aq) and $\text{AgNO}_3\text{(aq)}$ are mixed together:
 - NaNO_3 and AgCl can be formed
 - NaNO_3 is soluble so the Na^+ and NO_3^- ions remain in solution
 - AgCl , is insoluble so the ions join together and a precipitate is observed
 - the ionic equation for the reaction is $\text{Ag}^+\text{(aq)} + \text{Cl}^-\text{(aq)} \rightarrow \text{AgCl(s)}$

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Test your knowledge 3.4: Predicting Precipitation

Predict whether a precipitate will form when the following solutions are mixed. If precipitation occurs, write the ionic equation for the reaction taking place:

- (a) barium chloride solution and sodium sulphate solution
- (b) lead nitrate solution and sodium chloride solution
- (c) copper sulphate solution and sodium hydroxide solution
- (d) sodium carbonate solution and ammonium nitrate solution
- (e) potassium carbonate solution and calcium chloride solution
- (f) magnesium chloride solution and sulphuric acid
- (g) silver nitrate solution and hydrochloric acid
- (h) sodium chloride solution and nitric acid



Extension 3.5: Predicting Precipitation

Completed Exercise 3.4? Use the rules to predict some other pairs of solutions which will form a precipitate when mixed together, and write ionic equations for the reactions taking place!

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS

Lesson 4 – How can we prepare insoluble salts?



Practical 4.1: Observe precipitation reactions

You are provided with six bottles containing the following solutions:

A is silver nitrate (AgNO_3)

B is hydrochloric acid (HCl)

C is sulphuric acid (H_2SO_4)

D is barium chloride (BaCl_2)

E is copper sulphate (CuSO_4)

F is sodium hydroxide (NaOH)

Each bottle should come with its own 10 cm^3 measuring cylinder

- 1) Pour 5 cm^3 of solution A into a test tube
- 2) Add 5 cm^3 of solution B to the same test tube; note whether you see a precipitate or not
- 3) Write “precipitate” or “no precipitate” in the correct box of the table below
- 4) Repeat steps 2 – 3 using solutions C, D, E and F instead of B
- 5) Continue mixing the solutions until each combination has been mixed together once
- 6) Record your observations in the table below:

	A	B	C	D	E	F
A						
B						
C						
D						
E						

Write ionic equations for any reactions taking place

(iii) Preparing Insoluble Salts



Summary Activity 4.2: Preparation of soluble salts

In SS2 (Topic 5 – Acids, Bases and Salts and Topic 6 – Redox Reactions) you prepared some soluble salts.

- Which salts did you prepare and how did you prepare them?
- What types of reaction were used to prepare each salt?
- How was the salt separated from the rest of the reaction mixture?
- Which of the reactants should be in excess?

- Soluble salts can be prepared by neutralisation reactions or by redox reactions; you learned about the preparation of soluble salts in SS2 (Topic 5 - Acids, Bases and Salts and Topic 6 – Redox Reactions); now you will learn about the preparation of **insoluble** salts

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- Insoluble salts can be prepared very easily by precipitation reactions; if a solution containing the anion is added to a solution containing the cation, a precipitate of the insoluble salt is formed; a pure sample of the insoluble salt can be separated from the rest of the mixture by filtration
Eg the addition of lead nitrate solution to sodium chloride solution will result in the precipitation of the insoluble salt lead chloride



Practical 4.3: Prepare a sample of the insoluble salt lead chloride

(caution – lead compounds are toxic; wear gloves, wash your hands thoroughly after carrying out this experiment and do not pour the lead (II) nitrate solution down the sink)

- 1) Pour 5 cm³ of 1 mol dm⁻³ lead (II) nitrate solution into a boiling tube
- 2) Pour 5 cm³ of 2 mol dm⁻³ sodium chloride solution into the same boiling tube
- 3) Place a bung on the boiling tube and shake the contents thoroughly
- 4) Place a piece of filter paper into a funnel and pour the contents of the boiling tube into the funnel; allow the filtrate to run into a beaker
- 5) When the filtration is complete, pour a small quantity of cold distilled water into the funnel
- 6) Use a spatula to scrape the residue off the filter paper and onto another dry piece of filter paper
- 7) Fold the filter paper of the solid sample and squeeze gently

Write an ionic equation for the reaction taking place during this preparation.

