SECTION 5

SOLUBILITY AND ITS APPLICATIONS IN QUALITATIVE ANALYSIS

PHYSICAL CHEMISTRY

Equilibria

INTRODUCTION

Solubility tells us how much of a substance can dissolve in a liquid, like sugar in water. It helps explain why some things, like salt, mix well in water, while others, like oil, do not. We can also use solubility rules to identify which ions are in a solution by seeing if a compound dissolves or forms a solid when mixed with water.

At the end of this section, you will be able to:

1. Explain the term solubility and describe the factors that affect the solubility of substances.
2. Perform tests on water-soluble compounds to identify ions based on the solubility rules.

Key Ideas

- **Solubility** is the ability of a substance (solute) to dissolve in a solvent to form a homogeneous mixture or solution.
- **Precipitate**: is an insoluble solid that forms when two solutions are mixed, and a chemical reaction occurs.
- **Solvent** is the substance in which a solute is dissolved to form a solution, typically a liquid like water.
- **Solute** is the substance that is dissolved in a solvent to form a solution.
- **Aqueous solution** is a solution in which water is the solvent.
- **Insoluble** is a term used to describe a substance that does not dissolve in a particular solvent.
- **Saturation** is the point at which a solution can no longer dissolve any more solute at a given temperature and pressure.

- **Qualitative analysis:** is a process used to determine the presence of certain cations and anions in a substance by observing chemical reactions and their outcomes.
- **Spectator ions** are ions in a solution that do not participate directly in a chemical reaction but remain unchanged.
- **Effervescence:** is the release of gas bubbles during a reaction, often seen when acids react with carbonates.
- **Equilibrium** is the state in which the forward and reverse reactions occur at the same rate, so the concentrations of reactants and products remain constant.

SOLUBILITY

When a solution is formed from substances in different states of matter, the solvent is the substance that retains its state of matter. This implies that the solute changes its state. The constituent of the mixture that is present in a smaller amount is called the solute and the one present in a larger quantity is called the solvent. Demonstrate using table salt (NaCl) and water to prepare a salt solution.

Types of Solutions Based on States of Constituent

Solutions can be prepared in any of the three states of matter (solid, liquid, and gas). Even though the most common type of solution that is seen in our immediate environment is where the solute is solid and the solvent is in the liquid state (for example a salt or sugar solution), there are different types of solutions in nature. Any of the three states of matter can behave as a solute or a solvent depending on its quantity in the solution.

Table 5.1: Types of solutions and their constituents

Type of Solution	State of solute	State of solvent	Example
Solid - liquid	solid	liquid	Salt solution
Liquid - liquid	liquid	liquid	Alcohol in water
Gas - liquid	Gas	Liquid	dew
Gas - gas	Gas	Gas	Air
Gas - solid	Gas	Solid	H_2 with Platinum
Solid - solid	Solid	Solid	Alloys

Based on the extent of solute dissolution in a solvent at a given temperature, a solution can be categorised as **unsaturated, saturated or supersaturated**. An **unsaturated** solution is a solution that can dissolve more solute at a given temperature. A **saturated** solution is a solution that cannot dissolve any more solute at a given temperature. A **supersaturated** solution contains more solute in a solution than the maximum amount it can dissolve at a given temperature in the presence of undissolved solutes.

Demonstrate Unsaturated and Saturated Solutions Using a Named Salt

1. Pour about 100 cm^3 of water into a clean dry beaker.
2. Use a spatula to transfer some salt into the water in the beaker.
3. Stir with a rod until all dissolves.
4. Describe the type of solution formed.
5. Transfer more solute and stir to dissolve.
6. As more solute is added, observe and describe what happens.
7. Filter any solute that may be left and collect the filtrate in a beaker.
8. Describe the type of solution formed.
9. Transfer more solute into the filtrate and stir.
10. As more solute is added, observe and describe what happens.

The solubility of a solute at a given temperature is the maximum quantity of the solute in moles or grammes that dissolves in one (1) cubic decimetre of a solvent to form a saturated solution.

Factors that Affect Solubility

Nature of the solvent concerning the solute: Polar solvents have molecules with permanent dipoles, making them interact with polar solutes while non-polar solutes lack permanent dipoles and interact with non-polar solutes. When two substances are similar in terms of polarity, they are more likely to dissolve in each other. Polar solutes dissolve in polar solvents while non-polar solutes dissolve in non-polar solvents. Ionic compounds are more soluble in polar solvents than non-polar solvents. Inorganic acids, bases and ionic salts tend to be much more soluble in water.

Temperature: Generally, solubility increases with the temperature rise and decreases with the fall of temperature, although this is not always the case. Heat is needed to break the bonds holding the molecules or ionic lattices in the solid together. At the same time, the formation of new solute-solvent bonds gives off an amount of heat.

Pressure: Pressure can affect the solubility of a gas (solute) in a liquid solvent but has very little effect when the solute is a liquid or solid sample. Henry's Law states that the amount of gas that is dissolved in a liquid is directly proportional to the partial pressure of that gas.

The activities below will help you master hands-on exploration, reflection, and application of knowledge, fostering a deeper understanding of solutions and their properties.

