CHAPTER 15: REACTIONS IN AQUEOUS SOLUTIONS

Enhanced Introductory College Chemistry

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Chapter Contents

Please visit the web version of Enhanced Introductory College Chemistry (https://ecampusontario.pressbooks.pub/enhancedchemistry/) to access the complete book, interactive activities and ancillary resources.

In this chapter, you will learn about

- The properties of salts and give examples

- The properties of electrolytes and classifying them as strong or weak.
- Precipitation reactions in aqueous solutions
- Writing molecular equations, complete ionic equations, net ionic equations (for neutralization, precipitation, and gas formation reactions)
- Concepts relating to the ionization of water

To better support your learning, you should be familiar with the following concepts before starting this chapter:

- Nomenclature for ionic compounds: naming ionic compounds and writing its chemical formula
- How to use your scientific calculator including the following functions (buttons): scientific notation.
- Polyatomic ions
- Review double displacement reactions and balancing chemical equations
- Solubility concepts
- Properties of water

We have learned in the previous section(s) that solutions are of extreme importance in everyday life. The food we eat, the liquids we drink, the fluids in our body, the air we breathe, and the household products we use, such as household cleaners, hand sanitizers, and medications, are all solutions. We often focus on solutions only being table salt dissolving in water when, actually, particles like pollutants suspended in the air are also solutions. This particular example is a gaseous solution. Metal alloys, such as brass (which is mix of copper and zinc), is an example of a solid solution.

This chapter will focus on reactions occurring specifically in aqueous solutions. Water is by far the most important liquid solvent, partly because it is plentiful and partly because of its unique properties. In your body, in other living systems, and in the outside environment a tremendous number of reactions take place in aqueous solutions. Consequently this section, as well as significant portions of many other sections of this text, are devoted to developing an understanding of reactions which occur in water. Since ionic compounds and polar covalent compounds constitute the main classes which are appreciably soluble in water, reactions in aqueous solutions usually involve these types of substances. There are three important classes of reactions which occur in aqueous solution: precipitation reactions, acid-base reactions, and oxidation-reduction (redox) reactions.

- Precipitation reactions are useful for detecting the presence of various ions and for determining the concentrations of solutions.
- Acid-base reactions and redox reactions are similar in that something is being transferred from one species to another.
    - Acid-base reactions involve proton transfers, whereas redox reactions involve electron transfers.
    - Redox reactions are somewhat more complicated, though, because proton transfers and other bond-making and bond-breaking processes occur at the same time as electron transfer.

Below are demonstrations of each of the types of reactions. Some of these reactions will be explained in detail in this chapter; others will be covered in more depth in later chapters.

The first video demonstrates precipitation reaction between aqueous solutions of silver nitrate and sodium chloride.

Watch Double Displacement Reaction of $AgNO_3$ and NaCl (41 sec) (https://youtu.be/eGG3EI4mwok)

The next video demonstrates an acid and base reacting with aluminum (in the form of a soda can) at room temperature. When aluminum reacts with hydrochloric acid, it yields aqueous aluminum chloride and colourless hydrogen gas. When aluminum reacts with sodium hydroxide it forms water soluble sodium aluminate and hydrogen gas as products.

Watch Coke Cans in Acid and Base – Periodic Table of Videos (2min 49sec) (https://youtu.be/WnPrtYUKke8)

The last video shows a displacement reaction of zinc metal in aqueous copper (II) sulfate solution.

Watch Displacement Reaction of Metals – Zinc in Copper (II) Sulfate – with explanation at micro level (5min 45 sec) (https://youtu.be/6NrE6TZ8Fzw).

Attribution & References

Except where otherwise noted, this section is adapted by Jackie MacDonald from "11.1: Prelude to Aqueous Phase Reactions (https://chem.libretexts.org/Bookshelves/General_Chemistry/Book%3A_ChemPRIME_(Moore_et_al.)/11%3A_Reactions_in_Aqueous_Solutions/11.01%3A_Prelude_to_Aqueous_Phase_Reactions)" In *ChemPRIME (LibreTexts CHEMISTRY)* by Ed Vitz, John W. Moore, Justin Shorb, Xavier Prat-Resina, Tim Wendorff, & Adam Hahn, licensed under CC BY-NC-SA 4.0. / Adaptations and additions to content was updated for student comprehension.

15.1 SALTS

Learning Objectives

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

- Define the term salt and identify common salts
- Explain how a salt is formed

Salts are found all around us and within our human body. In nature, these ionic compounds are abundant in rocks and minerals of the Earth's mantel but also found in seawater and other bodies of water. We have already learned about salts when discussing ionic compounds. Salts are a chemical compound formed when ions form ionic bonds. In these reactions, one atom gives up one or more electrons, and thus becomes positively charged, whereas the other accepts one or more electrons and becomes negatively charged; overall the ionic compound has no net charge. Salts typically are crystalline, odourless, colourless/transparent or white, and have high melting and boiling points. As we learned in the previous chapter on solutions, many salts are soluble in water; however, some are not. If a given salt is soluble in water, it completely dissociates into ions other than a hydrogen ion (H^+) or hydroxide ion (OH^-) and forms an aqueous solution. This fact is elemental in distinguishing salts from acids and bases. Salts are derived from the neutralization reaction of an acid and base. Since acids and bases always contain either a metal cation or a cation derived from ammonium (NH_4^+) and a nonmetal anion, the two can combine to form a salt. For instance, when you mix aqueous solutions of hydrochloric acid and sodium hydroxide, aqueous sodium chloride and water are formed.

$$\mathrm{HCl(aq)+NaOH}(aq) \rightarrow \mathrm{NaCl}(aq)\mathrm{+H_2O}(l)$$

If the aqueous solution of sodium chloride undergoes a distillation process, a solid, crystalline NaCl salt would remain. Concepts related to acids and bases will be discussed in more detail in the next chapter.

A common salt, NaCl, also known as table salt, dissociates completely in water (Figure 15.1a). The positive and negative regions on the water molecule (the hydrogen and oxygen ends respectively) attract the negative chloride and positive sodium ions, pulling them away from each other. Non-polar and polar covalently bonded compounds break apart into molecules in solution; however, salts dissociate into ions. These ions are electrolytes, which will be discussed in the next section. This ionic property is critical to the function of ions in the human body when transmitting nerve impulses and prompting muscle contraction.

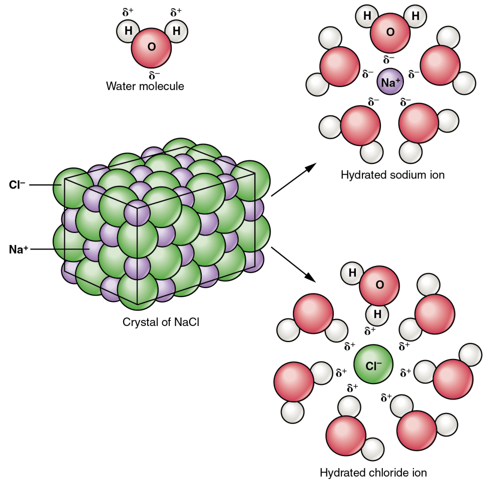

Figure 15.1a Dissociation of Sodium Chloride in Water: Notice that the solid crystals of sodium chloride dissociate not into molecules of NaCl, but into Na^+ cations and Cl^- anions, each completely surrounded by water molecules (credit: *Anatomy and Physiology 2e (Open Stax)*, CC BY 4.0 (http://creativecommons.org/licenses/by/4.0/))

Many other salts are important in the body. For example, bile salts produced by the liver help break apart dietary fats, and calcium phosphate salts, $Ca_3(PO_4)_2$ (*s*), form the mineral portion of teeth and bones. Some other commons salts include:

- Na_2CO_3 (sodium bicarbonate): Baking soda
- $MgSO_4$ – (magnesium sulfate): Epsom salt
- Fe_2O_3 (iron (III) oxide): Rust
- KNO_3 (potassium nitrate): component of crop fertilizer, used in rocket propellants, fireworks, gunpowder

Indigenous Perspective: Bone

Bone has two components: organic matter (which is mostly Type I collagen) and about 70% inorganic matter. The inorganic material of bone consists of calcium phosphate mineral crystals known as hydroxyapatite, $Ca_5(PO_4)_3(OH)$. This mineral is embedded throughout the organic bone matrix to give bone its structural stability and strength.

