Solutions and Reactions

Introduction ▶ Water is by far the most common and the most important liquid on our planet. It is vital to all living organisms, and it is indispensable to civilization. Water covers most of the earth and through its global cycle — the hydrologic cycle — permeates the atmosphere and the rocks. It shapes the face of the planet as it slowly dissolves and transports the land into the sea. Much of the importance of water lies in its ability to dissolve and suspend so many substrates and then transport them to different places. It is indeed the universal solvent for gases, liquids, and solids.

It is probably safe to say that all biochemical reactions in living systems require water as a solvent medium and often use water as a reactant or as a catalyst. For example, the human cardiovascular system is a marvelous multipurpose medium for the exchange and transport of dissolved gases, sugars, suspended organelles, and so on. The oceans, a global example, contain about 3.5% of dissolved inorganic solids, mostly in the form of hydrated cations and anions. Millions of species have evolved to make this salty solution the perfect medium for life. The billions of tiny bubbles continuously trapped in the crashing waves burst and fling microscopic droplets of sea solution into the atmosphere. The droplets evaporate, leaving solid particles that are so small that they remain in the air for long periods of time. These marine aerosols play a large part in controlling the weather because they act as cloud condensation nuclei, catalyzing the formation of the natural precipitates: rain, snow, and hail. The vast global ocean currents transport everything ranging from microscopic plankton to carbon dioxide and air pollutants, such as fluorocarbons, from one hemisphere to the other. At every scale, from the microscopic to the macroscopic, the solvent and fluid properties of water are essential to the removal of the toxic wastes generated by cells, organisms, people, factories, and cities.

Of equal importance are the substances that are not very soluble in water. In many biological systems, the fluid bag of cells is supported by a skeletal structure whose composition can vary from silicates in Radiolaria (tiny sea creatures) to calcium hydroxyphosphate in humans. The transport of hydrated calcium and phosphate ions in the complex system of bone and teeth formation is a marvel of solution chemistry. The survival of birds depends upon a fragile shell composed of tiny calcite crystals laid down on a light network of protein collagen. When things go awry and substances precipitate in the wrong places, serious health problems can occur. Cholesterol plaques in the blood stream can end the flow of life. Kidney, gall, and bladder stones and calcified deposits in the joints can cause severe
pain. Microscopic silica and asbestos and smoke particles of all kinds injure lung cells and can cause cancer and death. Much is now known about the solution chemistry of these complex biological processes, but there is still much to be learned. We are living in the modern era of microelectronics and photonics* in which the control of the flow of electrons and photons in water-insoluble solids is the basis for our communication and manufacturing machines. It is intriguing to speculate that the study of the chemistry and flow of ions in aqueous solution may lead to a revolution in the understanding of communication and evolution in living systems. Microionics may rival microelectronics in importance in the scientific future. In this series of experiments, you have the opportunity to study the fascinating chemistry of dissolution, reaction, and precipitation in aqueous solution.

**Background Chemistry**

A homogeneous mixture of two or more substances is called a *solution*. Usually, the substance that is present in a smaller amount is called the *solute*, and the substance present in the larger amount is called the *solvent*. Solutions may be gaseous (e.g., air), liquid (e.g., blood), and solid (e.g., steel).

Solutions are the places where chemical and biological reactions happen. Solutions are also the means by which substances are transported from one place to another. Liquid solutions are particularly important in geological, environmental, biological, and chemical systems; the universal solvent for these solutions is water. The solution is the medium that brings reactants close together long enough to allow new associations and bonds to form. Products appear as bond formation moves toward completion. The solvent almost always plays an active role in all of these processes, especially when water is the solvent. The solvent must first dissolve the substances that are eventually going to react. This deceptively simple process called *dissolution*, also involves the breaking of bonds and the formation of new associations. The solubility of a substance reflects how easily the solvent can make these changes occur. The *solubility* of a substance is, in fact, defined as the concentration of solute in a saturated solution at a specified temperature. The reverse of dissolution is the "coming out" of solution in which the product(s) of a reaction exceed the solubility and form a new phase that could be a gas, liquid, or solid. Again, the solvent often plays an active part in this process.

At the molecular level these dynamic aqueous solution processes can be pictured as shown in Figure 1.

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*The use of laser light in communication systems, sometimes termed optronics.*
where ◦ are solute and □ are water molecules. The extent to which each of these processes happens can be described in terms of three major interactions:

- Solute–solute interactions
  \[ ◦ \leftrightarrow ◦ \]

- Solute–water interactions
  \[ ◦ \leftrightarrow □ \]

- Water–water interactions
  \[ □ \leftrightarrow □ \]

If the solute–solute interaction is dominant, then solute will come out of solution and two phases will form. Which phase goes up (perhaps a gas) and which phase goes down (perhaps a solid) depends on external forces, such as gravity. If the water–water interaction is dominant, then the solute will be “squeezed out” of the solution and form another phase. In between these extremes, water–solute interactions can assure a stable, homogeneous solution.

