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2018 HIGHER SCHOOL CERTIFICATE COURSE MATERIALS

Preliminary Chemistry

Water

Term 1 – Week 3

Name …………………………………………………………….

Class day and time …………………………………………

Teacher name ……………………………………………...

Term 1 – Week 3 – Theory

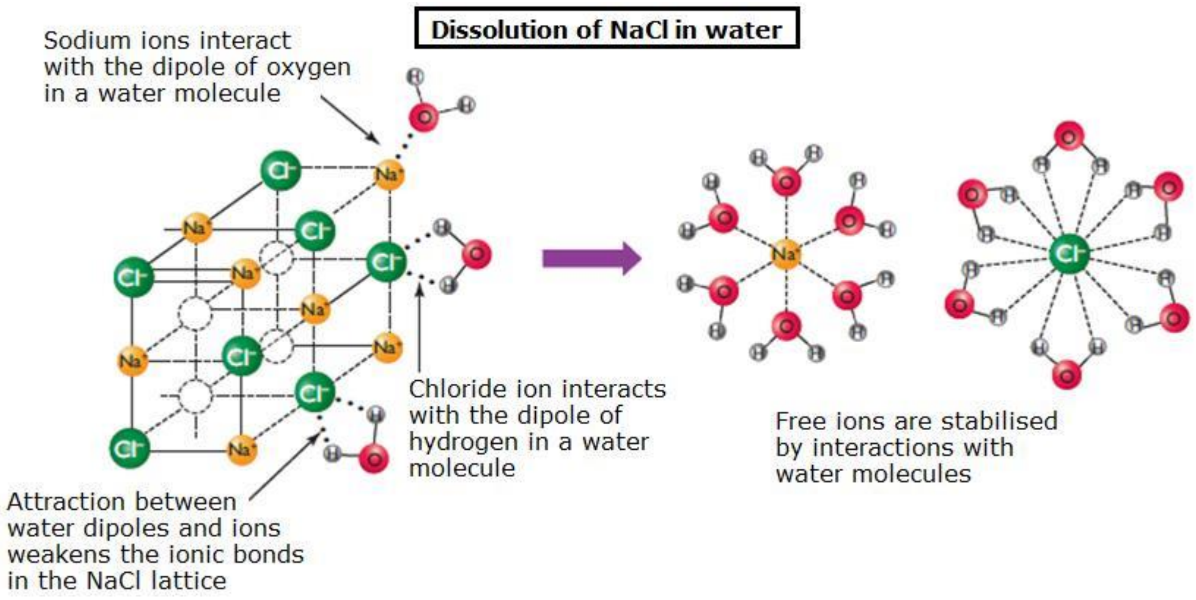
WATER IS AN IMPORTANT SOLVENT

- **Explain changes, if any, to particles and account for those changes when the following types of chemicals interact with water:**
    - **a soluble ionic compound such as sodium chloride**
    - **a soluble molecular compound such as sucrose**
    - **a soluble or partially soluble molecular element or compound such as iodine, oxygen or hydrogen chloride**
    - **a covalent network structure substance such as silicon dioxide**
    - **a substance with large molecules, such as cellulose or polyethylene**
- **Process information from secondary sources to visualise the dissolution in water of various types of substances and solve problems by using models to show the changes tha toccur in particle arrangement as dissolution occurs**

Soluble ionic compound

When a soluble ionic compound such as sodium chloride dissolves in water, the lattice breaks up into free ions which are able to move freely in the solution. This is because dipoles on water molecules cause them to attach onto free ions due to the strong electrostatic attraction. The attractive forces between water and the ions are stronger than the attractive forces between the positive and negative ions in the NaCl lattice, causing the NaCl crystal to dissolve.

The slightly negatively charged oxygen atom attaches to the positive ions (Na^+) while the slightly positively charged hydrogen atom attaches to the negative ions (Cl^-). Each ion is thus surrounded by several water molecules and becomes hydrated (see below).