One of the composite materials traditionally important to Inuit has been bone. This comes from a variety of natural sources, including caribou antler and walrus tusk. An important use of this strong carvable material is in snow goggles, called *ilgaak* or *iggaak*. Wearing these human-made goggles are invaluable in preventing snow blindness (Rayner-Canham et al., 2016, Composite Materials section).

Figure 15.1b Ilgaak/Iggaak – Inuit snow goggles: Snow goggles such as the ones illustrated are commonly made from caribou antler with caribou sinew for a strap (credit: work by CambridgeBayWeather, PD)

Not only are salts crucial compounds necessary for sustaining life, but they are also invaluable constituents of chemical reactions that occur around us on a daily basis.

Watch What are Salts? (5min 9s) (https://youtu.be/WnAKhtnJjz0).

Attribution & References

Except where otherwise noted, this section is adapted by Jackie MacDonald from "2.4 Inorganic Compounds Essential to Human Functioning" In *Anatomy and Physiology 2e (Open Stax)* by J. Gordon Betts, Kelly A. Young, James A. Wise, Eddie Johnson, Brandon Poe, Dean H. Kruse, Oksana Korol, Jody E. Johnson, Mark Womble, Peter DeSaix is licensed under CC BY 4.0 (http://creativecommons.org/licenses/by/4.0/). Access for free at *Anatomy and Physiology 2e (OpenStax).* / Adaptations and additions to content was updated for student comprehension.

References

Rayner-Canham, G., Taylor, R., & Lee, Y.-R. (2016, February). *Making chemistry relevant to Indigenous*

Peoples. Chem 13 News Magazine. Retrieved December 15, 2022, from https://uwaterloo.ca/chem13-news-magazine/february-2016/feature/making-chemistry-relevant-indigenous-peoples

15.2 ELECTROLYTES

Learning Objectives

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

- Define electrolyte and give examples of electrolytes
- Relate electrolyte strength to solute-solvent attractive forces

Pure water (purified water) is a pure substance made up of just three atoms: two hydrogen and one oxygen atom, H_2O. Since pure water has no constituents other than these atoms, it does not have any taste or smell, and it doesn't conduct electricity on its own. However, water can become a medium for conducting electricity. When ionic substances (salts) are dissolved in water, they undergo either a physical or a chemical change that yields free ions moving independently in the aqueous solution. This feature permits them to carry positive or negative electrical charges from one place to another and the solution can conduct an electrical current. These substances constitute an important class of compounds called **electrolytes**. Salts that ionize in aqueous solution are great conductors of electricity and are known as electrolytes. Substances that do not yield ions when dissolved are called **nonelectrolytes**. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a **strong electrolyte**. If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, it is called a **weak electrolyte**.

Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 15.2a).

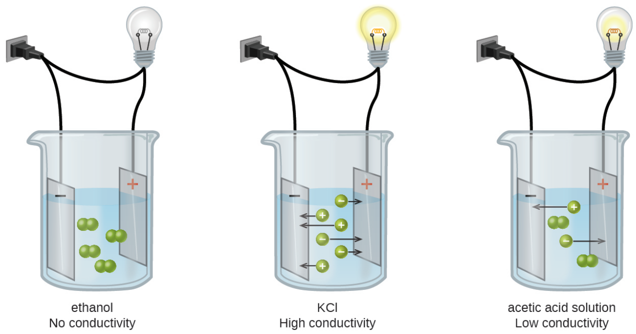

Figure 15.2a Solutions Containing Nonelectrolytes, Strong Electrolytes, or Weak Electrolytes: Solutions of nonelectrolytes such as ethanol do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Ionic Electrolytes

Water and other polar molecules are attracted to ions, as shown in Figure 15.2b. The electrostatic attraction between an ion and a molecule with a dipole is called an **ion-dipole attraction**. These attractions play an important role in the dissolution (ionization) of ionic compounds in water.

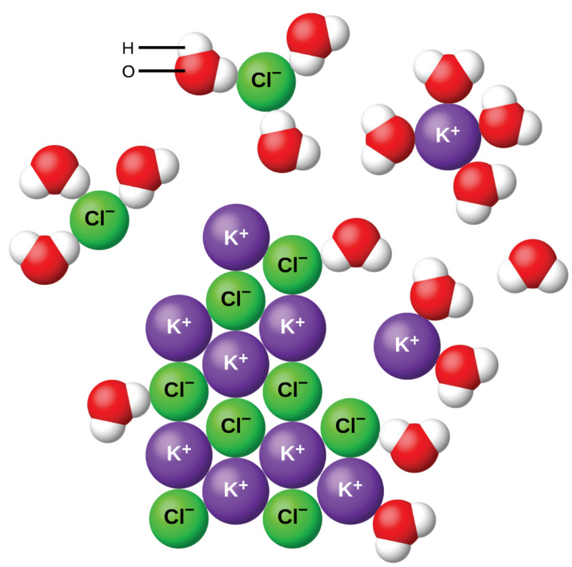

Figure 15.2b Representation of what Happens when Potassium Chloride is Dissolved in Water: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown (credit: *Chemistry (OpenStax)*, CC BY 4.0).

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as **dissociation**, which was discussed in the previous chapter – Solutions. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

Let us consider what happens at the microscopic level when we add solid KCl to water. Ion-dipole forces attract the positive (hydrogen) end of the polar water molecules to the negative chloride ions at the surface of the solid, and they attract the negative (oxygen) ends to the positive potassium ions. The water molecules penetrate between individual K^+ and Cl^- ions and surround them, reducing the strong interionic forces that bind the ions together and letting them move off into solution as solvated ions, as Figure 15.2b shows. The reduction of the electrostatic attraction permits the independent motion of each hydrated ion in a dilute solution, resulting in an increase in the disorder of the system as the ions change from their fixed and ordered positions in the crystal to mobile and much more disordered states in solution. This increased disorder is responsible for the dissolution of many ionic compounds, including KCl, which dissolve with the absorption of heat.

In other cases, the electrostatic attractions between the ions in a crystal are so large, or the ion-dipole attractive forces between the ions and water molecules are so weak, that the increase in disorder cannot compensate for the energy required to separate the ions, and the crystal is insoluble. Such is the case for salt compounds such as calcium carbonate (limestone), calcium phosphate (the inorganic component of bone), and iron oxide (rust).