Activity 5.1: Solutions Around Us

Part A: Observing and Identifying Solutions

1. Before starting, write down what you understand by the terms *solute*, *solvent*, and *solution*.
2. ***Preparation of Solutions:***
    - **a.** Sugar Solution: Dissolve a teaspoon of sugar in a glass of water. Stir it well.
    - **b.** Kerosene and Water: Pour a small amount of kerosene into a glass of water.

    Observe what happens.

    - **c.** **Coke**: Pour Coke into a clear glass. Observe the bubbles.
    - **d.** Write down your observations for each of the solutions above:
        - i. What happens to the sugar when it is added to water?
        - ii. How do the kerosene and water behave together?
        - iii. What do you notice about the Coke? How do the bubbles behave?

Part B: Categorisation of Solutions:

1. Based on your observations, classify each example (sugar solution, kerosene and water, Coke) into one of the following categories:
    - **a. Solid-Liquid Solution**
    - **b. Liquid-Liquid Solution**
    - **c. Gas-Liquid Solution**

2. For each type of solution, explain which substance is the solute and which is the solvent.
3. Provide examples of similar solutions from your everyday life.
4. Try to identify other types of solutions based on the materials available around you:
   a. **Solid-gas solution**: Can you think of an example?
   b. **Solid-solid Solution**: Do you have an example at home?
   c. **Gas-gas solution**: Where might you encounter this?

Reflection and Conclusion:

1. Summarise your findings and write a short paragraph explaining what you have learnt about solutions.
2. Reflect on how this knowledge might apply in real-world situations or other subjects.
3. If possible, share your findings with peers or family members and discuss whether they agree with your classifications.
4. Consider researching a real-world application of each type of solution you identified.

Activity 5.2: Hands-On Preparation of Solutions

Materials needed:

- Clear plastic beaker or glass jar
- Stirring rod or spoon
- Graduated cylinder or measuring cup
- Solutes: Sugar, salt (or other suitable solutes)
- Solvents: Water
- Hot water (for making supersaturated solutions)
- A balance (for measuring the mass of the solutes)
- Temperature-controlled hot plate (for heating water, if preparing supersaturated solutions)
- Thermometer (to monitor the temperature of hot water)
- Ice (to cool supersaturated solutions)

Procedure

1. **Unsaturated Solution:**
   - **a.** Fill a beaker or jar with water.
   - **b.** Add a small amount of sugar or salt to the water and stir until it dissolves completely. If needed, add more solute to see if it still dissolves.
   - **c.** Observe the solution and confirm its nature.

2. **Saturated Solution:**
   - **a.** Gradually add more sugar or salt to the same beaker, stirring continuously until no more solute will dissolve.
   - **b.** Look for undissolved solute at the bottom of the beaker, indicating a saturated solution. Reflect on the appearance and properties of the solution.

3. **Supersaturated Solution:**
   - **a.** *Preparation:* Heat water to around 60°C and dissolve as much sugar or salt as possible.
   - **b.** *Cooling:* Allow the hot solution to cool slowly at room temperature or in an ice bath. Be cautious, as supersaturated solutions can be delicate.
   - **c.** *Observation:* Once the solution is cooled, try adding a seed crystal or gently disturbing it. What do you observe, and what scientific process is explored?
   - **d.** *Observation and Discussion:* Take note of the appearance and behaviour of each type of solution. Reflect on the differences you observe based on the questions below.
     - i. What did you notice in each type of solution?
     - ii. How did the solubility change with temperature in the case of the supersaturated solution?
     - iii. Why do you think the solute behaves differently in each solution?
     - iv. Consider how these observations relate to the concepts of solubility and saturation.

Activity 5.3: Investigating Factors Affecting Solubility

In this activity, you will explore the factors that affect solubility, specifically *temperature, pressure, and the nature of solute and solvent*. You will get to understand how temperature, pressure, and the nature of solute and solvent affect the solubility of a substance.

Materials Needed:

- Beakers or clear cups
- Hot water, cold water
- Table salt (NaCl), Sugar
- Carbonated water (fizzy drinks)
- Different solvents (examples; water, oil)
- Different solutes (examples; sugar, salt, sand)
- Thermometer, Stirring rods
- Pressure container (if available for demonstration)
- Stopwatch or timer

Procedure:

(You are to work in smaller groups. Each group will investigate one or more factors affecting solubility)

Experiment 1: Investigate How Temperature Affects the Solubility of Solids in Liquids (Sugar in water)

1. Fill two beakers with equal volumes of water. One beaker with hot water and the other with cold water.
2. Add an equal mass of sugar gently to each beaker at the same time and stir.
3. Observe and record the time taken for the sugar to dissolve completely in both hot and cold water.

NB: Take particular care when using hot sugar solutions as they can cause nasty burns if spilled on the skin.

Experiment 2: Explore How Pressure Affects the Solubility of Gases in Liquids. (Carbon dioxide in water)

1. Open a bottle of carbonated water (fizzy drink) at room temperature and observe the release of carbon dioxide gas.
2. Discuss how carbon dioxide is more soluble under higher pressure, and how the release of pressure (opening the bottle) causes the gas to escape.

