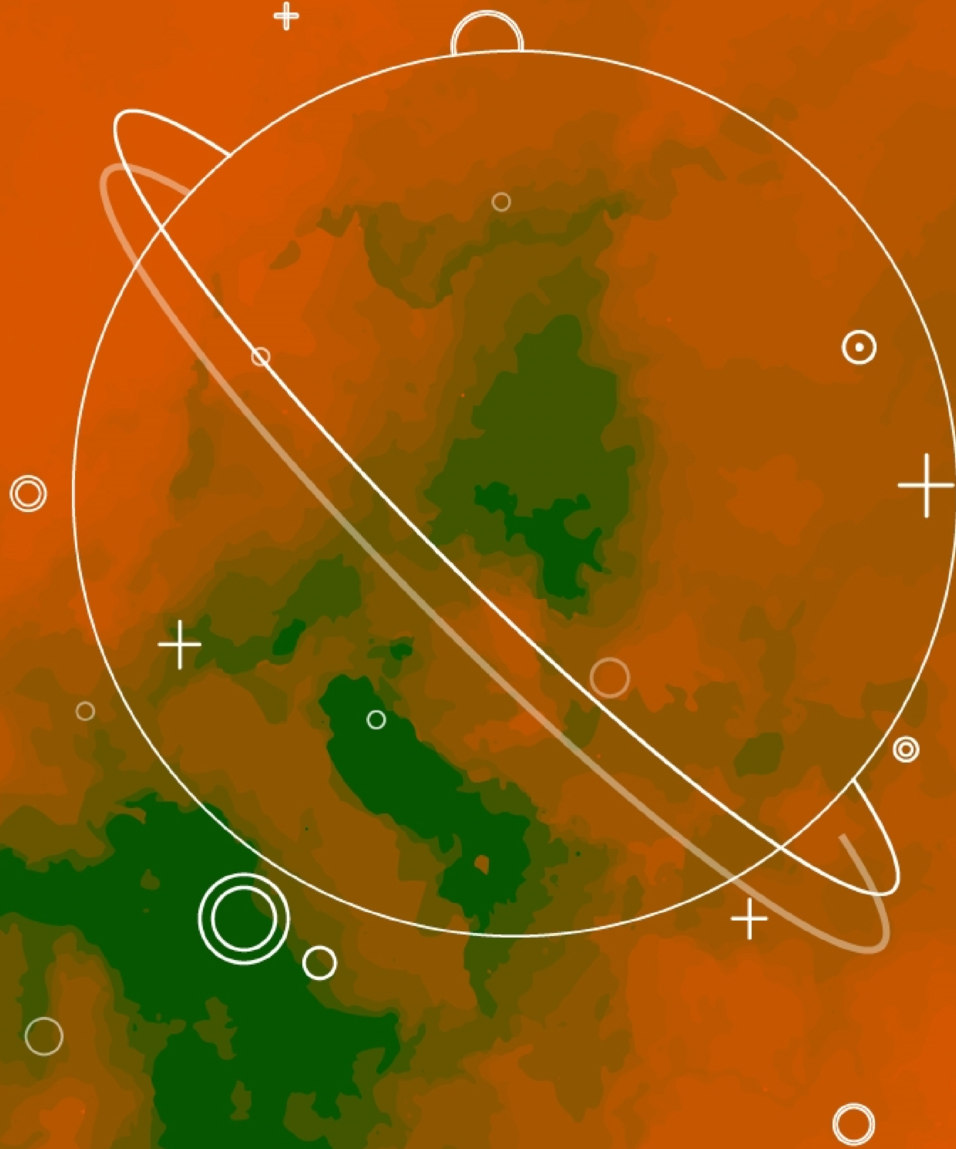

CHAPTER 18

Reactions In Aqueous Solution

MathU Teaching Emporium

CONTENTS

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1 INTRODUCTION

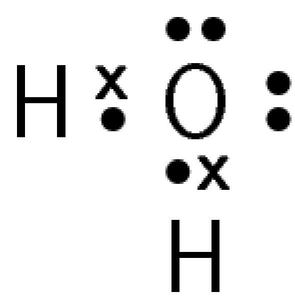
Many reactions in chemistry and all biological reactions (reactions in living systems) take place in water. We say that these reactions take place in aqueous solution. Water has many unique properties and is plentiful on Earth. For these reasons reactions in aqueous solutions occur frequently. In this chapter we will look at some of these reactions in detail. Almost all the reactions that occur in aqueous solutions involve ions. We will look at three main types of reactions that occur in aqueous solutions, namely precipitation reactions, acid-base reactions and redox reactions. Before we can learn about the types of reactions, we need to first look at ions in aqueous solutions and electrical conductivity.