What type of reaction is taking place?

Does it matter which reactant is in excess? Explain your answer.

Can't do this experiment? Watch this video instead: <https://www.youtube.com/watch?v=E1ODnGe9LnM>

Lesson 5 – How can we use precipitation reactions to identify cations in solution?

d) Qualitative Analysis Part 3 – Precipitation Reactions

- Qualitative analysis is the experimental identification of a substance or of a particular species present in a substance; you learned about some simple qualitative tests for cations, anions and gases in Unit 5 – Acids, Bases and Salts and Unit 6 – Redox Reactions



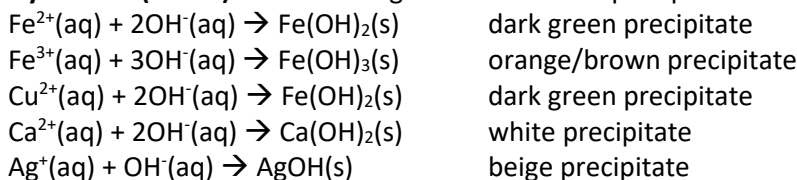
Summary Activity 5.1: Qualitative Analysis of Cations

- Describe two different tests to confirm the presence of H⁺ ions in solution
- Describe a test to confirm the presence of NH₄⁺ ions in solution

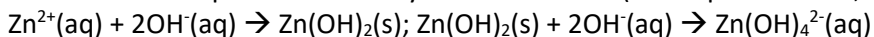
- Precipitation reactions are very useful in qualitative analysis – they are commonly used to identify particular cations and anions in solution:

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS

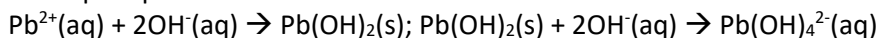
- **Sodium hydroxide (NaOH)** solution can give characteristic precipitates with many common metal ions:



In some cases, the hydroxide precipitates formed are amphoteric – this means that they react with alkali as well as acid, and dissolve in aqueous sodium hydroxide solution (see Topic 5 – Acids, Bases and Salts)



white precipitate which dissolves in excess NaOH



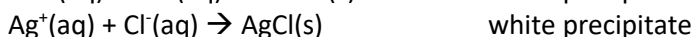
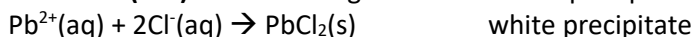
white precipitate which dissolves in excess NaOH



white precipitate which dissolves in excess NaOH

A small number of cations (NH_4^{+} , Na^{+} , K^{+} , Ba^{2+}) do not give a precipitate with hydroxide ions

- **Hydrochloric acid (HCl)** solution can give characteristic precipitates with a small number of metal ions:



Most other cations do not give a precipitate with chloride ions; this reaction allows Pb^{2+} to be distinguished from Zn^{2+} and Al^{3+}

- **Summary: precipitation tests for cations**

Cation	Test	Observation
Fe^{2+}	Add NaOH	Dark green precipitate
Fe^{3+}	Add NaOH	Orange brown precipitate
Cu^{2+}	Add NaOH	Pale blue precipitate
Ca^{2+}	Add NaOH	White precipitate insoluble in excess NaOH
Zn^{2+} or Al^{3+}	Add NaOH Add HCl	White precipitate soluble in excess NaOH No precipitate
Pb^{2+}	Add NaOH Add HCl	White precipitate soluble in excess NaOH White precipitate

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Practical 5.2: Use precipitation reactions to identify cations in solution

You are given eight solutions, labelled A, B, C, D, E, F, G and H; each solution contains one cation from the following: Pb^{2+} , Ca^{2+} , Fe^{2+} , Fe^{3+} , Zn^{2+} , Cu^{2+} , Al^{3+} and Na^+

- 1) Pour 2 cm³ of solution A into a test tube
- 2) Add a few drops of dilute NaOH, then add dilute NaOH until the test tube is almost full; note your observations
- 3) Repeat steps (a) to (b) using solutions B – H
- 4) Pour 2 cm³ of solution A into a test tube
- 5) Add a few drops of dilute HCl; record your observations
- 6) Repeat steps (d) to (e) using solutions B – H

Use your observations to identify which cation is in which solution; present your answers in a table similar to the one below:

Solution	Observations			cation present
	few drops NaOH	excess NaOH	few drops HCl	
A				
B				
C				
D				
E				
F				
G				
H				

Which two cations cannot be distinguished by this combination of tests?

