UNIT 8

SOLUBILITY AND PRECIPITATION REACTIONS

Answers

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

Summary Activity 1.1: What is a solution?

- Solution: mixture of two components, evenly distributed and in the same phase (usually liquid), solute = minor component of a solution; solvent = major component of a solution
- Water; aqueous solutions
- Sea water, brine, limewater
- Moles of solute per cubic decimetre of solution
- A substance which can form free ions in solution; strong electrolytes completely dissociate into ions in solution, weak electrolytes only partially dissociate

Test your knowledge 1.2: Understanding Solubility and Saturated Solutions

(a) 13.6 $moldm^{-3}$
(b) 61.2 g
(c) Yes, because the molarity would be 12.5 $moldm^{-3}$ which is less than a saturated solution
(d) By heating the solution until some of the water evaporates, or by cooling the solution
(e) $Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$
(f) 0.37 g
(g) No, because the molarity would be 0.27 $moldm^{-3}$ which is greater than the solubility of $Ca(OH)_2$
(h) Glucose is more soluble because its saturated solution has a higher molarity than a saturated solution of calcium hydroxide
(i) Calcium hydroxide is an electrolyte because when it dissolves it dissociates into its ions; glucose it not an electrolyte because it remains as molecules when it dissolves

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

Equipment needed per group: spatula, weighing boat, funnel, filter paper, 250 cm^3 volumetric flask, 250 cm^3 beaker, 25.0 cm^3 pipette with pipette filler, 50 cm^3 burette, clamp, stand, boss, access to mass balance, access to distilled water, access to $Ca(OH)_2$ (1 – 2 g per group), access to phenolphthalein (1 cm^3 per group), access to 0.05 $moldm^{-3}$ HCl (100 cm^3 per group)

- **$Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$**
- **The solubility of Ca(OH) in water is 0.011 $moldm^{-3}$; around 22 cm^3 of HCl should be needed**
- **Using 22 cm^3, moles of HCl = 22/1000 x 0.05 = 0.0011, so moles of $Ca(OH)_2$ = 5.5 x 10^{-4}, so molarity of solution = 2.2 x 10^{-4}/0.05 = 0.011 $moldm^{-3}$; this is the solubility of $Ca(OH)_2$**

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

Summary Activity 2.1: Preparing salts

- We heated the salt solution gently until most of the water had evaporated off
- Some of the salt crystallises out during heating because the concentration of the solution increases as the water is removed
- Most of the salt crystallises out during cooling; the water continues to evaporate so the concentration increases, and the solubility decreases as the solution cools down

Practical 2.2: Purify a sample of copper sulphate by recrystallisation

Equipment needed per group: 15 g of hydrated copper sulphate, spatula, 50 cm^3 measuring cylinder; stirring rod, 250 cm^3 beaker, 3 pieces of filter paper, funnel, evaporating dish, tripod, gauze and Bunsen burner OR sand bath, access to distilled water, access to fridge

- **The insoluble impurities are removed when the solid is filtered**
- **The soluble impurities are removed when the solid is decanted**

Test your knowledge 2.3: Using solubility curves

(a) Approx 9 $moldm^{-3}$

(b) Approx 2.5 $moldm^{-3}$

(c) Solubility = approx. 4 $moldm^{-3}$ so n = 4 x 0.05 = 0.2 so m = 20 g

(d) Solubility = approx. 14 $moldm^{-3}$, n = 20/101 = 0.2 so V = n/C = 0.014 dm^3 or 14 cm^3

(e) n = 10/101 = 0.1 so C = 0.1/0.01 = 10 $moldm^{-3}$ so T = 55 - 57 °C

(f) n = 15/101 = 0.15 so C = 0.15/0.02 = 7.5 $moldm^{-3}$ so T = 46 - 48 °C

(g) n = 30/101 = 0.3 so C = 0.3/0.025 = 12 $moldm^{-3}$ but solubility = 11 $moldm^{-3}$ so not all will dissolve

Extension 2.4: Using solubility curves

Open question so no answers available

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

Summary Activity 3.1: Solubility of Ionic Compounds

- Ionic compounds dissolve in water because the positive ions are attracted to the electronegative O atom in water and the negative ions are attracted to the electropositive H atom in water
- The attraction between the ions and water has to be stronger than the attraction of the ions to each other; in some cases the ions are attracted to each other more strongly than they are attracted to water
- sodium chloride, ammonium sulphate, copper sulphate
- silver chloride, calcium carbonate etc