Experiment 3: Determine How the Nature of a Solute and Solvent affects the solubility of the substance

1. Add equal amounts of table salt (ionic) and sugar (molecular) to separate beakers of water and oil.
2. Stir and observe which solute dissolves in which solvent.

Questions:

1. Define solubility.
2. Define a saturated solution.
3. Explain each of the following terms;
   a. Unsaturated solution
   b. Saturated solution
   c. Supersaturated solution
4. Explain how temperature affects the solubility of gases in liquids

SOLUBILITY RULES

Solubility rules are guidelines that are used to predict whether certain compounds will dissolve in water to form a solution. These rules are crucial in chemistry, particularly in analytical chemistry and environmental science, as they help identify ions in water-soluble compounds and predict the outcomes of chemical reactions.

Soluble Salts

1. The Na^+, K^+, and NH_4^+ ions form soluble salts. Thus, NaCl, KNO_3, $(NH_4)_2SO_4$, Na_2S, and $(NH_4)_2CO_3$ are soluble.
2. The nitrate (NO_3^-) ion forms soluble salts. Thus, $Cu(NO_3)_2$ and $Fe(NO_3)_3$ are soluble.

3. The chloride (Cl^-), bromide (Br^-), and iodide (I^-) ions generally form soluble salts. Exceptions to this rule include salts of the Pb^{2+}, Hg^{2+}, Ag^+, and Cu^+ ions. $ZnCl_2$ is soluble, but CuBr is not.
4. The sulphate (SO_4^{2-}) ion generally forms soluble salts. Exceptions include $BaSO_4$, $SrSO_4$ and $PbSO_4$, which are insoluble, and Ag_2SO_4, $CaSO_4$ and Hg_2SO_4, which are slightly soluble

Insoluble Salts

1. Sulphides (S^{2-}) are usually insoluble. Exceptions include Na_2S, K_2S, $(NH_4)_2S$, MgS, CaS, SrS, and BaS.
2. Oxides (O^{2-}) are usually insoluble. Exceptions include Na_2O, K_2O, SrO and BaO, which are soluble, and CaO, which is slightly soluble.
3. Hydroxides (OH^-) are usually insoluble. Exceptions include NaOH, KOH, $Sr(OH)_2$, and $Ba(OH)_2$, which are soluble, and $Ca(OH)_2$, which is slightly soluble.
4. Chromates (CrO_4^{2-}) are usually insoluble. Exceptions include Na_2CrO_4, K_2CrO_4, $(NH_4)_2CrO_4$, and $MgCrO_4$.
5. Phosphates (PO_4^{3-}) and carbonates (CO_3^{2-}) are usually insoluble. Exceptions include salts of the Na^+, K^+, and NH_4^+ ions.

Determination of the solubility of soluble substances

1. Put distilled water into a clean beaker and heat gently above room temperature.
2. Add the named salt a little at a time while stirring until a saturated solution is formed.
3. Cool to room temperature
4. A known volume of the saturated solution is transferred into a previously weighed clean and dry evaporating dish.
5. Weigh the evaporating dish and the solution.
6. Evaporate to dryness using a water bath.
7. Allow to cool and then weigh.
8. Repeat the drying and cooling process to ensure a constant mass of the salt is obtained.
9. Determine the number of moles of the salt in the volume of solution that was evaporated.

Importance of Solubility Rules

Understanding solubility rules is essential because they allow us to:

1. *Predict Reaction Products:* Knowing which compounds will dissolve helps chemists predict the products of reactions, especially double displacement reactions.
2. *Separate Mixtures:* In analytical chemistry, solubility rules aid in separating and identifying different ions in a mixture through precipitation reactions.
3. *Manage Environmental Impacts:* Environmental scientists use solubility rules to understand how pollutants behave in water and to design strategies for water treatment.

Solubility Curve

This is a graph that shows how temperature affects the solubility of salts. It is plotted with the solubility on the y-axis and temperature on the x-axis. Each point on the graph represents the maximum amount of solute that can dissolve in a solvent at a particular temperature. A change in temperature corresponds with a change in the maximum amount of the solute that can dissolve a solvent. The various points on the graph are then joined to form a curve better known as the solubility curve. The following deductions can then be made from the curve:

1. Any point below the curve represents an unsaturated solution at that particular temperature.
2. At any particular temperature, the solubility of the salt can be traced vertically to the curve.
3. Any point above the curve represents a supersaturated solution at that particular temperature.
4. To determine how much extra salt has been dissolved or crystallised out, compute the difference in the corresponding y-axis values at the respective temperatures.
5. Determine the point at which crystallisation starts and be able to calculate the mass of solute that will be crystallised out when a solution is cooled from one temperature to the other by simple subtraction on the y-axis axis.