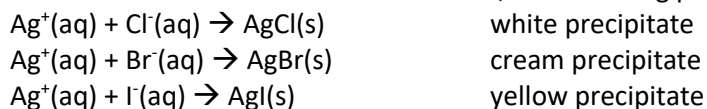
Lesson 6 – How can we use precipitation reactions to identify anions in solution?



Summary Activity 6.1: Qualitative Analysis of Anions

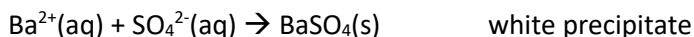
- a) Describe two different tests to confirm the presence of OH^- ions in solution
- b) Describe a test to confirm the presence of NO_3^- ions in solution
- c) Describe a test to confirm the presence of CO_3^{2-} ions in solution
- d) Describe a test to confirm the presence of SO_3^{2-} ions in solution

- **Silver nitrate (AgNO_3) solution** can give characteristic precipitates with most anions; however if these precipitates contain basic ions such as SO_3^{2-} , CO_3^{2-} or OH^- , the precipitates will react with acids; Ag_2SO_3 , AgOH and Ag_2CO_3 are all insoluble but react with acids; these precipitates therefore do not form if **dilute nitric acid (HNO_3)** is added before the silver nitrate solution; the following precipitates form even in acidic conditions:



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- **Barium chloride (BaCl₂) solution** can give characteristic precipitates with SO₃²⁻, SO₄²⁻ and CO₃²⁻, some anions. BaSO₃ and BaCO₃ contain basic ions so react with acids. These precipitates therefore do not form if **dilute hydrochloric acid (HCl)** is added before the barium chloride solution. Under those conditions, only BaSO₄ will form:



- **Summary: precipitation tests for anions**

Cation	Test	Observation
Cl ⁻	Add HNO ₃ and then AgNO ₃	white precipitate
Br ⁻	Add HNO ₃ and then AgNO ₃	cream precipitate
I ⁻	Add HNO ₃ and then AgNO ₃	yellow precipitate
SO ₄ ²⁻	Add HCl and then BaCl ₂	white precipitate
SO ₃ ²⁻	Add BaCl ₂ Then add HCl	White precipitate Dissolves in acid; gas evolved which turns blue litmus red and acidified dichromate paper green
CO ₃ ²⁻	Add BaCl ₂ Then add HCl	White precipitate soluble in excess NaOH Dissolves in acid; gas evolved which turns limewater milky



Practical 6.2: Use precipitation reactions to identify anions in solution

You are given six solutions, labelled A, B, C, D, E and F; each solution contains one anion from the following: NO₃⁻, I⁻, SO₄²⁻, Cl⁻, SO₃²⁻, CO₃²⁻

- 1) Pour 2 cm³ of solution A into a test tube
- 2) Add 2 cm³ of HNO₃, and then a few drops of AgNO₃; note your observations
- 3) Repeat steps (e) – (f) using solutions B – F
- 4) Pour 2 cm³ of solution A into a test tube
- 5) Add a few drops of BaCl₂; note your observations
- 6) Repeat steps (e) – (f) using solutions B – F
- 7) Pour 2 cm³ of each solution which gave a precipitate with BaCl₂ into separate test tubes
- 8) To each test tube, add 2 cm³ of HCl, and then a few drops of BaCl₂; note your observations

Use your observations to identify which anion is in which solution; present your answers in a table similar to the one below:

Solution	Observations			anion present
	HNO ₃ and AgNO ₃	BaCl ₂	HCl and BaCl ₂	
A				
B				
C				
D				
E				
F				

Which two anions cannot be distinguished by this combination of tests? What test could be used to distinguish between these two anions?