Test your knowledge 3.2: Predicting the Solubility of Ionic Compounds

(a) $Mg(NO_3)_2$; soluble (Rule 1)
(b) Na_2SO_4; soluble (Rule 2)
(c) $CuCl_2$; soluble (Rule 3)
(d) AgCl; insoluble (Rule 3)
(e) $PbBr_2$; insoluble (Rule 3)
(f) $CuSO_4$; soluble (Rule 4)
(g) $BaSO_4$; insoluble (Rule 4)
(h) $MgSO_4$; soluble (Rule 4)
(i) $BaCO_3$; insoluble (Rule 5)
(j) K_2CO_3; soluble (Rule 2)
(k) $CaCO_3$; insoluble (Rule 5)
(l) $Cu(OH)_2$; insoluble (Rule 5)
(m) LiOH; soluble (Rule 2)
(n) $Ba(OH)_2$; soluble (Rule 5)
(o) $Mg(OH)_2$; insoluble (Rule 5)

Extension 3.3: Predicting the Solubility of Ionic Compounds

Open question so no answers available

Test your knowledge 3.4: Predicting Precipitation

(a) precipitate: $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
(b) precipitate: $Pb^{2+}(aq) + 2Cl^{-}(aq)$ ($PbCl_2(s)$
(c) precipitate: $Cu^{2+}(aq) + 2OH^{-}(aq)$ (Cu(OH)2(s)
(d) no precipitate
(e) precipitate: Ca2+(aq) + CO32-(aq) ($CaCO_3(s)$
(f) no precipitate
(g) precipitate: $Ag^{+}(aq) + Cl^{-}(aq)$ (AgCl(s)
(h) no precipitate

Extension 3.5: Predicting Precipitation

Open question so no answers available

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Lesson 4 – How can we prepare insoluble salts?

Practical 4.1: Observe precipitation reactions

Equipment needed per group: access to labelled bottles containing 0.05 moldm^{-3} $AgNO_3$, and 0.1 moldm^{-3} HCl, H_2SO_4, $BaCl_2$, $CuSO_4$ and NaOH (25 cm^3 per group), with one 10 cm^3 measuring cylinder for each bottle, 15 x test tubes, one test tube rack

- **Expected observations:**

	A	B	C	D	E	F
A		precipitate	precipitate	precipitate	precipitate	precipitate
B			no precipitate	no precipitate	no precipitate	no precipitate
C				precipitate	no precipitate	no precipitate
D					precipitate	no precipitate
E						precipitate

- **Equations:**

AB: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$; AC: $2Ag^+(aq) + SO_4^{2-}(aq) \rightarrow Ag_2SO_4(s)$; AD: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$; AE: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$; AF: $Ag^+(aq) + OH^-(aq) \rightarrow AgOH(s)$; CD: $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$; DE: $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$; EF: $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$

Summary Activity 4.2: Preparation of soluble salts

- Copper sulphate was prepared by reacting excess copper oxide with sulphuric acid; this is an acid-base (or neutralisation) reaction; the excess copper oxide (which is insoluble) was removed by filtration and the soluble salt was extracted by crystallisation (the salt solution was heated and then allowed to cool)
  $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$
- Ammonium sulphate was prepared by reacting ammonia with sulphuric acid in a 2:1 ratio; this is an acid-base (or neutralisation) reaction; both reactants are also soluble so the exact quantities were needed; the soluble salt was extracted by crystallisation (the salt solution was heated and then allowed to cool)
  $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$
- Zinc sulphate was prepared by reacting excess zinc with dilute sulphuric acid; this is a redox reaction; the excess zinc was removed by filtration and the soluble salt was extracted by crystallisation (the salt solution was heated and then allowed to cool
  $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

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

Equipment needed per group: 5 cm^3 of 1 moldm^{-3} lead (II) nitrate solution, 5 cm^3 of 2 moldm^{-3} sodium chloride solution; 2 x 10 cm^3 measuring cylinders, boiling tube with bung; funnel, 2 x filter paper, small beaker (50 cm^3), spatula, access to distilled water

- **$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$**
- **Precipitation**
- **It doesn't matter; whichever reactant is in excess will be removed during the filtration process**

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

Summary Activity 5.1: Qualitative Analysis of Cations

- Add blue litmus paper; it will turn red; add a sample of calcium carbonate; a gas will be given off which turns limewater milky
- Add sodium hydroxide solution and warm; a pungent gas should be given off which turns red litmus paper blue