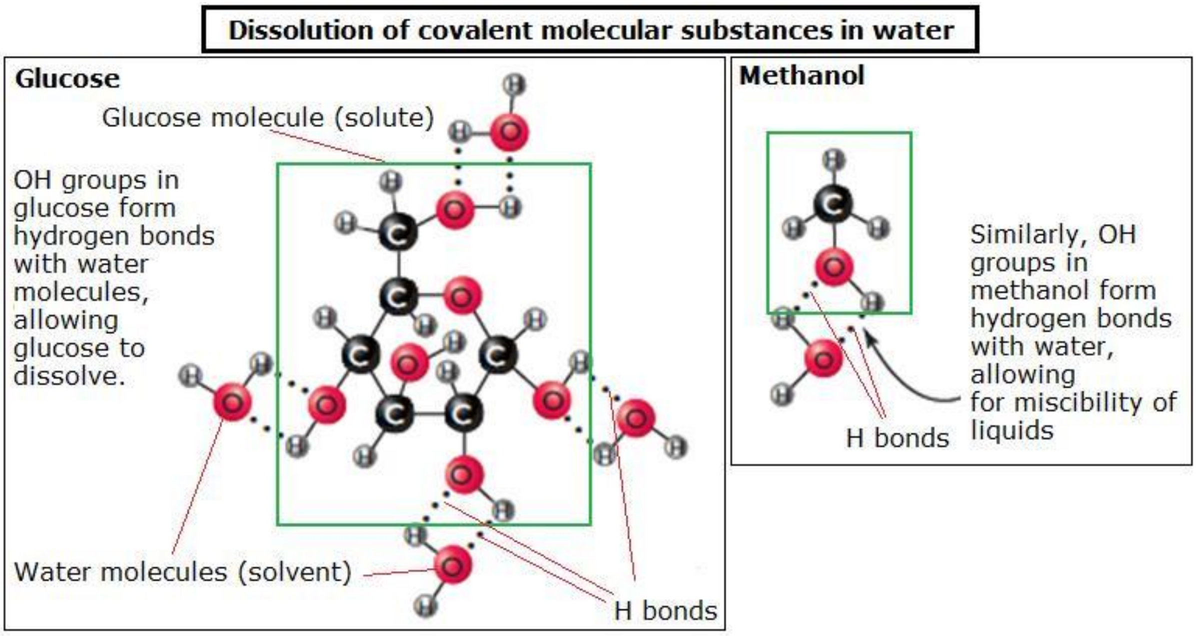


Soluble molecular compound

When a soluble molecular compound such as glucose dissolves into water, the crystal breaks up into individual molecules that form hydrogen bonds with individual water molecules. Some soluble molecular compounds include:

- Simple sugars (glucose, sucrose, fructose)
- Alcohols (methanol, ethanol, propenol)
- Fats (glycerols)

These substances all contain hydroxyl (OH) groups that can form strong hydrogen bonds with water molecules. When these substances dissolve, the intermolecular forces within the substance are overcome by stronger hydrogen bonding with water molecules. Individual molecules of the substance are then carried away by water molecules (see below).



Soluble or partially soluble molecular element or compound

Non-polar molecular substances like oxygen gas, nitrogen gas, bromine (Br_2) and iodine (I_2) are slightly soluble in water. Non-polar substances interact only through dispersion forces. As non-polar substances do not have net dipoles, no dipole-dipole forces or hydrogen bonding is possible.

Weak dispersion forces are enough to allow water to dissolve small amounts of non-polar substances like oxygen and nitrogen. (Remember, H2O cannot form hydrogen bonds with substances like O_2 and N_2 as they also need to be bonded to a H atom) However the solubility of these substances are much lower than substances that can form hydrogen bonds, or ionic substances.