Covalent Electrolytes

As mentioned previously, pure water is an extremely poor conductor of electricity because it is only very slightly ionized—only about two out of every 1 billion molecules ionize at 25 °C. The ionization of water will be discussed in more depth later in this chapter and these concepts will be important in understanding the behaviours of acids and bases. In brief, water ionizes when one molecule of water gives up a proton to another molecule of water, yielding hydronium and hydroxide ions.

$$\mathrm{H_2O}(l)\ +\ \mathrm{H_2O}(l) \leftrightharpoons \mathrm{H_3O^+}(aq)\ +\ \mathrm{OH^-}(aq)$$

In some cases, we find that solutions prepared from covalent compounds conduct electricity because the solute molecules react chemically with the solvent to produce ions. For example, pure hydrogen chloride is a gas consisting of covalently bonded HCl molecules. This gas contains no ions. Hydrogen chloride gas is very soluble in water and is dissolved in water to prepare hydrochloric acid. When we dissolve hydrogen chloride gas in water, we find that the solution is a very good conductor. The water molecules play an essential part in forming ions. However, it is important to note that solutions of hydrogen chloride in many other solvents, such as benzene, do not conduct electricity and do not contain ions.

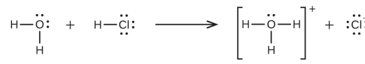

Figure 15.2c Reaction between Water and Hydrogen Chloride Gas to form Hydrochloric Acid: Hydrogen chloride gas dissolves in water and the ions react with water, transferring H^+ ions to form hydronium ions (H_3O^+) and chloride ions (Cl^-), which illustrates the ionization of this strong acid (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Hydrogen chloride gas dissolves in water and the ions react with water, transferring H^+ ions to form hydronium ions (H_3O^+) and chloride ions (Cl^-). This illustrates the ionization of a strong acid.

This reaction is essentially 100% complete for HCl (i.e., it is a strong acid and, consequently, a strong electrolyte). Likewise, weak acids and bases that only react partially generate relatively low concentrations of ions when dissolved in water and are classified as weak electrolytes. You will learn more about classifying the strength of acids and bases in later chapters.

For a summary **Watch Aqueous Solutions, Dissolving, and Solvation (14min 6sec). (https://youtu.be/AD22tefqhEQ)**

Attribution & References

Except where otherwise noted, this page is adapted by Jackie MacDonald from "11.2 Electrolytes (https://boisestate.pressbooks.pub/chemistry/chapter/11-2-electrolytes/)" In *General Chemistry 1 & 2* by Rice University, a derivative of *Chemistry (Open Stax)* by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson and is licensed under CC BY 4.0. Access for free at *Chemistry (OpenStax) (https://openstax.org/books/chemistry/pages/1-introduction)*. / Adaptations and additions were made to content in this section were made or student comprehension.

15.3 PRECIPITATION REACTIONS

Learning Objectives

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

- Predict the solubility of common inorganic compounds by using solubility rules
- Define precipitation reactions
- Recognize and identify examples of precipitation reactions
- Apply the solubility rules of common inorganic compounds to predict the products formed when two aqueous solutions are mixed

Scientists have found it convenient (or even necessary) to classify chemical interactions by identifying common patterns of reactivity. This section of this chapter will focus on a specific type of double displacement reaction called a precipitation reaction.

Precipitation Reactions and Solubility Rules

Solubility of Inorganic Compounds

The idea of solubility was introduced in the solutions chapter. The extent to which a substance may be dissolved in water, or any solvent, is quantitatively expressed as its **solubility**, defined as the maximum concentration of a substance that can be achieved under specified conditions. Substances with relatively large solubilities are said to be **soluble** and are found as dissolved ions in aqueous solution. A substance will **precipitate** when solution conditions are such that its concentration exceeds its solubility. Substances with relatively low solubilities are said to be **insoluble**, and these are the substances that readily precipitate from solution to form a solid (*s*). For purposes of predicting the identities of solids formed by precipitation reactions, one may simply refer to the solubility guidelines for many ionic compounds in Table 15.3a to predict whether a precipitation reaction will occur when solutions of soluble ionic compounds are mixed together. First, it is important to become familiar with using the solubility table to determine if a given salt will dissolve (*aq*), or not (*s*) in aqueous solution. If a salt is said to be insoluble, or has low solubility, or is slightly soluble, it will form a precipitate – a solid – and the symbol (*s*) will be used to represent that observation. If a salt is soluble, the salt will dissociate (ionize) in aqueous solution, and the symbol (*aq*) will be used to show that chemistry. It is important to mention that Table 15.3a does not include every possible soluble, insoluble salt combination. Since there are other possibilities, Table 14.2a and Table 14.2b can also be referenced to determine solubility of inorganic compounds.

Table 15.3a Solubility Guidelines for Inorganic Compounds in Water at 25°C

Negative Ion (Anion)	Positive Ion (Cation)	Solubility	Phase, Phase Symbol
All	Li^+, Na^+, K^+, Rb^+, Cs^+ NH_4^+	Soluble	aqueous, (*aq*)
Chloride (Cl^-), Bromide (Br^-), Iodide (I^-)	Ag^+, Pb^{2+}, Hg_2^{2+}, Cu^+	Low solubility	solid, (*s*)
Chloride (Cl^-), Bromide (Br^-), Iodide (I^-)	All others	Soluble	aqueous, (*aq*)
F^-	compounds with group 2 metal cations, Li^+, Al^{3+}, Pb^{2+}, $Fe^{2+}$and Fe^{3+}	Low solubility	solid, (*s*)
F^-	All others	Soluble	aqueous, (*aq*)
Hydroxide (OH^-)	Li^+, Na^+, K^+, Rb^+, NH_4^+, Sr^{2+}, Ba^{2+}	Soluble	aqueous, (*aq*)
Hydroxide (OH^-)	All others	Low solubility	solid, (*s*)
NO_3^-, NO_2^-	All (exception: $AgNO_2$ is insoluble)	Soluble	aqueous, (*aq*)
Phosphate (PO_4^{3-}), Carbonate (CO_3^{2-})	Na^+, K^+, Rb^+, NH_4^+	Soluble	aqueous, (*aq*)
Phosphate (PO_4^{3-}), Carbonate (CO_3^{2-})	All others	Low solubility	solid, (*s*)
Sulphate (SO_4^{2-})	Ca^{2+}, Sr^{2+}, Ba^{2+}, Ag^+, Pb^{2+}	Low solubility	solid, (*s*)
Sulphate (SO_4^{2-})	All others	Soluble	aqueous, (*aq*)

Source: “Table 15.3a Solubility Guidelines for Inorganic Compounds in Water at 25°C” was created by Jackie MacDonald,

Using the Solubility Table Guidelines to Determine Solubility

To qualitatively determine the solubility of a salt in aqueous solution, reference Table 15.3a and follow the steps outline below:

1. Identify the negative ion in the salt; locate the ion in the first column of Table 15.3a.
2. Determine the positive ion in the salt by moving into the second column, same row as your anion.
3. Finally, move horizontally in that row into the solubility column (column 3) to determine whether that compound is soluble or insoluble.
4. Identify the symbol used to represent its solubility

Example 15.3a

Use the solubility rules to determine the solubility of the following substances. Categorize each as soluble or insoluble and also write the symbol used to show its solubility.

1. Na_2S
2. AgBr
3. $Mg(Cl)_2$

Solution

1. Soluble, (*aq*) – Since S is not listed as an anion, it falls into the "all" option. It's cation is a Na^+ cation, so it is soluble, *aq*.
2. Insoluble, (*s*) – When Br^- forms a salt with Ag^+, it forms an insoluble salt, *s*.
3. Soluble, (*aq*) – When Cl^- forms a salt with Mg^{2+}, it fall into the "all others" category for cations, and it forms a soluble salt, *aq*.

Source: Example 15.3a created by Jackie MacDonald and David McCuaig, CC BY-NC-SA 4.0.

Exercise 15.3a

Check Your Learning Exercise (Text Version)
From the options provided, identify the salts that will form a precipitate in aqueous solution at 25 degrees Celsius.