2 IONS IN AQUEOUS SOLUTION

Water is seldom pure. Because of the structure of the water molecule, substances can dissolve easily in it. This is very important because if water wasn't able to do this, life would not be possible on Earth. In rivers and the oceans for example, dissolved oxygen means that organisms (such as fish) are able to respire (breathe). For plants, dissolved nutrients are available in a form which they can absorb. In the human body, water is able to carry dissolved substances from one part of the body to another.

2.1 Dissociation in water

Water is a **polar molecule**. If we represent water using Lewis structures we will get the following:



You will notice that there are two electron pairs that do not take part in bonding. This side of the water molecule has a higher electron density than the other side where the hydrogen atoms are bonded. This side of the water molecule is more negative than the side where the hydrogen atoms are bonded. We say this side is the delta negative (δ^-) side and the hydrogen side is the delta positive (δ^+) side. This means that one part of the molecule has a slightly positive charge (positive pole) and the other part has a slightly negative charge (negative pole). We say such a molecule is a dipole. It has two poles. Figure 18.1 shows this.

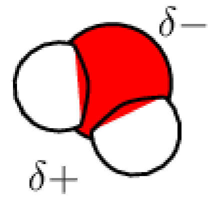


Figure 18.1: Water is a polar molecule

2.2 Dissociation of sodium chloride in water

It is the polar nature of water that allows ionic compounds to dissolve in it. In the case of sodium chloride (NaCl) for example, the positive sodium ions (Na^{+}) are attracted to the negative pole of the water molecule, while the negative chloride ions (Cl^{-}) are attracted to the positive pole of the water molecule. When sodium chloride is dissolved in water, the polar water molecules are able to work their way in between the individual ions in the lattice. The water molecules surround the negative chloride ions and positive sodium ions and pull them away into the solution. This process is called dissociation. Note that the positive side of the water molecule will be attracted to the negative chlorine ion and the negative side of the water molecule to the positive sodium ions. A simplified representation of this is shown in Figure 18.2. We say that dissolution of a substance has occurred when a substance dissociates or dissolves. Dissolving is a physical change that takes place. It can be reversed by removing (evaporating) the water.

DEFINITION

Dissociation

Dissociation is a general process in which ionic compounds separate into smaller ions, usually in a reversible manner.

DEFINITION

Dissolution

Dissolution or dissolving is the process where ionic crystals break up into ions in water.

DEFINITION

Hydration

Hydration is the process where ions become surrounded with water molecules.

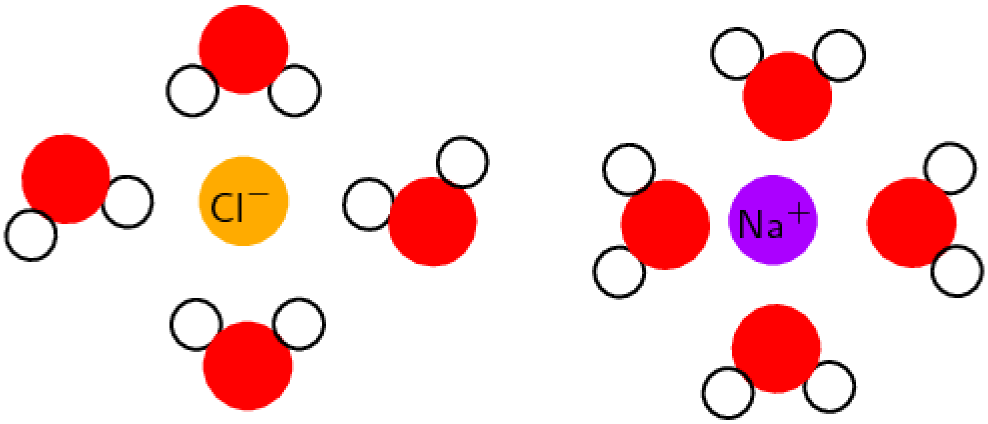


Figure 18.2: Sodium chloride dissolves in water

The dissolution of sodium chloride can be represented by the following equation:

$$NaCl(s) \longrightarrow Na^{+}(aq) + Cl^{-}(aq)$$

The dissolution of potassium sulfate into potassium and sulfate ions is shown below as another example:

$$K_2SO_4(s) \longrightarrow 2K^{+}(aq) + SO^{2-}{}_4(aq)$$

Remember that **molecular** substances (e.g. covalent compounds) may also dissolve, but most will not form ions. One example is glucose.

$$C_6H_{12}O_6(s) \longrightarrow C_6H_{12}O_6(aq)$$

There are exceptions to this and some molecular substances will form ions when they dissolve. Hydrogen chloride for example can ionise to form hydrogen and chloride ions.

$$HCl(g) + H_2O(l) \longrightarrow H_3O^{+}(aq) + Cl^{-}(aq)$$

You can try dissolving ionic compounds such as potassium permanganate, sodium hydroxide and potassium nitrate in water and observing what happens.