Practical 5.2: Use precipitation reactions to identify cations in solution

Chemicals needed: minimum 0.1 moldm^{-3} solutions of: $FeSO_4$ (labelled A), Na_2CO_3 (labelled B), $ZnSO_4$ (labelled C), $CaCl_2$ (labelled D), $CuSO_4$ (labelled E), $Pb(NO_3)_2$ (labelled F), $Al_2(SO_4)_3$ (labelled G) and $Fe_2(SO_4)_3$ (labelled H) - around 10 cm^3 per group prepared in a single bottle, each with its own dropping pipette; also 0.5 - 1 moldm^{-3} of the following solutions: NaOH, HCl - up to 20 cm^3 per group; each group needs its own bottle with its own dropping pipette

Apparatus needed per group: 16 test tubes, 1 test tube rack

Expected observations:

Solution	Observations			cation present
	few drops NaOH	excess NaOH	few drops HCl	
A	dark green precipitate	no change	no change	Fe^{2+}
B	no change	no change	no change	Na^+
C	white precipitate	dissolves – colourless solution	no change	Zn^{2+} or Al^{3+}
D	white precipitate	no change	no change	Ca^{2+}
E	pale blue precipitate	no change	no change	Cu^{2+}
F	white precipitate	dissolves – colourless solution	white precipitate	Pb^{2+}
G	white precipitate	dissolves – colourless solution	no change	Zn^{2+} or Al^{3+}
H	orange-brown precipitate	no change	no change	Fe^{3+}

- Zn^{2+} and Al^{3+} cannot be distinguished by this combination of tests

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

Summary Activity 6.1: Qualitative Analysis of Anions

a) Add red litmus paper; it will turn blue; add some ammonium chloride and warm; a pungent gas should be given off which turns red litmus paper blue
b) Add sodium hydroxide solution and aluminium powder and heat; a pungent gas should be given off which turns red litmus paper blue
c) Add HCl(aq); a gas will be given off which turns limewater milky
d) Add HCl(aq); a gas will be given off which turns blue litmus paper red and turns dichromate paper green

Practical 6.2: Use precipitation reactions to identify anions in solution

Chemicals needed: minimum 0.1 moldm^{-3} solutions of: $FeSO_4$ (labelled A), Na_2CO_3 (labelled B), KI (labelled C), KNO_3 (labelled D), $CaCl_2$ (labelled E), Na_2SO_3 (labelled F) - around 10 cm^3 per group prepared in a single bottle, each with its own dropping pipette; also access to 0.05 moldm^{-3} $AgNO_3$, 0.1 moldm $BaCl_2$, 1 moldm^{-3} HCl, 1 moldm^{-3} HNO_3 - up to 50 cm^3 per group; each bottle needs its own dropping pipette
Apparatus needed per group: 17 test tubes, 2 test tube racks
Expected observations:

Solution	Observations			anion present
	HNO_3 and $AgNO_3$	$BaCl_2$	HCl and $BaCl_2$	
A	white precipitate	white precipitate	white precipitate	SO_4^{2-}
B	no change	white precipitate	no change	CO_3^{2-} or SO_3^{2-}
C	yellow precipitate	no change	-	I^-
D	no change	no change	-	NO_3^-
E	white precipitate	no change	-	Cl^-
F	no change	white precipitate	no change	CO_3^{2-} or SO_3^{2-}

CO_3^{2-} and SO_3^{2-} cannot be distinguished by this combination of tests; they could be distinguished by adding $CaCl_2$(aq) and then adding HCl(aq) to the resulting precipitate; the gas evolved from CO_3^{2-} will turn limewater milky; the gas evolved from SO_3^{2-} will turn blue litmus red and turn dichromate paper green

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

(a) add NaOH (aq); no reaction with NaCl; white precipitate with Ca(OH)2
(b) add NaOH (aq) dropwise and then in excess; white precipitate with Ca(NO3)2 is insoluble in excess NaOH; white precipitate with Pb(NO3)2 dissolves in excess NaOH
(c) add NaOH (aq); dark green precipitate with FeSO4; pale blue precipitate with CuSO4: orange/brown precipitate with Fe2(SO4)3
(d) add HCl (aq); no reaction with Ca(NO3)2; white precipitate with Pb(NO3)2
(e) Add AgNO3 (aq); no reaction with NaNO3; white precipitate with NaCl
(f) Add BaCl2 (aq); no reaction with NaCl; white precipitate with Na2SO4
(g) Add BaCl2 (aq) then add HCl(aq); white precipitate with Na2SO4 is insoluble in HCl; white precipitate with Na2CO3 dissolves in HCl