1. $Fe(NO_3)_2$
2. $BaSO_4$
3. AgF
4. $Mg(OH)_2$

5. $AgNO_2$
6. $ZnCl_2$
7. PbI_2
8. KBr
9. $LiClO_3$
10. AgBr
11. KOH
12. $(NH_4)PO_4$
13. $CaCO_3$

Check Your Answers[1]

Source: "Exercise 15.3a" created by Jackie MacDonald and David McCuaig, licensed under CC BY-NC-SA 4.0.

Watch the video Precipitation Reactions (https://youtu.be/hVBsrwJFBTY?t=391) starting at 6min 31sec until 9min05sec. It demonstrates how to use a similar solubility table to determine solubility.

For a more in depth solubility table that is online and printable, link to "Solubility Rules Chart" by MilliporeSigma.

Source: Section titled "How to Use the Solubility Table Guidelines to Determine Solubility" was created by Jackie MacDonald and David McCuaig and is licensed under CC BY-NC-SA 4.0.

Precipitation Reactions

A **precipitation reaction** is one in which dissolved substances react to form one (or more) solid products. Many reactions of this type involve the exchange of ions between ionic compounds in aqueous solution and are sometimes referred to as double displacement, double replacement, or metathesis reactions. These reactions are common in nature and are responsible for the formation of coral reefs in ocean waters and kidney stones in animals. They are used widely in industry for production of a number of commodity and specialty chemicals. Precipitation reactions also play a central role in many chemical analysis techniques, including spot tests used to identify metal ions and gravimetric methods for determining the composition of matter.

A vivid example of precipitation is observed when aqueous solutions of potassium iodide and lead nitrate are mixed, resulting in the formation of solid lead iodide:

$$2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

This observation is consistent with the solubility guidelines: The only insoluble compound among all those involved is lead iodide, one of the exceptions to the general solubility of iodide salts. Lead iodide is a bright yellow solid that was formerly used as an artist's pigment known as iodine yellow (Figure 15.3a). The properties of pure PbI_2 crystals make them useful for fabrication of X-ray and gamma ray detectors.

Figure 15.3a Formation of Precipitate Lead(II) Iodide during a Precipitation Reaction: The precipitation reaction producing Lead(II) iodide is shown. It is known as "golden rain" because of the yellow hexagonal crystals forming throughout in the aqueous solution and the solid crystals settle at the bottom of the beaker. This picture was taken after cooling a heated lead(II) nitrate and potassium iodide solution on a Bunsen burner. (credit: work by Der Kreole, CC BY-SA 3.0)

The solubility table may be used to predict whether a precipitation reaction will occur when solutions of soluble ionic compounds are mixed together. One merely needs to identify all the ions present in the solution and then consider if possible cation/anion pairing could result in an insoluble compound.

For example, mixing aqueous solutions of silver nitrate and sodium chloride will yield a solution containing Ag^+, NO_3^-, Na^+, and Cl^- ions. Aside from the two ionic compounds originally present in the solutions, $AgNO_3$ and NaCl, two additional ionic compounds may be derived from this collection of ions: $NaNO_3$ and AgCl. The solubility table can be used to determine if either of these salt combinations are insoluble in aqueous solution. Insoluble salts will precipitate out of the solution to form a solid (*s*).

The solubility table indicates all nitrate salts are soluble, so sodium nitrate ($NaNO_3$) will remain ions in solution. However, silver chloride, AgCl, is one of the exceptions to the general solubility rules of chloride salts, and this combination of ions will form a solid in aqueous solution. Therefore, a precipitation reaction is predicted to occur, as described by the following molecular equation:

$$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

Watch Precipitation Reactions (10mins 13sec) (https://youtu.be/hVBsrwJFBTY).

Example 15.3b

Predict the result of mixing reasonably concentrated solutions of the following ionic compounds. If a reaction occurs and a precipitation is expected, write a balanced molecular equation for the following two reactions.

a. potassium sulfate and barium nitrate
b. lithium chloride and silver acetate)

Solution

a. The two possible products for this combination are KNO_3 and $BaSO_4$. The solubility guidelines indicate KNO_3 is soluble and $BaSO_4$ is insoluble, so a precipitation reaction is expected. The balanced molecular equation for this reaction is $K_2SO_4(aq) + Ba(NO_3)_2\ (aq) \rightarrow BaSO_4(s) + 2KNO_3(aq)$
b. The two possible products for this combination are $LiC_2H_3O_2$ and AgCl. The solubility guidelines indicate $LiC_2H_3O_2$ is soluble and AgCl is insoluble, and so a precipitation reaction is expected. The balanced molecular equation for this reaction is $LiCl\ (aq) + AgC_2H_3O_2(aq) \rightarrow AgCl(s) + LiC_2H_3O_2(aq)$

Exercise 15.3b

Predict the result of mixing reasonably concentrated aqueous solutions of the following ionic compounds. If a reaction occurs and a precipitation is expected, write a balanced molecular equation for the following reaction between lead(II) nitrate and ammonium carbonate.

Check Your Answer[2]

Exercise 15.3c

Which solution(s) could be used to precipitate the barium ion, Ba^{2+}, in a aqueous solution barium nitrate: sodium chloride, sodium hydroxide, or sodium sulfate? If a reaction occurs and a precipitation is expected, write a balanced molecular equation for that reaction

Check Your Answer[3]

Links to Interactive Learning Tools

Explore Precipitation Reactions (https://www.physicsclassroom.com/Concept-Builders/Chemistry/Precipitation-Reactions) from the Physics Classroom (https://www.physicsclassroom.com).

Attribution & References

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Notes

1. The salts that form precipitates are PbI_2, $AgBr$, $CaCO_3$, $Mg(OH)_2$, $BaSO_4$, $AgNO_2$
2. The two possible products for this combination are $PbCO_3$ and NH_4NO_3. The solubility guidelines indicate NH_4NO_3 is soluble and $PbCO_3$ is insoluble, and so a precipitation reaction is expected. The balanced molecular equation for this reaction is $Pb(NO_3)_2(aq) + (NH_4)_2CO_3(aq) \rightarrow PbCO_3(s) + 2NH_4NO_3(aq)$
3. The solubility table can be used to determine which ionic compounds containing barium will be insoluble in aqueous solution. There are three possibilities to consider: $BaCl_2$, $Ba(OH)_2$, and $BaSO_4$. The solubility guidelines indicate $BaCl_2$ and $Ba(OH)_2$ are soluble in water and will not form a precipitate; however, $BaSO_4$ is insoluble, so this combination will result in a precipitation reaction. The balanced molecular equation for this reaction is

$$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$$

15.4 DESCRIBING REACTIONS IN SOLUTIONS BY WRITING MOLECULAR, COMPLETE IONIC, AND NET IONIC EQUATIONS

Learning Objectives

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

- Write and balance chemical equations in molecular, complete ionic, and net ionic formats.

Ionic Compounds in Solution

We have learned that one important aspect about ionic compounds that differs from molecular compounds has to do with dissolving in a liquid, such as water. When molecular compounds, such as sugar, dissolve in water, the individual molecules drift apart from each other. When ionic compounds dissolve, the ions physically separate from each other. We can use a chemical equation to represent this process—for example, with NaCl:

$$NaCl(s) \xrightarrow{H_2O(l)} Na^{+}(aq) + Cl^{-}(aq)$$

Figure 15.4a Chemical Equation Illustrating Sodium Chloride Salt Ionizing in Water: NaCl will dissolve in water; each atom separates into their individual ions in the water solution. Na^+ and Cl^- ions disperse throughout the water and the salt has dissolved or ionized in the aqueous solution. (credit: work by Shirley Wacowich-Sgarbi, CC BY-NC-SA 4.0; / Adapted by Jackie MacDonald).