3 ELECTROLYTES, IONISATION AND CONDUCTIVITY

You have learnt that water is a polar molecule and that it can dissolve ionic substances in water. When ions are present in water, the water is able to conduct electricity. The solution is known as an electrolyte.

DEFINITION

Electrolyte
An electrolyte is a substance that contains free ions and behaves as an electrically conductive medium.

Because electrolytes generally consist of ions in solution, they are also known as ionic solutions. A strong electrolyte is one where many ions are present in the solution and a weak electrolyte is one where few ions are present. Strong electrolytes are good conductors of electricity and weak electrolytes are weak conductors of electricity. Non-electrolytes do not conduct electricity at all. Conductivity in aqueous solutions, is a measure of the ability of water to conduct an electric current. The more ions there are in the solution, the higher its conductivity. Also the more ions there are in solution, the stronger the electrolyte.

Factors that affect the conductivity of electrolytes

The conductivity of an electrolyte is therefore affected by the following factors:

- The **concentration of ions** in solution. The higher the concentration of ions in solution, the higher its conductivity will be.
- The **type of substance** that dissolves in water. Whether a material is a strong electrolyte (e.g. potassium nitrate, KNO_3), a weak electrolyte (e.g. acetic acid, CH_3COOH) or a non-electrolyte (e.g. sugar, alcohol, oil) will affect the conductivity of water because the concentration of ions in solution will be different in each case. Strong electrolytes form ions easily, weak electrolytes do not form ions easily and non-electrolytes do not form ions in solution.
- **Temperature**. The warmer the solution, the higher the solubility of the material being dissolved and therefore the higher the conductivity as well.

GENERAL EXPERIMENT

Electrical conductivity

Aim

To investigate the electrical conductivities of different substances and solutions.