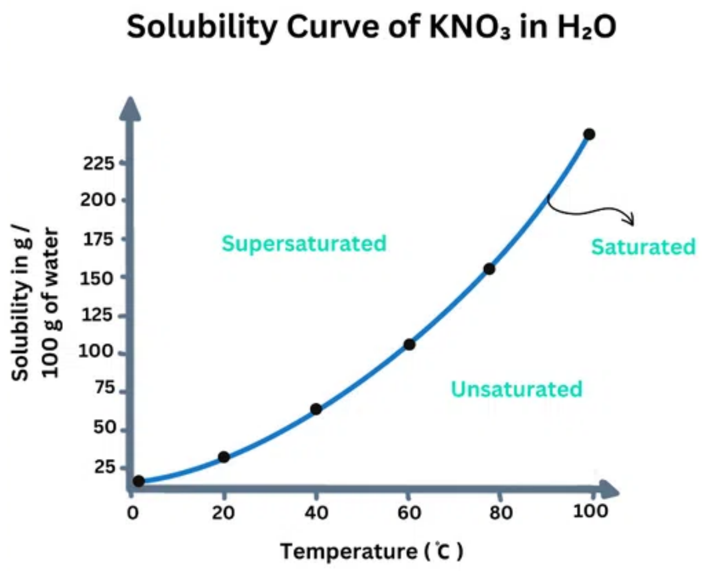


Figure 5.1: Sample of a solubility curve graph (showing areas of unsaturation, saturation and supersaturation solutions)

Questions:

1. Study the graph below and explain it to a colleague

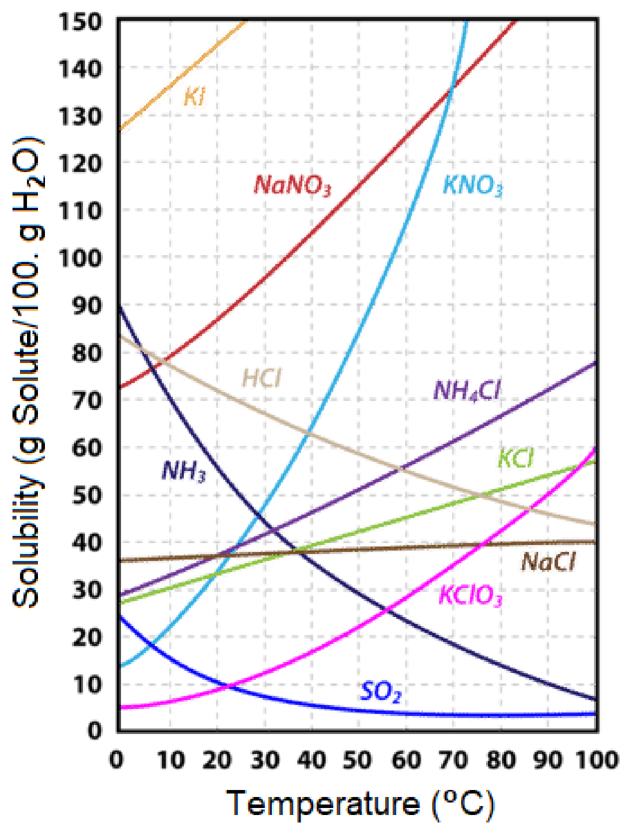


Figure 5.2: Solubility curves of several solid and gaseous solutes. **Image source:** https://chem.libretexts.org/@api/deki/files/275860/Solubility_Curves.png?revision=1

2. Use the graph below to determine the solubility of a specific salt at different temperatures in the solubility curves.

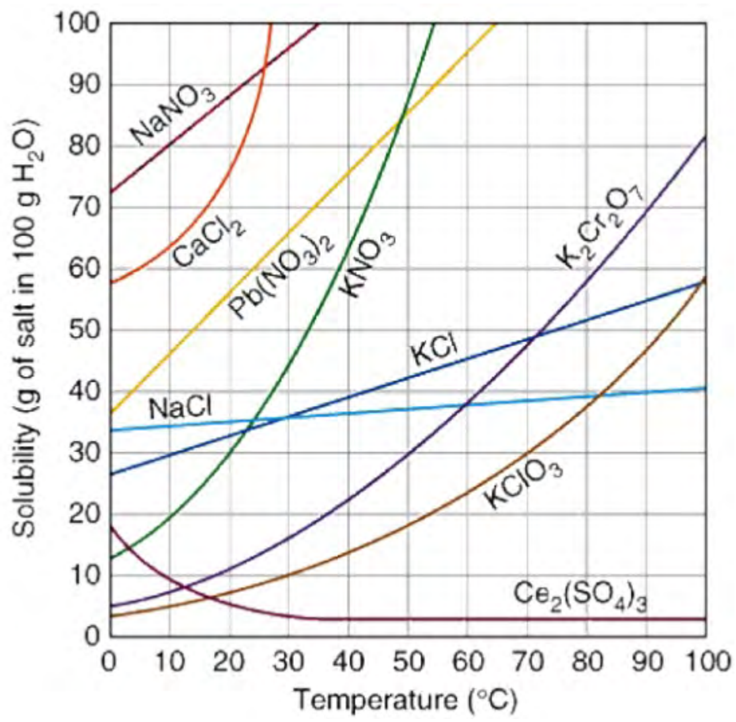


Figure 5.3: Solubility of various salts

a. What does the point above a solubility curve represent?

b. Use the information below to answer the questions that follow;

Temperature/°C	Solubility of NaCl (g/100g water)
0	33.0
20	33.5
40	34.0
60	34.5
80	35.0

i. Plot the solubility curve for NaCl

ii. What is the solubility of NaCl at 30°C?