Acidic molecules

Some molecules like HCl **react with water** to form an acidic solution, instead of dissolving in the usual way:

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

Upon exposure to water, the hydrogen chloride molecule becomes ionised as the covalent bond breaks. This ionisation reaction produced a hydronium ion as the hydrogen ions move to the water molecules and a chlorine ion. Substances that react with water to form hydronium ions are acidic while substances that from hydroxide ions are alkaline. Other common acids behave in the same way. For example:

- H_2SO_4: $H_2SO_{4(l)} + 2H_2O_{(l)} \rightarrow 2H_3O^+_{(aq)} + SO_4^{2-}{}_{(aq)}$
- H_2CO_3: $H_2CO_{3(l)} + 2H_2O_{(l)} \rightarrow 2H_3O^+_{(aq)} + CO_3^{2-}{}_{(aq)}$
- HNO_3: $HNO_{3(l)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$

Covalent network structure substance

A covalent network structure substance such as silicon dioxide **does not react or dissolve in water**. Covalent network substances are covalently bonded throughout the network. Covalent bonds are stronger than all intermolecular forces, even hydrogen bonding.

For example, silicon dioxide is essentially glass. When silicon dioxide is introduced to water, there is no observable dissolution or reaction.

Substance with large molecules

Substances with large molecules, such as cellulose or polyethylene also tend not to dissolve or react with water. They have large molecular weights in the hundreds of thousands, and are not soluble because the molecules are held strongly to one another in an orderly fashion by strong intermolecular forces.

For cellulose, long cellulose chains are strongly held together by hydrogen bonding, and there is no chance water can overcome this intermolecular force. This is the reason why wood (mostly cellulose) is so stiff and strong.

For polyethylene (common plastic), water is also unable to dissolve this, mainly because the individual molecules are long hydrocarbon chains which are non-polar. Water can only dissolve polar substances (hence why water is bad at removing oils). Also, at extremely high molecular weights such as in polymers like plastics, dispersion forces get strong enough to resist the effect of water's hydrogen bonding.

- **Analyse the relationship between the solubility of substances in water and the polar nature of the water molecule**

Like dissolves like

This statement refers to the fact that:

- **Polar solvents** (water, ammonia etc) are good at dissolving **polar or ionic substances** (glucose, ethanol, NaCl, $CaCl_2$)
- **Non-polar solvents** (vegetable oil, petrol, kerosene, acetone) are good at dissolving **non-polar substances** (plastics, other oils, dirt)

Below is a table of example substances and their solubility in water.

Bonding type	**Solubility in water**	**Examples**
Ionic	• Most are soluble	• NaCl, KNO_3, $MgSO_4$
Polar molecular	• Soluble if H bonds possible • Soluble by reaction • Otherwise insoluble	• Ethanol • HCl • dichloromethane
Non-polar molecular	• Some are very slightly soluble due to dispersion forces	• O_2 • Benzene
Large molecules	• Generally insoluble • Exceptions where extensive H-bonds are possible	• Cellulose (insoluble) • Starch (soluble)
Covalent lattices	• Insoluble	• Diamond • Glass • Quartz
Metals	• Insoluble • Unless they react with water	• Al, Zn, Au etc • Li, Na

- **<u>Perform a first-hand investigation to test the solubilities in water of a range of substances that include ionic, soluble molecular, insoluble molecular, covalent networks and large molecules</u>**

Materials that can be tested

In preparing for exams, it is a good idea to have an idea of what are common substances that can be tested in a school lab. Below is a list of common substances available in school labs that can be used in this solubility experiment.

Substance	Nature of bonding and structure
1. Sodium chloride	1. Ionic
2. Calcium carbonate	2. Ionic
3. Iodine	3. Covalent, non-polar
4. Sucrose	4. Covalent, polar
5. Ethanol	5. Covalent, polar
6. Polyethylene	6. Covalent, large molecules
7. Cyclohexane	7. Covalent, non-polar
8. Kerosene	8. Covalent, non-polar
9. Sand	9. Covalent lattice

Method

For all solids, pinch a sample and drop it into a test tube half filled with water. Observe whether it has dissolved and note whether it is totally or partially soluble, or totally insoluble.