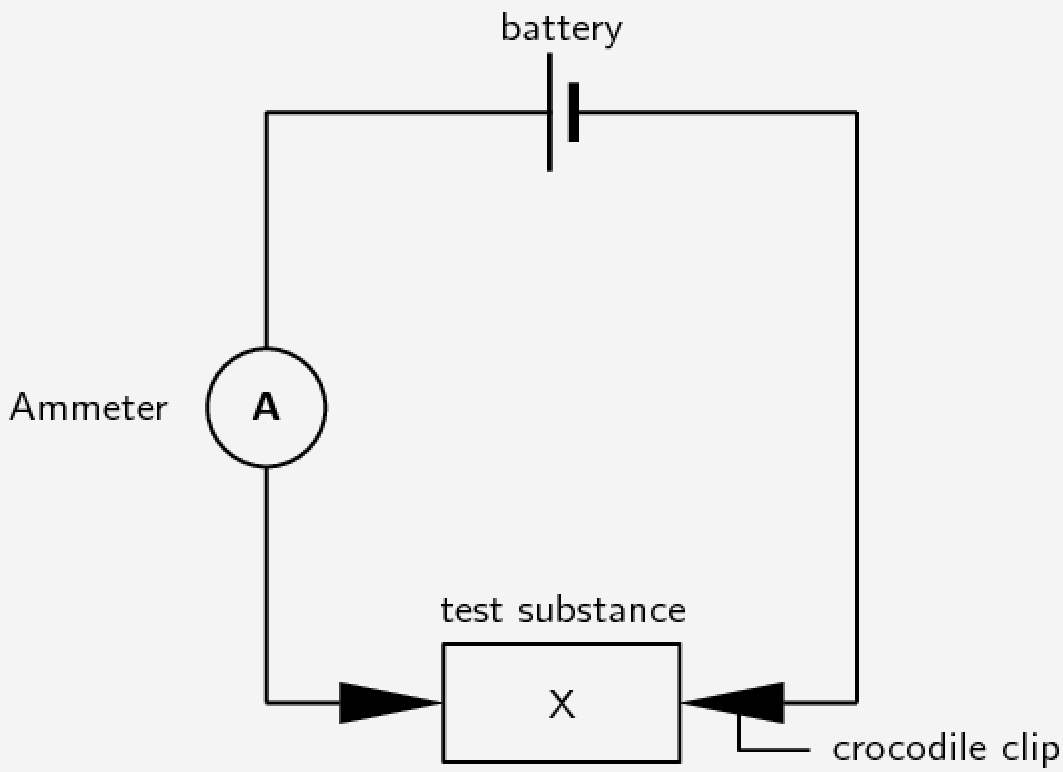
Apparatus

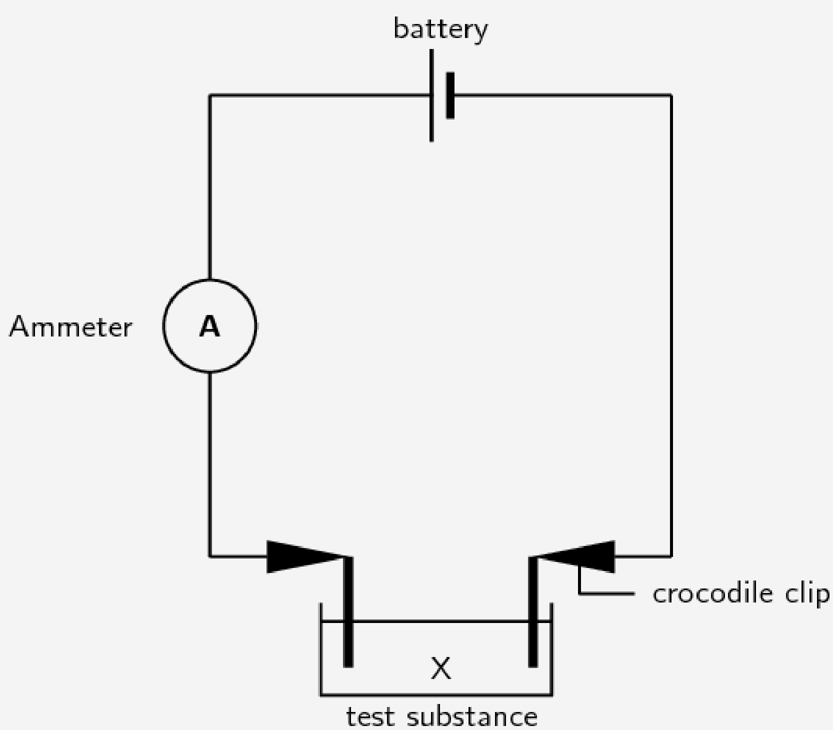
- Solid salt (NaCl) crystals
- different liquids such as distilled water, tap water, seawater, sugar, oil and alcohol
- solutions of salts e.g. NaCl, KBr, $CaCl_2$, NH_4Cl
- a solution of an acid (e.g. HCl) and a solution of a base (e.g. NaOH)
- torch cells
- ammeter
- conducting wire, crocodile clips and 2 carbon rods.

Method

1. Set up the experiment by connecting the circuit as shown in the diagram below. In the diagram, X represents the substance or solution that you will be testing.
2. When you are using the solid crystals, the crocodile clips can be attached directly to each end of the crystal. When you are using solutions, two carbon rods are placed into the liquid and the clips are attached to each of the rods.
3. In each case, complete the circuit and allow the current to flow for about 30 seconds.
4. Observe whether the ammeter shows a reading.

GENERAL EXPERIMENT (continued)





GENERAL EXPERIMENT (continued)

Results Record your observations in a table similar to the one below:

Test substance	Ammeter reading

What do you notice? Can you explain these observations?

Conclusions

Solutions that contain free-moving ions are able to conduct electricity because of the movement of charged particles. Solutions that do not contain free-moving ions do not conduct electricity.

Remember that for electricity to flow, there needs to be a movement of charged particles e.g. ions. With the solid NaCl crystals, there was no flow of electricity recorded on the ammeter. Although the solid is made up of ions, they are held together very tightly within the crystal lattice and therefore no current will flow. Distilled water, oil and alcohol also don't conduct a current because they are covalent compounds and therefore do not contain ions.

The ammeter should have recorded a current when the salt solutions and the acid and base solutions were connected in the circuit. In solution, salts dissociate into their ions, so that these are free to move in the solution. Look at the following examples:

Dissociation of potassium bromide:

$$KBr(s) \longrightarrow K^{+}(aq) + Br^{-}(aq)$$

Dissociation of table salt:

$$NaCl(s) \longrightarrow Na^{+}(aq) + Cl^{-}(aq)$$

Ionisation of hydrochloric acid:

$$HCl(l) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

Dissociation of sodium hydroxide:

$$NaOH(s) \longrightarrow Na^+(aq) + OH^-(aq)$$

4 PRECIPITATION REACTIONS

Sometimes, ions in solution may react with each other to form a new substance that is insoluble. This is called a precipitate. The reaction is called a precipitation reaction.

DEFINITION

Precipitate

A precipitate is the solid that forms in a solution during a chemical reaction.