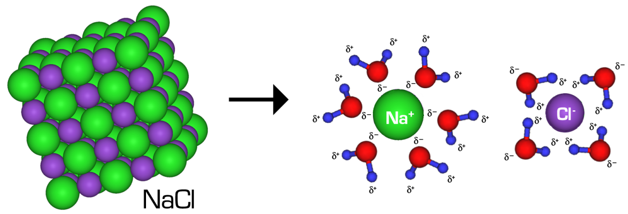

Figure 15.4b Water Hydration of a Crystal of Salt (NaCl): When immersed in water, the attractive forces of the water molecules to the NaCl separates the NaCl molecule to disperse into its separate ions. The negatively-charged oxygens of the water molecules are attracted to the positively-charged sodium ions. The positively-charged side of the water molecules (the hydrogen atoms) are attracted to the negatively-charged chloride ions. Water molecules pull the sodium and chloride ions apart, breaking the ionic bond that held them together. After the salt compounds are pulled apart, the sodium and chloride atoms are surrounded by water molecules. Once this happens, the salt is dissolved, resulting in a homogeneous solution. (credit: work by Andrea Hazard, CC-BY-SA 4.0)

When NaCl dissolves in water, the ions separate and go their own way in solution; the ions are now written with their respective charges, and the (aq) phase label emphasizes that they are dissolved (Figure 15.4a and Figure 15.4b). When an ionic compound dissociates in water, water molecules surround each ion and separate it from the rest of the solid. Each ion goes its own way in solution.

All ionic compounds that dissolve behave this way. (This behaviour was first suggested by the Swedish chemist Svante August Arrhenius [1859–1927] as part of his PhD dissertation in 1884. Interestingly, his PhD examination team had a hard time believing that ionic compounds would behave like this, so they gave Arrhenius a barely passing grade. Later, this work was cited when Arrhenius was awarded the Nobel Prize in Chemistry.)

Keep in mind that when the ions separate, *all* of the ions separate. Thus, when $CaCl_2$ dissolves, the one Ca^{2+} ion and the two Cl^- ions separate from each other:

$$CaCl_2(s) \rightarrow Ca^{2+}(aq) + Cl^-(aq) + Cl^-(aq)$$

which is simplified to

$$CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$$

That is, the two chloride ions go off on their own. They do not remain as Cl_2 (that would be elemental chlorine; these are chloride ions); they do not stick together to make Cl_2^- or Cl_2^{2-}. They become dissociated ions in their own right. It is important to note that polyatomic ions retain their overall identity when they are dissolved. It is important to understand and write the chemical equation that represents the dissociation of ions, as it is a crucial concept when writing complete ionic chemical equations, and net ionic equations.

Example 15.4a

Write the chemical equation that represents the dissociation of each ionic compound.

a. KBr
b. Na_2SO_4
c. $(NH_4)_3PO_4$

Solution

a. $KBr(s) \rightarrow K^+(aq) + Br^-(aq)$
b. Not only do the two sodium ions go their own way, but the polyatomic sulfate ion stays together as the sulfate ion. The dissolving equation is $Na_2SO_4(s) \rightarrow 2Na^+(aq) + SO_4^{2-}(aq)$
c. Not only do the three ammonium ions stay together, but the one phosphate ion stays together as well since they are both polyatomic ions. The dissolving equation is $(NH_4)_3PO_4(s) \rightarrow 3NH_4^+(aq) + PO_4^{3-}(aq)$

Exercise 15.4a

Exercise 15.4a.1

1) Write the chemical equation that represents the dissociation of $(NH_4)_2S$.

Check Your Answer[1]

Exercise 15.4a.2

2) Write the chemical equation that represents the dissociation of iron(III) sulfate, $Fe_2(SO_4)_3$.

Check Your Answer[2]

Equations for Ionic Reactions

Given the abundance of water on earth, it stands to reason that a great many chemical reactions take place in aqueous media. When ions are involved in these reactions, the chemical equations may be written with various levels of detail appropriate to their intended use. To illustrate this, consider a reaction between ionic compounds taking place in an aqueous solution. When aqueous solutions of $CaCl_2$ and $AgNO_3$ are mixed, a precipitation reaction takes place producing aqueous $Ca(NO_3)_2$ and solid AgCl as shown in the following molecular equation:

$$CaCl_2(aq) + 2AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$$

This balanced equation, derived by means describe in previous sections, is called a **molecular equation** because it doesn't explicitly represent the ionic species that are present in solution. When ionic compounds dissolve in water, they may dissociate into their constituent ions, which are subsequently dispersed homogenously throughout the resulting solution. Ionic compounds dissolved in water are, therefore, more realistically represented as dissociated ions, in this case:

$$CaCl_2(aq) \rightarrow Ca^{2+}(aq) + 2Cl^{-}(aq)$$
$$2AgNO_3(aq) \rightarrow 2Ag^{+}(aq) + 2NO_3^{-}(aq)$$
$$Ca(NO_3)_2(aq) \rightarrow Ca^{2+}(aq) + 2NO_3^{-}(aq)$$

Unlike these three ionic compounds, AgCl does not dissolve in water to a significant extent, as signified by its physical state notation, *s*.

Explicitly representing all dissolved ions results in a **complete ionic equation**. In this particular case, the formulas for the dissolved ionic compounds are replaced by formulas for their dissociated ions:

$$Ca^{2+}(aq) + 2Cl^{-}(aq) + 2Ag^{+}(aq) + 2NO_3^{-}(aq) \rightarrow Ca^{2+}(aq) + 2NO_3^{-}(aq) + 2AgCl(s)$$

NOTE: the order in which the products are written, does not matter. The solid may be written first or second. Examining this complete ionic equation shows that two chemical species are present in identical forms on both sides of the arrow, $Ca^{2+}(aq)$ and $NO_3^{-}(aq)$. These **spectator ions**—ions whose presence is required to maintain charge neutrality—are neither chemically nor physically changed by the process, and so they may be eliminated from the equation to yield a more concise representation called a **net ionic equation**. The net ionic equation summarizes the chemical change that occurred in the chemical reaction. Below, the complete ionic equation is shown first. The spectator ions ($Ca^{2+}(aq)$ and $2NO_3^{-}(aq)$) are bolded in blue. Then, the other reactants and products are included in the net ionic equation, which is shown second.

Complete Ionic Equation: $\mathbf{Ca^{2+}(aq)} + 2Cl^{-}(aq) + 2Ag^{+}(aq) + \mathbf{2NO_3^{-}(aq) \rightarrow Ca^{2+}(aq)} + \mathbf{2NO_3^{-}(aq)} + 2AgCl(s)$

Net Ionic Equation: $2Cl^{-}(aq) + 2Ag^{+}(aq) \rightarrow 2AgCl(s)$

which can be simplified to the smallest possible integers as coefficients and written as

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

This net ionic equation indicates that solid silver chloride may be produced from dissolved chloride and silver ions, regardless of the source of these ions. These molecular and complete ionic equations provide additional information, namely, the ionic compounds used as sources of Cl^{-} and Ag^{+}. The above reaction is a precipitation reaction.

Watch How to Write and Balance Net Ionic Equations (6min 18sec) (https://youtu.be/PXRH_IrN11Y).