3. The solubility values of Potassium chlorate at different temperatures are given below:

Temperature/°C	Solubility/(g/dm³)
0	5.3
10	5.9
20	6.5
30	7.0
40	7.5
50	8.0
60	8.5
70	9.0

a. Plot the solubility curve for $KClO_3$ from the result above.

b. From the graph, determine the solubility of $KClO_3$ at 35°C

QUALITATIVE CHEMICAL ANALYSIS

In chemistry, qualitative analysis is the process of identifying which elements or groups of elements are present in a sample. This branch of chemistry focuses on finding out what is in the sample, rather than measuring how much of each substance is there. The methods used can vary in complexity depending on the sample and how it interacts with other substances.

Preliminary Examination

The first step in qualitative analysis is often a preliminary examination. By observing the appearance of the substance and testing its reaction with litmus paper, you can get clues about which ions might be present. Checking the solubility of the sample is also part of this initial examination. If the sample is insoluble, noting the colour and nature of any precipitate (solid formed in the solution) and its behaviour when more of the substance is added can provide important information.

Examination of Cations

To systematically identify the cations (positively charged ions) in a sample, chemists use a series of steps to precipitate groups of cations from the solution using specific reagents. Each group of precipitates is then examined further to identify the individual cations. Common reagents for this purpose include aqueous solutions of sodium hydroxide – $NaOH_{(aq)}$ and ammonia - $NH_{3(aq)}$.The effects of these reagents on different cations can help to identify which cations are present in the sample.The table below describes the effect of **$NaOH_{(aq)}$** on cations in ***drops*** and then in ***excess***.

Table 5.2: Effect of **$NaOH_{(aq)}$** on cations

Cation	Drops of $NaO_{H(aq)}$	Excess $NaOH_{(aq)}$
Ca^{2+}	White chalky precipitate	precipitate insoluble
Pb^{2+}	White chalky precipitate	precipitate soluble
Al^{3+}	White gelatinous precipitate	precipitate soluble
Zn^{2+}	White gelatinous precipitate	precipitate soluble
Cu^{2+}	Blue gelatinous precipitate	precipitate insoluble
Fe^{2+}	Dark green gelatinous precipitate	precipitate insoluble
Fe^{3+}	Reddish brown precipitate	precipitate insoluble
NH_4^+	No visible reaction	No visible reaction

To identify cations, the test with an aqueous solution of NaOH should be repeated with an aqueous ammonia solution (also known as ammonium hydroxide, NH_4OH). The observations with ammonia solution are usually similar to those of NaOH, but not always the same and the differences can be as important as identifying metal ions. The table below describes the effect of $NH_{3\ (aq)}$ on cations in ***drops*** and then in ***excess.***

Table 5.3: Effect of **$NH_{3(aq))}$** on cations

Cations	Drops of $NH_{3(aq)}$	Excess $NH_{3(aq)}$
Ca^{2+}	No visible reaction	No precipitate formed
Pb^{2+}	White chalky precipitate	precipitate insoluble

Cations	Drops of $NH_{3(aq)}$	Excess $NH_{3(aq)}$
Al^{3+}	White gelatinous precipitate	precipitate insoluble
Zn^{2+}	White gelatinous precipitate	precipitate soluble
Cu^{2+}	Blue gelatinous precipitate	precipitate dissolves
Fe^{2+}	Dark green gelatinous precipitate	precipitate insoluble
Fe^{3+}	Reddish brown precipitate	precipitate insoluble

Examination of Anions

The procedure for detecting anions (negatively charge ions) in aqueous solutions is called anion analysis. Some preliminary tests are done before anion analysis. The preliminary tests that are done include physical examination (i.e. colour and smell), effect on litmus paper and dry heating test (for the evolution of gas).

The other preliminary tests are based on the fact that CO_3^{2-}, S^{2-}, NO_3^- and SO_3^{2-} react with dilute H_2SO_4 to produce CO_2, H_2S, NO_2 and SO_2 gases respectively. These gases on identification indicate the nature of the anion present in the salt.

Cl^-, Br^-, I^- and NO_3^- react with concentrated H_2SO_4 but not with diluted H_2SO_4 to produce Cl_2, Br_2, I_2 and NO_2 gases respectively. SO_4^{2-} does not react with diluted H_2SO_4 or concentrated H_2SO_4. The table below describes the ***smell, colour*** and ***effect*** of the various gases on wet litmus paper.

Table 5.4: Smell, colour and effect of gases on wet litmus paper.