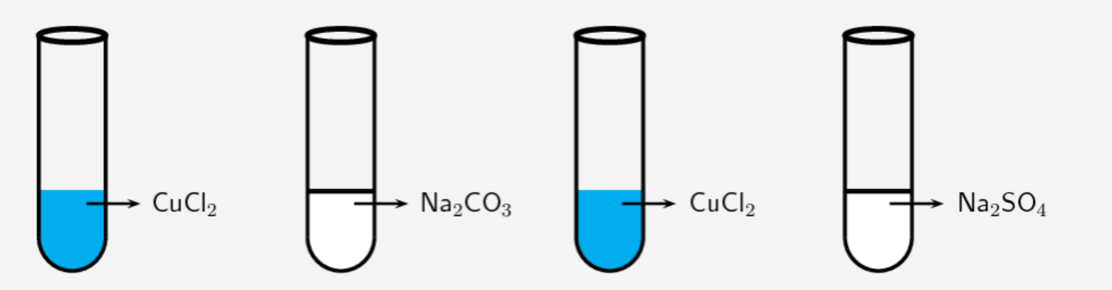
GENERAL EXPERIMENT

The reaction of ions in solution

Aim

To investigate the reactions of ions in solutions. **Apparatus**

4 test tubes; copper(II) chloride solution; sodium carbonate solution; sodium sulphate solution



GENERAL EXPERIMENT (continued)

Method

1. Prepare 2 test tubes with approximately 5 mL of dilute copper(II) chloride solution in each
2. Prepare 1 test tube with 5 mL sodium carbonate solution
3. Prepare 1 test tube with 5 mL sodium sulphate solution
4. Carefully pour the sodium carbonate solution into one of the test tubes containing copper(II) chloride and observe what happens
5. Carefully pour the sodium sulphate solution into the second test tube containing copper(II) chloride and observe what happens

Results

1. A light blue precipitate forms when sodium carbonate reacts with copper(II) chloride.
2. No precipitate forms when sodium sulphate reacts with copper(II) chloride. The solution is light blue.

It is important to understand what happened in the previous demonstration. We will look at what happens in each reaction, step by step.

For reaction 1 you have the following ions in your solution: Cu^{2+}, Cl^{-}, Na^{+} and CO^{2-}_{3}. A precipitate will form if any combination of cations and anions can become a solid. The following table summarises which combination will form solids (precipitates) in solution.

Salt	Solubility
Nitrates	All are **soluble**
Potassium, sodium and ammonium salts	All are **soluble**
Chlorides, bromides and iodides	All are **soluble** except silver, lead(II) and mercury(II) salts (e.g. silver chloride)
Sulphates	All are **soluble** except lead(II) sulphate, barium sulphate and calcium sulphate
Carbonates	All are **soluble** except those of potassium, sodium and ammonium
Compounds with fluorine	Almost all are **soluble** except those of magnesium, calcium, strontium (II), barium (II) and lead (II)
Perchlorates and acetates	All are **soluble**
Chlorates	All are **soluble** potassium chlorate
Metal hydroxides and oxides	Most are **insoluble**

Table 18.1: General rules for the solubility of salts

TIP

Salts of carbonates, phosphates, oxalates, chromates and sulphides are generally insoluble.

If you look under carbonates in the table it states that all carbonates are insoluble except potassium sodium and ammonium. This means that Na_2CO_3 will dissolve in water or remain in solution, but CuCO3 will form a precipitate. The precipitate that was observed in the reaction must therefore be $CuCO_3$. The balanced chemical equation is:

$$2Na^{+}(aq) + CO_3{}^{2-}(aq) + Cu^{+}(2+)(aq) + 2Cl^{-}(aq) \longrightarrow CuCO_3 + 2Na^{+}(aq) + 2Cl^{-}(aq)$$

Note that sodium chloride does not precipitate and we write it as ions in the equation. For **reaction 2** we have Cu^{2+}, Cl^{-}, Na^{+} and $SO^{2-}{}_4$ in solution. Most chlorides and sulphates are soluble according to the table. The balanced chemical equation is:

$$2Na^{+}(aq) + SO_4{}^{2-}(aq) + Cu^{2+}(aq) + 2Cl^{-}(aq) \longrightarrow SO_4{}^{2-}(aq) + 2Na^{+}(aq) + 2Cl^{-}(aq) + Cu^{2+}(aq)$$

Both of these reactions are ion exchange reactions.

Tests for anions

We often want to know which ions are present in solution. If we know which salts precipitate, we can derive tests to identify ions in solution. Given below are a few such tests.