Test your knowledge 6.3: Using precipitation to distinguish between different solutions

Describe a simple test to show how you would distinguish between the following solutions:

- (a) sodium chloride and calcium chloride
- (b) zinc nitrate and calcium nitrate
- (c) iron (II) sulphate, copper sulphate and iron (III) sulphate
- (d) zinc nitrate and lead nitrate
- (e) sodium chloride and sodium nitrate
- (f) sodium chloride and sodium sulphate
- (g) sodium sulphate and sodium carbonate



Extension 6.4: Further qualitative analysis

Completed Exercise 6.3? Describe simple tests to distinguish between the following solutions: (these questions require you to use qualitative analysis from previous units)

- (a) sodium nitrate and nitric acid
- (b) sodium hydroxide and sodium carbonate
- (c) sodium nitrate and ammonium nitrate
- (d) sodium nitrate and water

Lesson 7 – What is hard water?

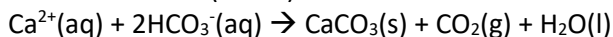
d) Hard and Soft Water

(i) definition of hard water

- Hard water is water which contains a high concentration of dissolved Ca^{2+} , Mg^{2+} and Fe^{2+} ions; these ions are commonly found in rocks and can get into the water supply when water comes into contact with these rocks; the presence of these ions in water causes a number of precipitation reactions to take place which can cause problems; soft water is water which contains only a very low concentration of these ions

(ii) disadvantages of hard water

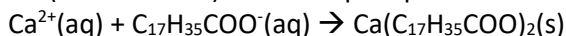
- The first problem caused by the presence of these ions is that when heated, these ions react with dissolved hydrogencarbonate ions (HCO_3^-) in water to form calcium carbonate, which is insoluble:



The compound $\text{CaCO}_3(\text{s})$ is known as limescale and can form on the surface of heating elements, especially in kettles, washing machines, dishwashers and industrial water heaters; limescale reduces the efficiency heat transfer and it can also restrict the flow of water if it builds up in pipes

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- The second problem caused by the presence of these ions is they react with the anions present in soap, such as $C_{17}H_{35}COO^-$ (stearate ions) to form a precipitate:



This precipitate is known as scum and it removes the stearate ions from the water; as a result the lather which should be formed in water by the stearate ions does not form; so soap is less effective in hard water

- The structure of soap will be covered in more detail in Unit 11 – Further Organic Chemistry

(iii) advantages of hard water

- Ca^{2+} , Mg^{2+} and Fe^{2+} are important minerals for the human body, and hard water can be a useful source of these minerals; if the concentration of these ions are very high, however, the disadvantages of hard water are judged to outweigh the benefits

(iv) temporary and permanent hardness

- The hardness of water is described as temporary if the main anion present in the water is HCO_3^- ; when heated, HCO_3^- decomposes into CO_3^{2-} which forms a precipitate with Ca^{2+} , Mg^{2+} and Fe^{2+} ions;



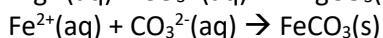
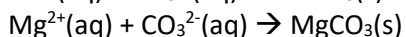
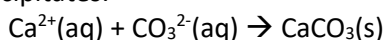
This precipitate can be removed by filtration and the resulting water is soft; temporary hardness can therefore be removed simply by heating the water

- The hardness of water is described as permanent if the main anion present in the water is not HCO_3^- ; in such cases it is usually SO_4^{2-} or Cl^- ; these ions do not form precipitation reactions on heating; permanent hardness therefore cannot be removed just by heating the water

(v) softening water

- Other than simply boiling (which only removes temporary hardness), there are three main techniques available for softening water (ie removing Ca^{2+} , Mg^{2+} and Fe^{2+} ions):

- **Precipitation:** all of the ions responsible for water hardness form insoluble carbonates; the addition of a solution of sodium carbonate (Na_2CO_3) to the water will therefore remove all three of these ions as precipitates:



The resulting water, when filtered, is soft

- **Distillation:** when water is boiled, any dissolved ions do not boil but remain in the residue (as ionic compounds have a high boiling point); when condensed again, the boiled water has lost its ions and is therefore soft
- **Ion Exchange:** ion exchange resins contain weakly attached sodium ions (Na^+); when hard water is passed through these resins, the Ca^{2+} and Mg^{2+} ions attach more strongly to the resin and replace the Na^+ ions; the Ca^{2+} and Mg^{2+} ions in the water are therefore replaced with Na^+ ions and the water is now soft

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS

(vi) testing the hardness of water

- The degree of hardness in different water samples can be compared simply by adding soap solution gradually until a lather is formed; the greater the quantity of soap solution required, the greater the levels of hardness in the water



Practical 7.1: Test the Hardness of Water

Collect four different samples of water from four different sources (such as tap water, rainwater, sea water and bottled water); you will need access to a solution of soap in ethanol

- Using a measuring cylinder, pour 10 cm^3 of the first water sample into a conical flask
- Fill a burette with the solution of soap in ethanol; record the initial burette reading
- Add 1 cm^3 of the soap solution to the conical flask and shake the flask well; observe whether or not a lather forms
- If no lather forms, add a further 1 cm^3 and shake the flask well; continue this process until a lather forms which lasts for 30 seconds
- Once a lather has formed, record the final burette reading and calculate the volume of soap solution required
- Rinse out your conical flask and repeat steps (a) to (e) for the other water samples
- Record your results in a simple table

Water source	Volume of soap required / cm^3

Which water is the hardest?



Extension 7.2: Testing for Temporary and Permanent Hardness in Water

Completed Exercise 7.1? How would you adapt Experiment 7.1 to determine how much of the hardness in each water sample was temporary and how much was permanent?



Test your knowledge 7.3: Understanding the difference between hard and soft water

- Which ions are responsible for the hardness of water?
- State two problems caused by hard water
- State one advantage of hard water
- Describe three ways of reducing the hardness of water

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS

Lesson 8 – How much have I learned about solubility and precipitation reactions?



8.1 END-OF-UNIT QUIZ

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS

1. What is a saturated solution?
2. Explain how, in a saturated solution of NaCl, the solid NaCl and the aqueous NaCl form a dynamic equilibrium
3. Describe how the solubility of most solids varies with temperature
4. Describe how the solubility of most gases varies with temperature
5. Use graph paper to draw a solubility curve for NH₄Cl given the following solubility data:

Temperature / °C	Solubility in water / moldm ⁻³
20	6.7
30	7.5
40	8.2
50	9.2
60	11.4
70	12.7

Use your graph to answer the following questions:

- (a) What is the molarity of a saturated solution of NH₄Cl at 25 °C?
 - (b) What mass of NH₄Cl can be dissolved in 50 cm³ of water at 45 °C?
 - (c) What volume of water is needed to dissolve 10 g of NH₄Cl at 55 °C?
 - (d) At what temperature will a 10.0 moldm⁻³ solution of NH₄Cl start to crystallise if it is cooled slowly from 90 °C?
6. Predict whether the following ionic compounds will be soluble in water:
 - (a) copper sulphate
 - (b) barium sulphate
 - (c) sodium carbonate
 - (d) copper carbonate
 - (e) calcium chloride
 7. Describe what you would observe if you mixed the following pairs of solutions together:
 - (a) copper sulphate and sodium hydroxide
 - (b) calcium nitrate and sodium chloride
 - (c) lead nitrate and sulphuric acid
 - (d) calcium chloride and sodium carbonate
 - (e) silver nitrate and hydrochloric acid
 8. Give a simple chemical test which would distinguish between the following solutions:
 - (a) iron (II) sulphate and iron (III) sulphate
 - (b) lead nitrate and zinc nitrate
 - (c) sodium carbonate and sodium sulphate
 9.
 - (a) Which ions are responsible for the hardness of water
 - (b) Give two reasons why hardness of water is a problem
 - (c) Describe three ways in which water can be softened
 - (d) Describe a simple experiment to compare the hardness of different water samples