Gas	Smell	Colour	Litmus paper
SO_2	Irritating, Pungent	Colourless	Turns from blue to red
NO_2	Irritating, Pungent	Reddish-brown	Turns from blue to red
H_2S	Rotten egg	Colourless	Turns from blue to Faintly red
CO_2	Odourless	Colourless	Turns from blue to Faintly red
HCl	Choking	Colourless	Turns from blue to red
HBr	Choking	Colourless	Turns from blue to red
HI	Choking	Colourless	Turns from blue to red

Anions that give white precipitation

The following bench solutions produce white precipitate with anions -***$BaCl_2$, $AgNO_3$, $Ca\ (OH)_2$ and $Pb(NO_3)_2$***

Actions of $BaCl_2$ on some anions

This is a precipitation reaction caused by barium ions and sulphate, sulphite and carbonate ions clumping together to give a white solid mass (precipitate).

$Ba^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)}\ BaSO_{4(s)}$

$BaSO_4$ is insoluble in dilute HCl.

$Ba^{2+}{}_{(aq)} + SO_3{}^{2-}{}_{(aq)}\ BaSO_3(s)$

$BaSO_3$ dissolves when dilute HCl is added to produce a soluble $BaCl_2$, with the evolution of a colourless gas that has a choking smell. The gas, SO_2, produced turns the purple colour of acidified $KMnO_4$ to colourless.

$Ba^{2+}{}_{(aq)} + CO_3{}^{2-}(aq)\ BaCO_3(s)$

$BaCO_3$ dissolves when dilute HCl is added to produce a soluble $BaCl_2$ with the evolution of a colourless odourless gas. The gas CO_2 produced turns lime water milky.

$BaSO_4$, $BaSO_3$, and $BaCO_3$ all produce a white precipitate upon reacting with $BaCl_{2.}$ However, each of the resulting products has a peculiar reaction with dilute HCl.

Actions of $AgNO_3$ on some anions

This is a precipitation reaction caused by silver ions with chloride, sulphite and carbonate ions clumping together to give a white solid mass (precipitate).

$Ag^{+}{}_{(aq)} + Cl^{-}{}_{(aq)}\ AgCl(s)$

$AgCl_{(s)}$ is insoluble in dilute HNO_3 but becomes soluble when excess NH_3 is added to the resulting insoluble precipitate.

$Ag^{+}{}_{(aq)} + SO_3{}^{2-}{}_{(aq)}\ Ag_2SO_{3(s)}$

$Ag_2SO_{3(s)}$ dissolves when dilute HNO_3 is added to produce a soluble $AgNO_3$, with the evolution of a colourless gas that has a choking smell.

The gas, SO_2, produced turns the purple colour of acidified $KMnO_4$ to colourless.

$Ag^{+}{}_{(aq)} + CO_3{}^{2-}{}_{(aq)}\ Ag_2\ CO_{3(s)}$

$Ag_2CO_{3(s)}$ dissolves when dilute HNO_3 is added to produce a soluble $AgNO_3$, with the evolution of a colourless odourless gas. The gas CO_2 produced turns lime water milky.

Actions of $Ca(OH)_2$ on some anions

$Ca(OH)_2$ reacts with sulphites and carbonates to give calcium $CaSO_3$ and $CaCO_3$, which are seen as a white precipitate. Further tests or reactions with dilute hydrochloric acid help to evolve characteristic gases.

$CaSO_3$ dissolves when dilute HCl is added to produce a soluble $CaCl_2$, with the evolution of a colourless gas that has a choking smell. The gas, SO_2, produced turns the purple colour of acidified $KMnO_4$ to colourless.

$CaCO_3$ dissolves when dilute HCl is added to produce a soluble $CaCl_2$, with the evolution of a colourless odourless gas, CO_2, which turns lime water milky.

Actions of $Pb(NO_3)_2$ on some anions

Lead (II) nitrate, $Pb(NO_3)_2$, can have various effects on different anions, depending on their chemical properties. It is often used in qualitative analysis to test for the presence of specific anions. Here are some common effects of lead (II) nitrate on certain anions such as halogens, sulphate sulphites and carbonate. Regarding the halogens, different colours of precipitates are formed with respect to the type of the halogen being tested.

$Pb(NO_3)_2 + 2Cl^-_{(aq)}\ PbCl_{2\,(s)} + 2NO_3^-$

Lead (II) nitrate reacts with chloride ions to form a white precipitate of lead chloride ($PbCl_2$).

$Pb(NO_3)_2 + 2I^-_{(aq)}\ PbI_{2(s)} + 2NO_3^-$

Lead (II) nitrate reacts with iodide ions to form a yellow precipitate of lead iodide (PbI_2).

$Pb(NO_3)_2 + 2Br^-(aq)\ PbBr_2(s) + 2NO_3^-$

Lead (II) nitrate reacts with bromide ions to form a creamy-white precipitate of lead bromide ($PbBr_2$).

$SO_4^{2-}{}_{(aq)} + Pb(NO_3)_{2(aq)}\ PbSO_{4(s)} + 2NO_3^-$Lead(II) nitrate reacts with sulphate ions to form a white precipitate of lead sulphate ($PbSO_4$).

$Pb(NO_3)_{2(aq)} + CO_3^{2-}{}_{(aq)}\ PbCO_{3(s)} + 2NO_3^-{}_{(aq)}$

Lead (II) nitrate reacts with carbonate ions to form a white precipitate of lead (II) carbonate ($PbCO_3$), which is slightly soluble in water.