Test for a chloride

Prepare a solution of the unknown salt using distilled water and add a small amount of silver nitrate solution. If a white precipitate forms, the salt is either a chloride or a carbonate.

$$NO_3^-(aq) + Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s) + NO_3(aq)^-$$

(AgCl is white precipitate)

$$CO_3^{2-}(aq) + 2Ag^+(aq) + 2NO_3^-(aq) + Cl^-(aq) \longrightarrow Ag_2CO_3(s) + 2NO_3(aq)^-$$

(Ag_2CO_3 is white precipitate)

The next step is to treat the precipitate with a small amount of **concentrated nitric acid**. If the precipitate remains unchanged, then the salt is a chloride. If carbon dioxide is formed and the precipitate disappears, the salt is a carbonate.

$$HNO_3(l) + AgCl(s) \longrightarrow \text{no reaction, precipitate is unchanged}$$

$$Ag_2CO_3(s) + 2HNO_3(l) \longrightarrow 2Ag^+(aq) + 2NO_3^-(aq) + H_2O(l) + CO_2(g) \text{ precipitate disappears}$$

Test for bromides and iodides

As was the case with the chlorides, the bromides and iodides also form precipitates when they are reacted with silver nitrate. Silver chloride is a white precipitate, but the silver bromide and silver iodide precipitates are both pale yellow. To determine whether the precipitate is a bromide or an iodide, we use chlorine water and carbon tetrachloride (CCl_4).

Chlorine water frees bromine gas from the bromide and colours the carbon tetrachloride a reddish brown.

$$2Br^-(aq) + Cl_2(aq) \longrightarrow Br_2(g) + 2Cl^-(aq)$$

Chlorine water frees iodine gas from an iodide and colours the carbon tetrachloride purple.

$$2I^{-}(aq) + Cl_2(aq) \longrightarrow I_2(g) + 2Cl^{-}(aq)$$

Test for a sulphate

Add a small amount of barium chloride solution to a solution of the test salt. If a white precipitate forms, the salt is either a sulfate or a carbonate.

$$S0^{2-}{}_4(aq) + Ba^{2+}(aq) + Cl^{-}(aq) \longrightarrow BaSO_4(s) + Cl^{-}(aq)$$

($BaSO_4$ is a white precipitate)

$$C0^{2-}{}_3(aq) + Ba^{2+}(aq) + Cl^{-}(aq) \longrightarrow BaCO_3(s) + Cl^{-}(aq)$$

($BaCO_4$ is a white precipitate)

If the precipitate is treated with nitric acid, it is possible to distinguish whether the salt is a sulphate or a carbonate (as in the test for a chloride).

$$BaSO_4(s) + HN0_3(l) \longrightarrow \text{no reaction, precipitate is unchanged}$$
$$BaCO_3(s) + 2HNO_3(l) \longrightarrow Ba^{2+}(aq) + 2NO_3{}^{-}(aq) + H_2O(l) + CO_2(g) \text{ precipitate disappears}$$

Test for a carbonate

If a sample of the dry salt is treated with a small amount of acid, the production of carbon dioxide is a positive test for a carbonate.

$$2HCl + K_2CO_3(aq) \longrightarrow CO_2(g) + 2KCl(aq) + H_2O(l)$$

If the gas is passed through limewater (an aqueous solution of calcium hydroxide) and the solution becomes milky, the gas is carbon dioxide.

$$Ca^{2+}(aq) + 2OH^{-}(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$$

(It is the insoluble $CaCO_3$ precipitate that makes the limewater go milky)

5 OTHER TYPES OF REACTIONS

We will look at two types of reactions that occur in aqueous solutions. These are ion-exchange reactions and redox reactions. Ion exchange reactions include precipitation reactions, gas forming reactions and acid-base reactions. Redox reactions are electron transfer reactions. It is important to remember the difference between these two types of reactions. In ion exchange reactions ions are exchanged, in electron transfer reactions electrons are transferred. These terms will be explained further in the following sections.

Ion exchange reactions

Ion exchange reactions can be represented by:

$$AB(aq) + CD(aq) \longrightarrow AD + CB$$

Either AD or CB may be a solid or a gas. When a solid forms this is known as a precipitation reaction. If a gas is formed then this may be called a gas forming reaction. Acid-base reactions are a special class of ion exchange reactions and we will look at them separately.

The formation of a precipitate or a gas helps to make the reaction happen. We say that the reaction is driven by the formation of a precipitate or a gas. All chemical reactions will only take place if there is something to make them happen. For some reactions this happens easily and for others it is harder to make the reaction occur.

DEFINITION

Ion exchange reaction

A type of reaction where the positive ions exchange their respective negative ions due to a driving force.

DID YOU KNOW?

Ion exchange reactions are used in ion exchange chromatography. Ion exchange chromatography is used to purify water and as a means of softening water. Often when chemists talk about ion exchange, they mean ion exchange chromatography.

We have already looked at precipitation reactions.