For liquids, use a pipette or eyedropper to apply a few drops of the liquid into a test tube of water. Observe whether the liquids are miscible.

Results

Substance	Solubility in water
1. Sodium chloride	1. Fully soluble
2. Calcium carbonate	2. Slightly soluble
3. Iodine	3. Slightly soluble
4. Sucrose	4. Fully soluble
5. Ethanol	5. Fully miscible
6. Polyethylene	6. Insoluble
7. Cyclohexane	7. Immiscible
8. Kerosene	8. Immiscible
9. Sand	9. Insoluble

Note that calcium carbonate may appear insoluble, but in fact, all ionic substances dissolve to a very slight extent.

Risk assessment

Most of the substances in the above list are harmless. Ethanol is relatively harmless, but technically you should not touch it with bare hands, as pro-longed exposure can be toxic. Hydrocarbons tested, such as cyclohexane, ethanol and kerosene are highly flammable, so this test should be done away from flames and sources of heat or sparks. Lastly, cyclohexane is an irritant to mucous membranes such as the eyes, so it is a good idea to wear safety goggles when testing cyclohexane.

THE CONCENTRATION OF SALTS IN WATER WILL VARY ACCORDING TO THEIR SOLUBILITY, AND PRECIPITATION CAN OCCUR WHEN THE IONS OF AN INSOLUBLE SALT ARE IN SOLUTION TOGETHER

- **Identify some combinations of solutions which will produce precipitates, using solubility data**
- **Present information in balanced chemical equations and identify the appropriate phase descriptors (s), (l), (g), and (aq) for all chemical species**

Precipitates

Precipitation is the formation of a solid in a solution during a chemical reaction. When the reaction occurs, the solid formed is called the precipitate. Most commonly, precipitates form when we mix two solutions containing a combination of anion and cation that form an insoluble solid when mixed. For example, if we mix copper nitrate solution with sodium carbonate solution, we get:

$$Cu(NO_3)_{2(aq)} + Na_2(CO_3)_{2(aq)} \rightarrow 2NaNO_{3(aq)} + CuCO_{3(s)}$$

Copper carbonate forms as a precipitate, as it is an insoluble salt.

Solubility rules

Below is a summary of solubility rules for the ions you will encounter for the Preliminary and HSC course. Note that this table does not cover the more complex ions that will not be studied in the HSC.

Ions that form **soluble** compounds	**Exceptions**
Group I ions (Li^+, Na^+, K^+ etc)	
Ammonium (NH_4^+)	
Nitrate (NO_3^-)	
Acetate (CH_3COO^-)	
Hydrogen carbonate (HCO_3^-)	
Chlorate (ClO_3^-)	
Perchlorate (ClO_4^-)	
Group VII ions, excluding F^- (Cl^-, Br^-, I^-)	When combined with Ag^+, Pb^{2+} and Hg^{2+}
Sulfates (SO_4^{2-})	When combined with Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+} and Pb^{2+}

Ions that form **insoluble** compounds	**Exceptions**
Carbonate (CO_3^{2-})	When combined with Group I ions, or ammonium (NH_4^+)
Chromate (CrO_4^{2-})	When combined with Group I ions, Ca^{2+}, Mg^{2+}, or ammonium (NH_4^+)
Phosphate (PO_4^{3-})	When combined with Group I ions or ammonium (NH_4^+)
Sulfide (S^{2-})	When combined with Group I ions or ammonium (NH_4^+)
Hydroxide (OH^-)	When combined with Group I ions, Ca^{2+}, Ba^{2+}, Sr^{2+}, or ammonium (NH_4^+)

It is a good idea to **remember the above two tables**, as these rules will be a great help to you throughout your study in HSC Chemistry and in exams (solubility rules may not be given to you in exams).