It is essential to remember that the formation of these precipitates indicates the presence of the corresponding anions in the solution. Confirmatory tests or additional analyses may be necessary to verify the presence of specific anions, especially if multiple anions produce similar precipitates. Additionally, some anions might require specific pre-treatments or adjustments in pH before performing the test to ensure accurate results.

Self Assessment

1. What are the general solubility rules for common ions in aqueous solutions?
2. Imagine you have a solution containing an unknown compound. Describe a step-by-step procedure to identify the ions present in the compound using solubility rules. Include the tests you would perform and how you would interpret the results.
3. A student has conducted a series of tests on a water-soluble compound and determined that it contains either chloride ions (Cl^-) or carbonate ions (CO_3^{2-}). Design an experimental procedure to identify which ion is present in the compound. Explain your rationale for each procedure step and how it will lead to a conclusive result.

REVIEW QUESTIONS

1. Imagine you are a scientist working in a laboratory to develop a new carbonated beverage. You need to understand the solubility of different gases and how they behave in the liquid solution under varying conditions.
   a. Before you begin your experiments, explain what solubility is in the context of the gases dissolving in the liquid base of the beverage.
   b. During the preparation, you reach a point where no more gas can dissolve in the liquid at a given temperature and pressure. Explain what you think may happen.
   c.
      i. Describe what an unsaturated solution would look like in your beverage preparation.
      ii. Revisit the concept of a saturated solution and explain how it applies when preparing the beverage with a specific gas concentration.
      iii. Imagine you cool the beverage rapidly after dissolving more gas than usual. Explain what a supersaturated solution is in this context and the potential consequences.
      d. As you experiment with different temperatures, explain how temperature changes would affect the amount of gas that can dissolve in the beverage. How might this influence the final product's quality and consistency?
2. Use the data below to answer the questions that follow;

Solubility of NaCl g/100g water	33.0	33.5	34.0	34.5	35.0
Temperature /°C	0	20	40	60	80

   a. Plot the solubility curve for NaCl.
   b. What does a point above a solubility curve represent?
   c. What is the solubility of NaCl at 30°C?

3. The solubility values of Potassium chlorate – $KClO_3$ at different temperatures are given below: Use the data or information given to answer the questions that follow.

Solubility g/d^{m}3	5.3	5.9	6.5	7.0	7.5	8.0	8.5	9.0
Temperature /^{o}C	0	10	20	30	40	50	60	70

a. Plot the solubility curve for $KClO_3$.

b. What does a point on the solubility curve represent?

c. What is the solubility of $KClO_3$ at 35^{o}C.?

d. What will be the temperature if the solubility is 8.75g/dm^3?

4. A student has conducted a series of tests on a water-soluble compound and determined that it contains either chloride ions (Cl^-) or carbonate ions (CO_3^{2-}). Design an experimental procedure to identify which ion is present in the compound. Explain your rationale for each procedure step and how it will lead to a conclusive result.

5. Imagine you have a solution containing an unknown compound. Describe a step-by-step procedure to identify the ions present in the compound using solubility rules. Include the tests you would perform and how you would interpret the results.

ANSWERS TO REVIEW QUESTIONS

1. a. **Solubility** refers to the ability of a gas to dissolve in a liquid. In the case of this carbonated beverage, it describes how much of the gas can dissolve in the liquid base, creating the fizz and bubbles characteristic of carbonated drinks.

   b. At this point, the solution is saturated, meaning the liquid has dissolved as much gas as it can under the given temperature and pressure conditions.

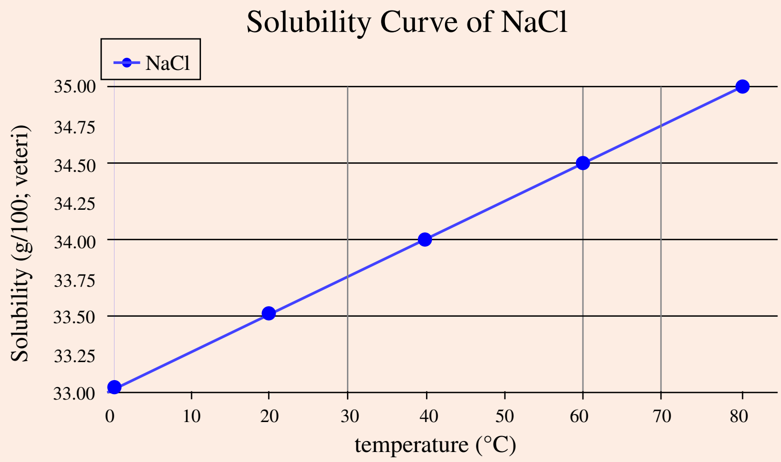
   c. i. An unsaturated solution in your beverage would mean that the liquid can still dissolve more gas. The liquid has not yet reached its capacity for gas solubility.

      ii. A saturated solution occurs when the liquid has dissolved the maximum amount of gas possible at the current temperature and pressure. In your beverage, this means no more gas can dissolve, and any additional gas would not integrate into the liquid.

      iii. A supersaturated solution happens when the liquid holds more dissolved gas than it normally would at a given temperature. This can occur if you rapidly cool the beverage after dissolving an excess amount of gas. This state is unstable, and the excess gas may eventually escape, leading to fizzing or even the liquid bubbling over.

   d. **Effect of temperature on gas solubility**: As the temperature increases, the solubility of gases in the liquid typically decreases, meaning less gas can dissolve in the beverage. This could lead to less carbonation and a flatter drink. Conversely, lower temperatures allow more gas to dissolve, potentially enhancing the beverage's fizz and creating a more refreshing product. However, temperature fluctuations could also affect the consistency of carbonation, impacting the beverage's overall quality.