Gas forming reactions

These reactions are similar to precipitation reactions with the exception that instead of a precipitate forming, a gas is formed instead. An example of a gas forming reaction is sodium carbonate in hydrochloric acid. The balanced equation for this reaction is:

$$Na_2CO_3(s) + 2HCl(aq) + \longrightarrow 2CO_2(g) + 2NaCl(aq) + H_2O(l)$$

Acid-base reactions

Acid-base reactions take place between acids and bases. In general, the products will be water and a salt (i.e. an ionic compound). An example of this type of reaction is:

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

This is an special case of an ion exchange reaction since the sodium in the sodium hydroxide swaps places with the hydrogen in the hydrogen chloride forming sodium chloride. At the same time the hydroxide and the hydrogen combine to form water.

Redox reactions

Redox reactions involve the exchange of electrons. One ion loses electrons and becomes more positive, while the other ion gains electrons and becomes more negative. To decide if a redox reaction has occurred we look at the charge of the atoms, ions or molecules involved. If one of them has become more positive and the other one has become more negative then a redox reaction has occurred. For example, sodium metal is oxidised to form sodium oxide (and sometimes sodium peroxide as well). The balanced equation for this is:

$$4Na + O_2 \longrightarrow 2Na_2O(aq)$$

In the above reaction sodium and oxygen are both neutral and so have no charge. In the products however, the sodium atom has a charge of 1+ and the oxygen atom has a charge of 2−. This tells us that the sodium has lost electrons and the oxygen has gained electrons. Since one species has become more positive and one more negative we can conclude that a redox reaction has occurred. We could also say that electrons have been transferred from one species to the other. (In this case the electrons were transferred from the sodium to the oxygen).

GENERAL EXPERIMENT

Demonstration: Oxidation of sodium metal

You will need a Bunsen burner, a small piece of sodium metal and a metal spatula. Light the Bunsen burner. Place the sodium metal on the spatula. Place the sodium in the flame. When the reaction finishes, you should observe a white powder on the spatula. This is a mixture of sodium oxide (Na_2O) and sodium peroxide (Na_2O_2).

WARNING

Sodium metal is very reactive. Sodium metal reacts vigorously with water and should never be placed in water. Be very careful when handling sodium metal.

GENERAL EXPERIMENT

Reaction types

Aim To use experiments to determine what type of reaction occurs.

Apparatus Soluble salts (e.g. potassium nitrate, ammonium chloride, sodium carbonate, silver nitrate, sodium bromide); hydrochloric acid (HCl); sodium hydroxide (NaOH); bromothymol blue; zinc metal; copper (II) sulphate; beakers; test-tubes

Method

- For each of the salts, dissolve a small amount in water and observe what happens.
- Try dissolving pairs of salts (e.g. potassium nitrate and sodium carbonate) in water and observe what happens.
- Dissolve some sodium carbonate in hydrochloric acid and observe what happens.
- Carefully measure out 20 cm^3 of sodium hydroxide into a beaker.
- Add some bromothymol blue to the sodium hydroxide
- Carefully add a few drops of hydrochloric acid to the sodium hydroxide and swirl. Repeat until you notice the colour change.
- Place the zinc metal into the copper sulphate solution and observe what happens.

Results

Answer the following questions:

- What did you observe when you dissolved each of the salts in water?
- What did you observe when you dissolved pairs of salts in the water?
- What did you observe when you dissolved sodium carbonate in hydrochloric acid?
- Why do you think we used bromothymol blue when mixing the hydrochloric acid and the sodium hydroxide? Think about the kind of reaction that occurred.
- What did you observe when you placed the zinc metal into the copper sulphate?
- Classify each reaction as either precipitation, gas forming, acid-base or redox.
- What makes each reaction happen (i.e. what is the driving force)? Is it the formation of a precipitate or something else?
- What criteria would you use to determine what kind of reaction occurs?
- Try to write balanced chemical equations for each reaction

GENERAL EXPERIMENT (continued)

Conclusion

We can see how we can classify reactions by performing experiments.

In the experiment above, you should have seen how each reaction type differs from the others. For example, a gas forming reaction leads to bubbles in the solution, a precipitation reaction leads to a precipitate forming, an acid-base reaction can be seen by adding a suitable indicator and a redox reaction can be seen by one metal disappearing and a deposit forming in the solution.