Extension 6.4: Further qualitative analysis

(a) add blue litmus paper; it turns red in nitric acid but not in sodium nitrate
(b) add HCl; observe bubbles with sodium carbonate but not with sodium hydroxide, or add magnesium chloride solution (or any solution containing a +2 ion); a precipitate forms in both cases; with the carbonate, the precipitate will give off bubbles when it dissolves, but the hydroxide will dissolve without giving off bubbles
(c) Add NaOH and heat; with ammonium nitrate, a pungent gas will be given off which turns red litmus blue; with sodium nitrate there will be no reaction
(d) Add aluminium powder and sodium hydroxide and hear; with sodium nitrate, a pungent gas will be given off which turns red litmus blue; with water there will be no reaction

Lesson 7 – What is hard water?

Practical 7.1: Test the Hardness of Water

Chemicals needed: water from different sources and a solution of soap in ethanol (around 50 cm^3 per group)
Apparatus needed per group: one conical flask, one measuring cylinder (10 cm^3), one burette with clamp, boss and stand and one funnel
The seawater should be the hardest (need the greatest quantity of soap) and the rainwater should be the softest (need the least quantity of soap)

Extension 7.2: Testing for Temporary and Permanent Hardness in Water

Take 10 cm^3 of water from each source and boil them before adding the soap; then add the soap as in the original experiment; if less soap is required with the boiled sample, some of its hardness is temporary; the bigger the difference, the greater the amount of temporary hardness in the water

Test your knowledge 7.3: Understanding the Difference Between Hard and Soft Water

(a) Ca^{2+}, Mg^{2+} and Fe^{2+} ions
(b) Forms limescale when heated, forms scum instead of lather with soap
(c) It is a good source of minerals for humans
(d) Distillation, ion exchange, precipitation using sodium carbonate

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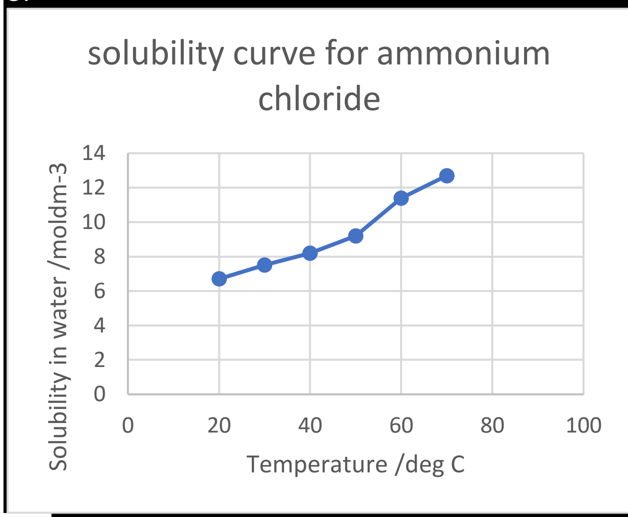
Lesson 8 – How much have I learned about solubility and precipitation reactions?

8.1 END-OF-UNIT QUIZ
UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS

1. A solution which contains the maximum amount of dissolved solute which it is possible to dissolve in that quantity of solvent
2. The solid NaCl dissolves and the aqueous NaCl crystallises at equal rates
3. Solubility of solids usually increases with temperature
4. Solubility of gases usually decreases with temperature
5.



(a) 7.1 – 7.2 moldm-3 (b) solubility = 8.5 – 8.7 moldm^{-3} so mass = 23 – 24 g
(c) solubility = 10.0 – 10.2 moldm^{-3} so volume = 18 – 19 cm^3 (d) around 54 °C

6. (a) soluble; (b) insoluble; (c) soluble; (d) insoluble; (e) soluble
7. (a) pale blue precipitate; (b) no reaction; (c) white precipitate; (d) white precipitate; (e) white precipitate
8. (a) Add NaOH (aq); $FeSO_4$ gives dark green precipitate, $Fe_2(SO_4)_3$ gives orange/brown precipitate
(b) Add HCl (aq); $Pb(NO_3)_2$ gives white precipitate; $Zn(NO_3)_2$ gives no reaction
(c) Add HCl (aq) and then $BaCl_2$(aq); Na_2CO_3 gives no reaction; Na_2SO_4 gives white precipitate
9. (a) Fe^{2+}, Ca^{2+}, Mg^{2+}
(b) causes limescale when heated, causes scum instead of lather with soap
(c) ion exchange, distillation, precipitation with sodium carbonate
(d) Add soap from a burette to a fixed quantity of different water sample; measure how much is needed to form a lather; the more soap needed, the harder the water