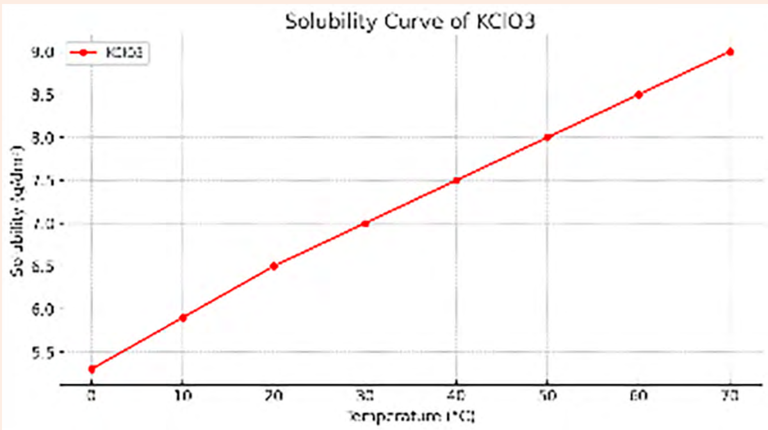
2. **a.** Plot of a solubility curve for NaCl



b. A point above the solubility curve represents a supersaturated solution.

c. The solubility of NaCl at 30°C is approximately 33.75 g/100g water.

3. **a.** **A plot of the Solubility curve of $KClO_3$**



b. A point on the solubility curve represents a saturated solution.

c. Solubility of $KClO_3$ at 35°C is approximately 7.25 g/100g water.

d. The temperature will be 67°C if the solubility is 8.75 g/dm^3?

4. Experimental Procedure to Differentiate Between Chloride (Cl^-) and Carbonate (CO_3^{2-}) Ions samples.

Tests	Observations	Inferences
Add distilled water to a small sample of the compound in test tubes.	*The sample dissolved in distilled water to form a colourless aqueous solution.*	*The compounds are water-soluble salts.*
Add a few drops of silver nitrate solution to a portion of the solution.	*A white precipitate of silver chloride (AgCl) will form.*	*Chloride ions (Cl^-) are present.*
Adding excess ammonia solution to the precipitate.	*White precipitate AgCl dissolves in excess ammonia.*	*Chloride ion (Cl^-) is confirmed.*
Add dilute HCl to another portion of the solution.	*Effervescence (bubbles of CO_2) will occur.*	*Carbonate ions (CO_3^{2-}) are present.*
Pass the gas through limewater ($Ca(OH)_2$).	*The limewater turns milky.*	*Carbonate ion is confirmed due to the formation of calcium carbonate ($CaCO_3$).*

Rationale

- *Silver Nitrate Test*: $AgNO_3$ reacts specifically with Cl^- to form a white precipitate, while CO_3^{2-} does not react in the same way.
- *Hydrochloric Acid Test:* HCl reacts with CO_3^{2-} to release CO_2 gas, which can be visually observed as effervescence.
- *Confirmatory Tests:* Additional tests confirm the identity of the ion by further reacting the products in a predictable manner, ensuring accurate identification.

5. Procedure to Identify Ions in an Unknown Compound Step-by-Step Using Solubility Rules

- Dissolve the unknown compound in distilled water.
- Use litmus paper or a pH meter to check the pH of the solution.

- Test for halides (Cl^-, Br^-, I^-) using silver nitrate ($AgNO_3$).
- Test for sulphate ions (SO_4^{2-}) using barium chloride ($BaCl_2$).
- Test for carbonate ions (CO_3^{2-}) using dilute hydrochloric acid.
- Test for ammonium ions (NH_4^+) using sodium hydroxide.
- Test for metal ions using aqueous sodium hydroxide or ammonia.
- Based on initial observations, perform additional confirmatory tests specific to the suspected ions.
- Document all observations and interpret them based on the solubility rules and the specific reactions observed.

EXTENDED READING

Refer to the links below to watch videos on how to prepare a solution:

- https://www.youtube.com/watch?v=iPYyRNjXkgY
- https://www.youtube.com/watch?v=A2YyIo8vSCA

REFERENCES

1. Brønsted, J.N., 1922. Studies on solubility. IV. The principle of the specific interaction of ions. *Journal of the American Chemical Society*, *44*(5), pp.877-898.
2. Voigt, W., 2011. Chemistry of salts in aqueous solutions: Applications, experiments, and theory. *Pure and Applied Chemistry*, *83*(5), pp.1015-1030.

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