Example 15.4b

The molecular equations for two chemical reactions are given below: Write the complete ionic equation for each chemical reaction.

a. $KBr(aq) + AgC_2H_3O_2(aq) \rightarrow KC_2H_3O_2(aq) + AgBr(s)$
b. $MgSO_4(aq) + Ba(NO_3)_2(aq) \rightarrow Mg(NO_3)_2(aq) + BaSO_4(s)$

Solution

For any ionic compound that is aqueous, we will write the compound as separated ions.

a. The complete ionic equation is
   $K^{+}(aq) + Br^{-}(aq) + Ag^{+}(aq) + C_2H_3O_2^{-}(aq) \rightarrow K^{+}(aq) + C_2H_3O_2^{-}(aq) + AgBr(s)$

b. The complete ionic equation is
$Mg^{2+}(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + BaSO_4(s)$

Exercise 15.4b

Write the complete ionic equation for the following chemical chemical reaction:

$$CaCl_2(aq) + Pb(NO_3)_2(aq) \rightarrow Ca(NO_3)_2(aq) + PbCl_2(s)$$

Check Your Answer[3]

Example 15.4c

Write the net ionic equation for each chemical reaction below. Identify the spectator ions in each of these reactions.

a. $K^+(aq) + Br^-(aq) + Ag^+(aq) + C_2H_3O_2^-(aq) \rightarrow K^+(aq) + C_2H_3O_2^-(aq) + AgBr(s)$
b. $Mg^{2+}(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2\ NO_3^-(aq) \rightarrow Mg^{2+}(aq) + 2\ NO_3^-(aq) + BaSO_4(s)$

Solution

a. In the first equation, the $K^+(aq)$ and $C_2H_3O_2^-(aq)$ ions are spectator ions (bolded in blue), so they are cancelled.
$\mathbf{K^+(aq)} + Br^-(aq) + Ag^+(aq) + \mathbf{C_2H_3O_2^-(aq)} \rightarrow \mathbf{K^+(aq)} + \mathbf{C_2H_3O_2^-(aq)} + AgBr(s)$;
The net ionic equation is
$Br^-(aq) + Ag^+(aq) \rightarrow AgBr(s)$
b. In the second equation, the $Mg^{2+}(aq)$ and $NO_3^-(aq)$ ions are spectator ions (bolded in blue), so they are cancelled.
$\mathbf{Mg^{2+}(aq)} + SO_4^{2-}(aq) + Ba^{2+}(aq) + \mathbf{2NO_3^-(aq)} \rightarrow \mathbf{Mg^{2+}(aq)} + \mathbf{2NO_3^-(aq)} + BaSO_4(s)$;
The net ionic equation is
$SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$

Exercise 15.4c

Write the net ionic equation for the chemical reaction below. Identify the spectator ions in this reaction.

$Ca^{2+}(aq) + 2Cl^{-}(aq) + Pb^{2+}(aq) + 2NO_3^{-}(aq) \rightarrow Ca^{2+}(aq) + 2NO_3^{-}(aq) + PbCl_2(s)$

Check Your Answer[4]

Example 15.4d

When aqueous barium chloride is mixed with an aqueous solution of sodium sulfate, the mixture reacts to yield solid barium sulfate and aqueous sodium chloride. Write balanced molecular, complete ionic, and net ionic equations for this process. Name the spectator ions.

Solution

Begin by identifying formulas for the reactants and products and arranging them properly in chemical equation form:

$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + NaCl(aq)$ (unbalanced equation)

Balance is achieved easily in this case by changing the coefficient for NaCl to 2, resulting in the molecular equation for this reaction:

$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$ (balanced equation)

The dissolved ionic compounds, $BaCl_2$, Na_2SO_4, and NaCl, can be represented as dissociated ions to yield the complete ionic equation:

$Ba^{2+}(aq) + 2Cl^{-}(aq) + 2Na^{+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + 2Na^{+}(aq) + 2Cl^{-}(aq)$

Finally, identify the spectator ion(s), in this case $Na^{+}(aq)$ and $Cl^{-}(aq)$, and remove them from each side of the equation to generate the net ionic equation:

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

Exercise 15.4d

When an aqueous solution of $AgNO_3$ is added to an aqueous solution of $CaCl_2$, insoluble AgCl precipitates. Write three equations (complete molecular equation, complete ionic equation, and net ionic equation) that describe this process.

Check Your Answer[5]

Net Ionic Equations of Gas Forming Reactions

Sometimes a gas will be involved as one of the reactants or products in a solution reaction. Net ionic equations can be written for these chemical reactions, as well. For example,

$$2HCl(aq) + Na_2S(aq) \rightarrow H_2S(g) + 2NaCl(aq)$$

In this example, since hydrochloric acid fully dissociates in aqueous solution, the complete ionic equation would be:

$$2H^+(aq) + 2Cl^-(aq) + 2Na^+(aq) + S^{2-}(aq) \rightarrow H_2S(g) + 2Na^+(aq) + 2Cl^-(aq)$$

Removing the spectator ions we obtain the net ionic equation:

$$2H^+(aq) + S^{2-}(aq) \rightarrow H_2S(g)$$

Exercise 15.4e

When carbon dioxide gas combines with an aqueous solution of sodium hydroxide, the mixture reacts to yield aqueous sodium carbonate and liquid water. Write balanced molecular, complete ionic, and net ionic equations for this process.

Check Your Answer[6]

Net Ionic Equations of Single Replacement Reactions

Involving Aqueous Solutions

Recall, that during a **single replacement reaction** (SR) there is an exchange of elements; typically cations; an element becomes a compound and a compound becomes an element. SR reactions are always oxidation reduction (redox) reactions. These types of reactions will be discussed more in the Oxidation Reduction chapter, but we will cover the basics of writing net ionic equations for redox reactions undergoing single replacement.

Consider the following SR reaction between iron metal and copper(II) nitrate:

Balanced Molecular Equation: $Fe(s) + Cu(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$

Complete Ionic Equation: $Fe(s) + Cu^{2+}(aq) + 2NO_3^-(aq) \rightarrow Fe^{2+}(aq) + 2NO_3^-(aq) + Cu(s)$

Since nitrate ions are dissociated in the reactant and product side of the complete ionic equation, these are removed to produce the net ionic equation.

Net Ionic Equation: $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

Exercise 15.4f

Write the complete ionic and net ionic equations for the following chemical reaction.

$Ni(s) + 2HCl(aq) \rightarrow NiCl_2(aq) + H_2(g)$

Check Your Answer[7]

Net Ionic Equations of Neutralization Reactions

A **neutralization reaction** is a type of double displacement/replacement reaction. There is a double exchange of two cations, or two anions to form two new compounds. Neutralization reactions will be discussed more in the Acids and Bases chapter, but we will cover the basics of writing net ionic equations for neutralization reactions, now. In a neutralization reaction, the reactants are an acid and a base, and the products are often a salt (soluble or insoluble) and water, and neither reactant is the water itself:

acid + base → salt + water

Consider the following reaction between a strong acid, hydrochloric acid, and strong base, sodium hydroxide. The balanced molecular equation can be used to generate a complete ionic equation, and its net ionic equation:

Balanced Molecular Equation: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

Complete Ionic Equation: $H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$

Since nitrate ions are dissociated in the reactant and product side of the complete ionic equation, these are removed to produce the net ionic equation.

Net Ionic Equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Links to Interactive Learning Tools

Explore Net Ionic Equations (https://www.physicsclassroom.com/Concept-Builders/Chemistry/Precipitation-Reactions) from the Physics Classroom (https://www.physicsclassroom.com).

Attribution & References

Except where otherwise noted, this page is adapted from "4.1 Writing and Balancing Chemical Equations (https://pressbooks.bccampus.ca/chem1114langaracollege/chapter/4-1-writing-and-balancing-chemical-equations/)" In *CHEM 1114 – Introduction to Chemistry* by Shirley Wacowich-Sgarbi and Langara Chemistry Department is licensed under CC BY-NC-SA 4.0. Adaptations and additions to content in this section were made by Jackie MacDonald for student comprehension.