6 CHAPTER SUMMARY

- The **polar** nature of water means that **ionic compounds** dissociate easily in aqueous solution into their component ions.
- Dissociation is a general process in which ionic compounds separate into smaller ions, usually in a reversible manner.
- Dissolution or dissolving is the process where ionic crystals break up into ions in water.
- Hydration is the process where ions become surrounded with water molecules.
- **Conductivity** is a measure of a solution's ability to conduct an electric current.
- An **electrolyte** is a substance that contains free ions and is therefore able to conduct an electric current. Electrolytes can be divided into **strong** and **weak** electrolytes, based on the extent to which the substance ionises in solution.
- A **non-electrolyte** cannot conduct an electric current because it does not contain free ions.
- The **type of substance**, the **concentration of ions** and the **temperature** of the solution affect its conductivity.
- There are three main types of reactions that occur in aqueous solutions. These are precipitation reactions, acid-base reactions and redox reactions.
- Precipitation and acid-base reactions are sometimes known as ion exchange reactions. Ion exchange reactions also include gas forming reactions. Ion exchange reactions are a type of reaction where the positive ions exchange their respective negative ions due to a driving force.
- A **precipitate** is formed when ions in solution react with each other to form an insoluble product. Solubility rules help to identify the precipitate that has been formed.
- A number of tests can be used to identify whether certain **anions** (chlorides, bromides, iodides, carbonates, sulphates) are present in a solution.

- An acid-base reaction is one in which an acid reacts with a base to form a salt and water.
- A redox reaction is one in which electrons are transferred from one substance to another.

7 EXERCISES

7.1 Exercise 1

1. For each of the following, say whether the substance is ionic or molecular.
   1.1 potassium nitrate (KNO_3)
   1.2 ethanol (C_2H_5OH)
   1.3 sucrose (a type of sugar) ($C_{12}H_{22}O_{11}$)
   1.4 sodium bromide ($NaBr$)
2. Write a balanced equation to show how each of the following ionic compounds dissociate in water.
   2.1 sodium sulphate (Na_2SO_4)
   2.2 potassium bromide (KBr)
   2.3 potassium permanganate ($KMnO_4$)
   2.4 sodium phosphate (Na_3PO_4)
3. Draw a diagram to show how KCl dissolves in water.

7.2 Exercise 2

1. Silver nitrate ($AgNO_3$) reacts with potassium chloride (KCl) and a white precipitate is formed.
   1.1 Write a balanced equation for the reaction that takes place. Include the state symbols.
   1.2 What is the name of the insoluble salt that forms?
   1.3 Which of the salts in this reaction are soluble?
2. Barium chloride reacts with sulphuric acid to produce barium sulphate and hydrochloric acid.
   2.1 Write a balanced equation for the reaction that takes place. Include the state symbols.
   2.2 Does a precipitate form during the reaction?
   2.3 Describe a test that could be used to test for the presence of barium sulphate in the products.
3. A test tube contains a clear, colourless salt solution. A few drops of silver nitrate solution are added to the solution and a pale yellow precipitate forms. Chlorine water and carbon tetrachloride were added, which resulted in a purple solution.

3.1 Which one of the following salts was dissolved in the original solution?

i. NaI

ii. KCl

iii. K_2CO_3

iv. Na_2SO_4

3.2 Write the balanced equation for the reaction that took place between the salt and silver nitrate.

8 ANSWERS FOR EXERCISES

8.1 Exercise 1

1.1 Ionic

1.2 Molecular

1.3 Molecular

1.4 Ionic

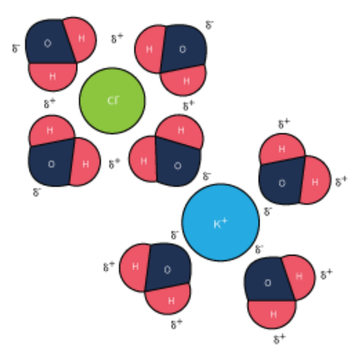
2.1 $Na_2SO_4(s) + H_2O(l) \rightarrow 2Na^+(aq) + SO_4^{2-}(aq)$

2.2 $KBr(s) + H_2O(l) \rightarrow K^+(aq) + Br^-(aq)$

2.3 $KMnO_4(s) + H_2O(l) \rightarrow K^+(aq) + MnO_4^-(aq)$

2.4 $Na_3PO_4(s) + H_2O(l) \rightarrow 3Na^+(aq) + PO_4^{3-}(aq)$

3.



8.2 Exercise 2

1.1 $AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$

1.2 Silver chloride ($AgCl$)

1.3 Potassium chloride(KCl) , silver nitrate($AgNO_3$) , potassium nitrate(KNO_3)

2.1 $BaCl_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$

2.2 Yes, barium sulfate($BaSO_4$)

2.3 Add a small amount of barium chloride to the solution. If a white precipitate forms the salt is either a sulphate or carbonate. To test if the salt is a sulphate the precipitate is then treated with nitric acid. If the precipitate remains unchanged the salt is a sulphate.

3.1 i. NaI (Sodium Iodide)

3.2 $NaI + AgNO_3 \rightarrow NaNO_3 + AgI$