Solubility examples

Example 1

Barium nitrate reacts with zinc sulfate:

$$Ba(NO_3)_{2(aq)} + ZnSO_{4(aq)} \rightarrow BaSO_{4(s)} + Zn(NO_3)_{2(aq)}$$

Because barium sulfate is insoluble (from solubility rules), when barium ions come into contact with sulfate ions, they form a precipitate.

Example 2

Lead nitrate reacts with potassium iodide:

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow PbI_{2(s)} + 2KNO_{3(aq)}$$

Again, from the solubility rules, we know lead iodide is insoluble, so if we mix a lead solution with an iodide solution, lead iodide will form a precipitate.

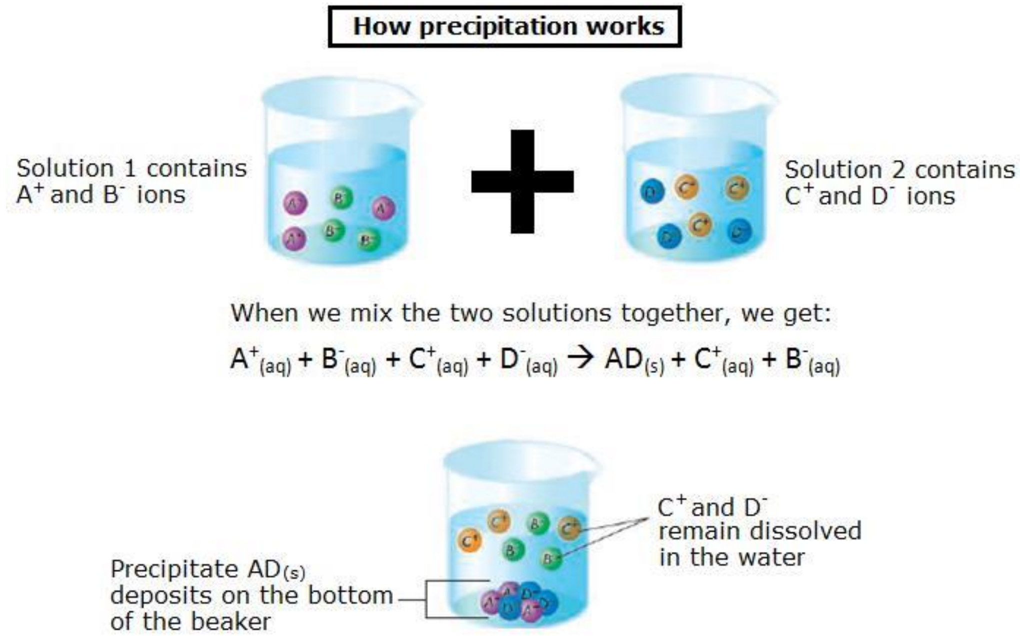
Example 3

Calcium nitrate reacts with sodium carbonate:

$$Ca(NO_3)_{2(aq)} + Na_2CO_{3\,(aq)} \rightarrow CaCO_{3(s)} + 2NaNO_{3(aq)}$$

- **<u>Describe a model that traces the movement of ions when solution and precipitation occur</u>**

This dot-point requires students to understand how precipitation actually works, in terms of movement of ions in solution. The diagram below illustrates the precipitation process:



Precipitates form when a combination of cation and anion are insoluble together, and come in contact in the same solution. When ions are in solution, they are in constant motion and in contact with each other. If the electrostatic attraction between any combination of cation and anion is stronger than the electrostatic attraction with water molecules, the cation and anion join up to form an insoluble salt. This is how precipitation occurs.

- **<u>Identify the dynamic nature of ion movement in a saturated dissolution</u>**
- **<u>Construct ionic equations to represent the dissolution and precipitation of ionic compounds in water</u>**

Ions in a saturated solution

A saturated solution is a solution that is at the point of maximum concentration where it cannot dissolve further salt. For example, if you keep adding NaCl to a glass of water, eventually the salt will appear to stop dissolving, and additional salt added remains a solid.