Notes

1. $(NH_4)_2S\,(s) \rightarrow 2NH_4^+(aq) + S^{2-}(aq)$
2. $Fe_2(SO_4)_3\,(s) \rightarrow 2Fe^{3+}(aq) + 3SO_4^{2-}(aq)$
3. Complete ionic equation is $Ca^{2+}(aq) + 2Cl^-(aq) + Pb^{2+}(aq) + 2NO_3^-(aq) \rightarrow Ca^{2+}(aq) + 2NO_3^-(aq) + PbCl_2(s)$
4. $Ca^{2+}(aq)$ and $NO_3^-(aq)$ ions are spectator ions (bolded in blue), so they are cancelled. $\mathbf{Ca^{2+}(aq)} + 2Cl^-(aq) + Pb^{2+}(aq) + \mathbf{2NO_3^-(aq)} \rightarrow \mathbf{Ca^{2+}(aq)} + \mathbf{2NO_3^-(aq)} + PbCl_2(s)$; The net ionic equation is $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$
5. Begin by identifying formulas for the reactants and products and arranging them properly in chemical equation form: $AgNO_3(aq) + CaCl_2(aq) \rightarrow AgCl(s) + Ca(NO_3)_2(aq)$ (unbalanced equation) Balancing this reaction results in the molecular equation for this reaction: $2AgNO_3(aq) + CaCl_2(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$ (balanced equation) The dissolved ionic compounds, $AgNO_3$, $CaCl_2$, and $Ca(NO_3)_2$, can be represented as dissociated ions to yield the complete ionic equation: $2Ag^+(aq) + 2NO_3^-(aq) + Ca^{2+}(aq) + 2Cl^-(aq) \rightarrow 2AgCl(s) + Ca^{2+}(aq) + 2NO_3^-(aq)$ Finally, identify the spectator ion(s), in this case $Ca^{2+}(aq)$ and $NO_3^-(aq)$, and remove them from each side of the equation to generate the net ionic equation: $2Ag^+(aq) + 2Cl^-(aq) \rightarrow 2AgCl(s)$ which can be simplified to the smallest possible integers as coefficients and written as $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
6. Begin by identifying formulas for the reactants and products and arranging them properly in chemical equation form:

$CO_2(g) + NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$ (unbalanced) Balance is achieved easily in this case by changing the coefficient for NaOH to 2, resulting in the molecular equation for this reaction: $CO_2(g) + 2NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$ (balanced) The two dissolved ionic compounds, NaOH and Na_2CO_3, can be represented as dissociated ions to yield the complete ionic equation: $CO_2(g) + 2Na^+(aq) + 2OH^-(aq) \rightarrow 2Na^+(aq) + CO_3^{2-}(aq) + H_2O(l)$ Finally, identify the spectator ion(s), in this case $Na^+(aq)$, and remove it from each side of the equation to generate the net ionic equation: $CO_2(g) + 2OH^-(aq) \rightarrow CO_3^{2-}(aq) + H_2O(l)$ To review a video of this solution watch **How to Write the Net Ionic Equation for NaOH + CO_2 = Na_2CO_3 + H_2O (2mins 38sec).** (https://youtu.be/QKXeIdCVqYQ)

7. Complete Ionic Equation: $Ni(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Ni^{2+}(aq) + 2Cl^-(aq) + H_2(g)$ Since chloride ions are dissociated in both the reactant and product side of the complete ionic equation, these are removed to produce the net ionic equation.

   Net Ionic Equation: $Ni(s) + 2H^+(aq) \rightarrow Ni^{2+}(aq) + H_2(g)$

CHAPTER 15 - SUMMARY

15.1 Salts

Salts are a chemical compound formed when ions form ionic bonds. In these reactions, one atom gives up one or more electrons, and thus becomes positively charged, whereas the other accepts one or more electrons and becomes negatively charged; overall the ionic compound has no net charge. Salts typically are crystalline, odourless, colourless/transparent or white, and have high melting and boiling points. Many salts are soluble in water; however, some are not. If a given salt is soluble in water, it completely dissociates into ions other than a hydrogen ion (H^+) or hydroxide ion ($OH^−$) and forms an aqueous solution. This fact is elemental in distinguishing salts from acids and bases. Salts are derived from the neutralization reaction of an acid and base. Since acids and bases always contain either a metal cation or a cation derived from ammonium (NH_4^+) and a nonmetal anion, respectively, the two can combine to form a salt.

15.2 Electrolytes

Substances that dissolve in water to yield ions are called electrolytes. Electrolytes may be covalent compounds that chemically react with water to produce ions (for example, acids and bases), or they may be ionic compounds that dissociate to yield their constituent cations and anions, when dissolved. Dissolution of an ionic compound is facilitated by ion-dipole attractions between the ions of the compound and the polar water molecules. Soluble ionic substances and strong acids ionize completely and are strong electrolytes, while weak acids and bases ionize to only a small extent and are weak electrolytes. Nonelectrolytes are substances that do not produce ions when dissolved in water.

15.3 Precipitation Reactions

Chemical reactions are classified according to similar patterns of behaviour. Precipitation is one type of chemical reaction which involves the formation of one or more insoluble products. Precipitation reactions, also called double displacement reactions can be summarized with the following reaction equation:

$AB(aq) + CD\ (aq) \rightarrow AD(s)+CB(aq)$ or (s)

The formation of the solid from combining two aqueous solutions is the DRIVING FORCE of the reaction (the factor that makes the reaction go). A precipitation reaction can be predicted to occur with the help of a solubility table.

15.4 Describing Reactions in Solutions by Writing Molecular, Complete Ionic, and Net Ionic Equations

Chemical equations are symbolic representations of chemical and physical changes. Chemical reactions in aqueous solution that involve ionic reactants or products may be represented more realistically by complete ionic equations and, more succinctly, by net ionic equations. Complete ionic and net ionic equations can be used to illustrate what is happening during precipitation reactions, neutralization reactions, gas evolving reactions, and single replacement reactions when these reactions occur in aqueous solutions.

Attributions and References

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Adaptations to aid in student comprehension by Jackie MacDonald.

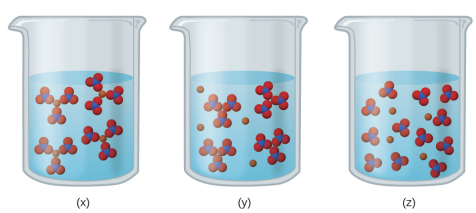
CHAPTER 15 - REVIEW

15.1 Salts

1. Explain how a salt is formed.
   Check answers: [1]
2. What salt will form when you combine hydrobromic acid and lithium hydroxide?
   Check answers: [2]
3. What salt will form when you combine sulfuric acid, H_2SO_4, and sodium hydroxide?
   Check answers: [3]

15.2 Electrolytes

1. Differentiate between strong and weak electrolytes.
   Check Answers: [4]
2. Explain why the ions Na^+ and Cl^- are strongly solvated in water but not in hexane, a solvent composed of nonpolar molecules. **Check Answer:** [5]
3. Explain why solutions of HBr in benzene (a nonpolar solvent) are nonconductive, while solutions in water (a polar solvent) are conductive.
4. Consider the solutions presented:
   a. Which of the following sketches, shown in the figure below, best represents the ions in a solution of $Fe(NO_3)_3(aq)$?



   b. Write a balanced chemical equation showing the products of the dissolution of $Fe(NO_3)_3$.
   Check Answer: [6]
5. Compare the processes that occur when methanol (CH_3OH), hydrogen chloride (HCl), and sodium hydroxide (NaOH) dissolve in water. Write equations and prepare sketches showing the form in which each of these compounds is present in its respective solution.

6. What is the expected electrical conductivity of the following solutions?
   a. $NaOH(aq)$
   b. $HCl(aq)$
   c. $C_6H_{12}O_6(aq)$ (glucose)
   d. $NH_3(l)$
   Check Answer: [7]
7. Why are most *solid* ionic compounds electrically nonconductive, whereas aqueous solutions of ionic compounds are good conductors? Would you expect a *liquid* (molten) ionic compound to be electrically conductive or nonconductive? Explain.
8. Indicate the most important type of intermolecular attraction responsible for solvation in each of the following solutions:
   a. the solutions in Figure 14.2c
   b. methanol, CH_3OH, dissolved in ethanol, C_2H_5OH
   c. methane, CH_4, dissolved in benzene, C_6H_6
   d. the polar halocarbon CF_2Cl_2 dissolved in the polar halocarbon $CF_2ClCFCl_2$
   e. $O_2(l)$ in $N_2(l)$
   Check Answer: [8]

15.3 Precipitation Reactions

1. What are the general characteristics that help you recognize double replacement reactions?
   Check Answers: [9]
2. What are the general characteristics that help you recognize a precipitation reaction?
   Check Answers: [10]
3. Assuming that the following is a precipitation reaction, determine the products (and identify their phases, *aq* or *s*) and write the balanced chemical equation.
   $Zn(NO_3)_2 + NaOH \rightarrow ?$
   Check Answers: [11]
4. Assuming that the following is a precipitation reaction, determine the products (and identify their phases, *aq* or *s*) and write the balanced chemical equation.
   $MgCl_2 + NaOH \rightarrow$
   Check Answers: [12]
5. Use the solubility table to predict if the following double replacement reaction will occur and, if so, write a balanced chemical equation. $Pb(NO_3)_2 + KBr \rightarrow ?$
   Check Answers: [13]
6. Use the solubility table to predict if the following double replacement reaction will occur and, if so, write a balanced chemical equation.

$KCl + Na_2CO_3 \rightarrow ?$
Check Answers:[14]

7. Which solution could be used to precipitate the barium ion, Ba^{2+}, in a water sample: sodium chloride, sodium hydroxide, or sodium sulfate? What is the formula for the expected precipitate?
Check Answers: [15]

15.4 Net Ionic Equations

1. From the balanced molecular equations, write the complete ionic and net ionic equations for the following:
(a) $K_2C_2O_4(aq) + Ba(OH)_2 \rightarrow 2KOH(aq) + BaC_2O_2(s)$
(b) $Pb(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow PbSO_4(s) + 2HNO_3(aq)$
(c) $CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$
Check Answers: [16]
2. Predict the result of mixing reasonably concentrated solutions of the following ionic compounds. If precipitation is expected, write a balanced net ionic equation for the reaction.
(a) potassium sulfate and barium nitrate
(b) lithium chloride and silver acetate
(c) lead nitrate and ammonium carbonate
Check Answers: [17]
3. From the balanced molecular equation, write the net ionic equations for the following reaction:
$2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + H_2O(l)$
Check Answers: [18]
4. From the balanced molecular equation, write the net ionic equations for the following reaction:
$2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$
Check Answers: [19]

Attribution & References

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Notes

1. A salt is formed in a neutralization reaction of an acid and a base. Acids and bases always contain either a metal cation or a cation derived from ammonium (NH_4^+) and a nonmetal anion, the two can combine to form a salt.
2. the anion from the HBr is Br^-. The cation from LiOH is Li^+. The salt formed is LiBr, lithium bromide.
3. the anion from the H_2SO_4 is SO_4^{2-}. The cation from NaOH is Na^+. The salt formed is Na_2SO_4, sodium sulfate.
4. A strong electrolyte is a salt that fully ionizes (100%) in the aqueous solution and are great conductors of electricity. Whereas, with a weak electrolyte only a relatively small fraction of the dissolved substance undergoes the ion-producing process and therefore is a poor conductor of electricity.
5. Crystals of NaCl dissolve in water, a polar liquid with a very large dipole moment, and the individual ions become strongly solvated. Hexane is a nonpolar liquid with a dipole moment of zero and, therefore, does not significantly interact with the ions of the NaCl crystals.
6. (a) $Fe(NO_3)_3$ is a strong electrolyte, thus it should completely dissociate into Fe^{3+} and ($NO_3^{\ -}$) ions. Therefore, (z) best represents the solution. (b) $Fe(NO_3)_3(s) \longrightarrow Fe^{3+}(aq) + 3NO_3^{\ -}(aq)$
7. (a) high conductivity (solute is an ionic compound that will dissociate when dissolved); (b) high conductivity (solute is a strong acid and will ionize completely when dissolved); (c) nonconductive (solute is a covalent compound, neither acid nor base, unreactive towards water); (d) low conductivity (solute is a weak base and will partially ionize when dissolved)
8. (a) ion-dipole; (b) hydrogen bonds; (c) dispersion forces; (d) dipole-dipole attractions; (e) dispersion forces
9. A double replacement reaction occurs when parts of two ionic compounds are exchanged, making two new compounds. A characteristic of a double-replacement equation is that there are two compounds as reactants and two different compounds as products.
10. In a precipitation reactions, the reactants will be dissolved substances in aqueous solutions and will react to form one (or more) solid products. Many reactions of this type involve the exchange of ions between ionic compounds in aqueous solution and a type of double displacement/replacement reaction.
11. $Zn(NO_3)_2(aq) + 2NaOH(aq) \rightarrow 2NaNO_3(aq) + Zn(OH)_2(s)$
12. $MgCl_2(aq) + 2NaOH(aq) \rightarrow 2NaCl(aq) + Mg(OH)_2(s)$
13. The reaction will occur since a precipitate will form. $Pb(NO_3)_2(aq) + 2KBr(aq) \rightarrow 2KNO_3(aq) + PbBr_2(s)$
14. No reaction occurs; both products are soluble in water. $2KCl(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + K_2CO_3(aq)$
15. sodium sulfate, $BaSO_4$
16. (a) complete ionic: $2K^+(aq) + C_2O_4^{2-}(aq) + Ba^{2+}(aq) + 2OH^-(aq) \rightarrow 2K^+(aq) + 2OH^-(aq) + BaC_2O_4(s)$ net ionic: $Ba^{2+}(aq) + C_2O_4^{2-}(aq) \rightarrow BaC_2O_4(s)$ (b) complete ionic: $Pb^{2+}(aq) + 2NO_3^-(aq) + 2H^+(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2H^+(aq) + 2NO_3^-(aq)$ net ionic: $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$ (c) complete ionic: $CaCO_3(s) + 2H^+(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$ net ionic: $CaCO_3(s) + 2H^+(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$

17. **(a)** The two possible products for this combination are KNO_3 and $BaSO_4$. The solubility guidelines indicate $BaSO_4$ is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction is$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$ **(b)** The two possible products for this combination are $LiC_2H_3O_2$ and AgCl. The solubility guidelines indicate AgCl is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction is$Ag^{+}(aq) + Cl^{-}(aq) \rightarrow AgCl(s)$ **(c)** The two possible products for this combination are $PbCO_3$ and NH_4NO_3. The solubility guidelines indicate $PbCO_3$ is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction is$Pb^{2+}(aq) + CO_3^{2-}(aq) \rightarrow PbCO_3(s)$
18. $H+(aq) + OH^{-}(aq) \rightarrow H_2O(l)$
19. $2Ag^{+}(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + 2Ag(s)$