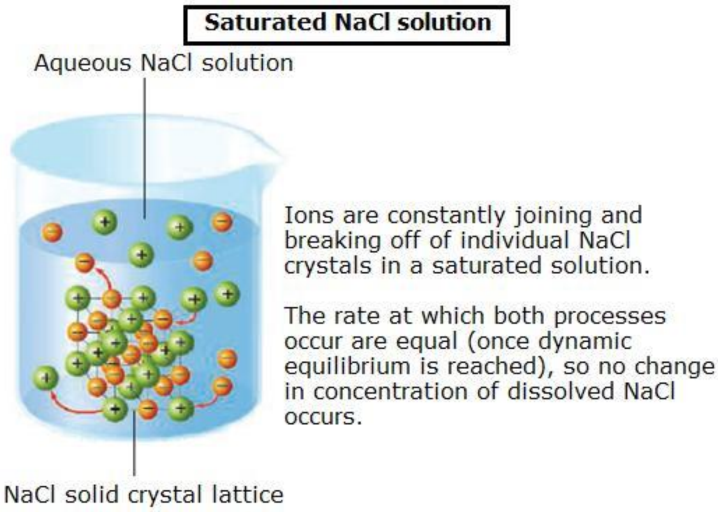
When an ionic substance is immersed in water, it breaks up into ions which move independently through the solution. When a solution becomes saturated, ions of the solid **still continue to dissolve** and go into solution but also ions in solution **precipitate out of solution** at an equal rate. This is called a **dynamic equilibrium**, because the rate of dissolution is equal to the rate of precipitation, in a saturated solution. Both reactions are occurring at equal rates so that there is no overall change in concentration in the solution.

To represent a dynamic equilibrium of a saturated solution, we use the ↔ symbol to denote a reversible reaction.

Example 1

Consider a saturated aqueous solution of NaCl:

$$NaCl_{(s)} \leftrightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$



Example 2

Consider a saturated aqueous solution of $PbCl_2$ (lead chloride is slightly soluble in water, so it requires less salt to make a saturated solution):

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

Again, the forward and reverse reactions occur at the same rate, in a saturated solution. If the solution is not yet saturated (i.e. concentration < maximum concentration), the forward reaction occurs slightly faster than the reverse reaction, until saturation occurs.

Example 3

Consider a saturated solution of lead sulfate:

$$Pb^{+2}_{(aq)} + SO_4^{2-}{}_{(aq)} \leftrightarrow PbSO_{4\,(s)}$$

- The forward reaction represents the precipitation of $PbSO_{4\,(s)}$.
- The reverse reaction represents the dissolution of $PbSO_{4\,(s)}$ to form Pb^{+2} and SO_4^{2-} ions in solution.

- **Describe the molarity of a solution as the number of moles of solute per litre of solution using:**

$$\boldsymbol{c = \frac{n}{V}}$$

- **Carry out simple calculations to describe the concentration of given solutions, given masses of solute and volumes of solution**
- **Calculate mass and concentration relationships in precipitation reactions as they are encountered**

Molarity

The molarity of a solution is defined as the **number of moles of solute per litre of solution**. Molarity is a **measure of concentration** of solutes in a solution (this is commonly a salt, acid, base or specific anions or cations).

Molarity is given by:

$$c = \frac{n}{V}$$

Where:

c is the molarity (mol/L)

n is the number of moles (no units)

V is the volume of solution (L)

Worked examples

Example 1

Calculate the molarity of the solution when 5.00g of silver nitrate is mixed with 1.20L of water.

First we calculate the number of moles of silver nitrate:

$$n = \frac{5}{molar\ mass\ of\ silver\ nitrate}$$

$$= \frac{5}{107.9 + 14 + 16 \times 3} = 0.0294mol$$

Next we use the known volume to calculate the molarity:

$$c = \frac{n}{V}$$

$$= \frac{0.0294}{1.2} = 0.0245mol/L$$

Example 2

Calculate the amount of silver nitrate in 1.5L of a 0.4mol/L solution.

First we calculate the number of moles in the solution by rearranging the formula:

$$c = \frac{n}{V}$$

$$n = cV$$

$$= 0.4 \times 1.5 = 0.6mol$$

Next we figure out the mass of $AgNO_3$ in grams, by multiplying the number of moles with the molecular mass of $AgNO_3$:

$$m = n \times (107.9 + 14 + 16 \times 3)$$

$$= 0.6 \times 169.9 = 101.94g$$

Example 3

How much anhydrous sodium carbonate is needed to make 1.2L of a 0.01M solution?

(Anhydrous means assuming sodium carbonate is in pure unhydrated form)

First we calculate the number of moles of Na_2CO_3 needed:

$$n = cV = 0.01 \times 1.2 = 0.012mol$$

Next we calculate the mass of Na_2CO_3 required:

$$m = n \times (2 \times 23 + 12 + 3 \times 16)$$

$$= 0.012 \times 106 = 1.272g$$

Hence 1.272g of anhydrous sodium carbonate is needed to make 1.2L of a 0.01M solution.

Example 4

How many mL of water is needed to dilute 200mL of a 0.5M solution of HCl to a 0.01M solution?

First we calculate the number of moles in the first solution:

$$n = cV = 0.5 \times 0.2 = 0.1mol$$

We know this is the same amount of moles in the diluted solution. We rearrange the equation to get V:

$$V = \frac{n}{c} = \frac{0.1}{0.01} = 10L$$

That is we need the second solution to be made up to 10 litres. Therefore we need an additional 9,800mL of water to be added.

Example 5

Find the molarity of a HCl solution if 25mL of this solution reacts completely with 0.256g of anhydrous sodium carbonate.

First we write out the chemical reaction between HCl and Na_2CO_3:

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$$

From this equation, we know that for each mole of Na_2CO_3 there needs to be 2 moles of HCl.

Now we calculate the number of moles of Na_2CO_3:

$$n_{sodium\ carbonate} = \frac{0.256}{2 \times 23 + 12 + 3 \times 16} = 2.415 \times 10^{-3} mol$$

We multiply this by 2 to get the number of moles of HCl required (from the chemical equation):

$$n_{hydrogen\ chloride} = 2 \times n_{sodium\ carbonate} = 4.830 \times 10^{-3} mol$$

Now we calculate the molarity of the HCl solution:

$$c = \frac{n}{V} = \frac{4.830 \times 10^{-3}}{0.025} = 0.1932 mol/L$$

Term 1 – Week 3 – Homework

WATER IS AN IMPORTANT SOLVENT

- **Explain changes, if any, to particles and account for those changes when the following types of chemicals interact with water:**
    - **a soluble ionic compound such as sodium chloride**
    - **a soluble molecular compound such as sucrose**
    - **a soluble or partially soluble molecular element or compound such as iodine, oxygen or hydrogen chloride**
    - **a covalent network structure substance such as silicon dioxide**
    - **a substance with large molecules, such as cellulose or polyethylene**
- **Process information from secondary sources to visualise the dissolution in water of various types of substances and solve problems by using models to show the changes that occur in particle arrangement as dissolution occurs**

1. Potassium nitrate (KNO_3) is a very soluble salt. Draw a diagram showing the anion and cation in water solution and illustrate how water molecules attach onto dissociated ions. Label your diagram. **[3 marks]**

2. Of the following covalent molecules, identify those which can dissolve in water: **[2 marks]**
    a. Hydrogen chloride (HCl)

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    b. Glucose ($C_6H_{12}O_6$)

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    c. Carbon dioxide (CO_2)

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- **<u>Analyse the relationship between the solubility of substances in water and the polar nature of the water molecule</u>**

1. Explain what is meant by the term "like dissolve like". **[2 marks]**

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2. Adding detergent is important to washing clothes. If dirt in dirty clothes are mostly non-polar, predict the role detergent plays in washing clothes. **[3 marks]**

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- **<u>Perform a first-hand investigation to test the solubilities in water of a range of substances that include ionic, soluble molecular, insoluble molecular, covalent networks and large molecules</u>**

1. In your study of Chemistry, you have performed a first-hand investigation to test the solubilities in water of a range of substances.

   Identify 2 sources of risk in this experiment. For each risk, suggest ways the risk can be managed. **[3 marks]**

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THE CONCENTRATION OF SALTS IN WATER WILL VARY ACCORDING TO THEIR SOLUBILITY, AND PRECIPITATION CAN OCCUR WHEN THE IONS OF AN INSOLUBLE SALT ARE IN SOLUTION TOGETHER

- **Identify some combinations of solutions which will produce precipitates, using solubility data**
- **Present information in balanced chemical equations and identify the appropriate phase descriptors (s), (l), (g), and (aq) for all chemical species**

1. Every salt in every solvent has a limit to how much the salt can dissolve. Beyond this limit, the solution is saturated, and precipitates will form. Describe 2 ways how precipitates are formed. **[3 marks]**

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2. For each solution mixture below, predict whether a precipitate will form. If you predict a precipitate will form, identify the precipitate's chemical formula. **[1 mark each]**
   a. $KNO_3 + (NH_4)_2CO_3$

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   b. $CuSO_4 + BaCl_2$

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3. The following are word equations of solutions to be mixed. Convert the following word equations into balanced chemical equations, with the proper state labelled for each species. **[1 mark each]**
    a. Sodium carbonate + calcium chloride

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    b. Potassium phosphate + ammonium carbonate

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    c. Ammonium hydroxide + iron nitrate

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- **<u>Describe a model that traces the movement of ions when solution and precipitation occur</u>**

1. When a lead nitrate solution is added to a sodium carbonate solution, a precipitate forms. Describe the movement of ions as the solutions are mixed. **[2 marks]**

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2. When a solution of calcium nitrate is added to a solution of sodium hydroxide, a precipitate forms. Describe the movement of ions as the solutions are mixed. **[2 marks]**

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- **<u>Identify the dynamic nature of ion movement in a saturated dissolution</u>**
- **<u>Construct ionic equations to represent the dissolution and precipitation of ionic compounds in water</u>**

1. Dynamic equilibriums can be described as having separate reactions of dissolution and precipitation occurring at equal rates simultaneously.
    a. If a solution is not yet saturated, identify which reaction rate (dissolution or precipitation) proceeds faster. **[1 mark]**

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    b. If a solution is supersaturated (e.g. by changing the temperature), identify which reaction rate (dissolution or precipitation) proceeds faster. **[1 mark]**

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    c. Hence explain in terms of competing reaction rates how dynamic equilibrium is reached. **[2 marks]**

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2. Construct ionic equations of the following processes: **[1 mark each]**
   a. Dissolution of ammonium chloride in water

   b. Dynamic equilibrium of sodium chloride in water

   c. Dissolution of copper sulfate in water

- **<u>Describe the molarity of a solution as the number of moles of solute per litre of solution using:</u>**

$$c = \frac{n}{V}$$

- **<u>Carry out simple calculations to describe the concentration of given solutions, given masses of solute and volumes of solution</u>**
- **<u>Calculate mass and concentration relationships in precipitation reactions as they are encountered</u>**

1. Calculate the mass of salt required to make the following concentrations: **[1 mark each]**
   a. Amount of silver nitrate in 2.6L of 0.6M solution

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   b. Amount of sodium carbonate in 1.5L of 0.8mol/L solution

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2. Find the molarity of a sulfuric acid solution if 150mL reacts completely with 5.42g of $NaOH.2H_2O$. (Note: questions 8 and 9 are important for titration calculations, which will be covered in year 12) **[3 marks]**

End of homework