Water is an excellent solvent for many ionic and polar covalent compounds because the water molecule is a polar covalent molecule. Each \( \text{H}_2\text{O} \) molecule has a small negative charge on the oxygen atom and a small positive charge on the hydrogen atoms (Refer to Figure 2 for illustration.)
This permanent dipole (two poles, a positive one and a negative one) is due to the electronegativity of the oxygen atom. The electrons in the covalent bonds between O and H are pulled slightly towards the O atom, creating the dipole. In liquid water (and ice), the water molecules are fairly close together and attract each other, forming weak bonds called hydrogen bonds, which are pictured in Figure 3.

Figure 3 Hydrogen Bonding in Liquid Water

Although these hydrogen bonds are weak, they are responsible for many of the interesting and unusual properties of water, including some of its versatile solvent properties. Solutes that interact with the charges on the water molecule are generally soluble provided that the solute–solute interactions are not very strong. Ammonia (NH₃), for example, is very soluble in water, forming hydrogen bonds in the process, whereas dinitrogen (N₂), a nonpolar molecule, is not (see Figure 4).

Figure 4 Hydrogen Bonding to NH₃ and not N₂

Large nonpolar substances, such as gasoline, are not very soluble in water because the long nonpolar hydrocarbon chains disrupt the hydrogen bonding between water molecules. Solvent–solute interactions dominate, and the gasoline molecules are “squeezed out” of solution and form a separate liquid floating on top of the water, as shown in Figure 5.
The solubility of solids in water is also governed by the three major interactions discussed earlier. A simple example is the dissolution of sodium nitrate (NaNO₃) in water. Solid sodium nitrate has a crystal structure that consists of sodium ions (Na⁺) ionically bonded to nitrate ions (NO₃⁻). Water is a good solvent for salts such as sodium nitrate because water molecules are able to move in between the cation and anion and screen the charges from each other. When this happens the ionic bond weakens, and the water molecules can then orient around the dissociated (separated) ions with the negative end of the water dipole pointed towards the cation and the positive end pointed towards the anion. The water molecules are bonded quite strongly to the ions by the attraction of unlike charges. All cations and anions dissolved in aqueous solution are surrounded by bound water molecules, and they are said to be hydrated. Figure 6 illustrates the process of dissolution and hydration.

Note: Most chemistry texts, including this one assume that the reader understands that cations and anions in aqueous solution exist as hydrated ions; therefore the formulas for ions are usually written without the bonded solvent molecules. The extent to which the
dissolution process occurs — i.e., how soluble the salt is — depends on the magnitude of the attraction between ions and water molecules (hydration energy) compared with the attraction between ions in the solid salt (lattice energy). The battle of forces within the crystal versus those between ions and water molecules is discussed in a later section on precipitation reactions.

The strength of the attraction between ions and water molecule dipoles depends on the charge and the radius of the ion. Small ions with high charge and electronegativity have the greatest attraction for water molecules; conversely, large ions with a single charge have the smallest attraction. If the attraction of the metal ion for the negative end of the water molecule dipole is strong enough, the water molecule may be ripped apart, releasing a hydrogen ion (H+) that is then hydrated. A good example is the aluminum ion (radius 67 pm and a +3 charge), which can react with water to produce the hydroxy cation [Al(H₂O)₆OH]³⁺ and the hydronium ion H₃O⁺:

\[
[\text{Al(H}_2\text{O)}_6\text{OH}]^{3+} + \text{H}_2\text{O} \rightarrow [\text{Al(H}_2\text{O)}_5\text{OH}]^{2+} + \text{H}_3\text{O}^+
\]

Compounds that react with water to produce hydronium ions are called acids. The aluminum ion is therefore called an acidic cation, whereas the sodium ion discussed earlier is called a nonacidic cation. This type of reaction can continue with more water molecules to produce hydrated cations containing several hydroxy groups. Eventually, these hydroxy cations can generate insoluble hydroxides — e.g., Al(OH)₃(aq) — that precipitate from solution.

Anions in aqueous solution also interact with water, except of course in this instance, the positive end of the water dipole is attracted to the anion. If the anion has a small size and high charge, then water molecules are pulled apart and the hydrogen atom bonds to the anion, releasing a hydroxide ion (OH⁻). A simple example is the reaction of a carbonate ion (CO₃²⁻) with water,

\[
\text{CO}_3^{2-} + \text{H}_2\text{O} \rightarrow \text{HCO}_3^- + \text{OH}^-
\]

which produces a bicarbonate ion (HCO₃⁻) and a basic (the opposite of acidic) solution. The carbonate ion is called a basic anion, whereas the nitrate ion (NO₃⁻) discussed earlier is called a nonbasic anion.

The chemistry of ionic solutes in aqueous solution becomes considerably more interesting and complicated when two solutions, each containing a soluble ionic compound, are mixed. Generally, in this situation the resulting solution, at the instant of mixing, contains two different hydrated cations and two different hydrated anions. If strong solute–solute interaction occurs between any two of the ions, then a chemical reaction results and a new product is formed. Often the two remaining ions do not interact and remain in solution in the same hydrated state as before the reaction. These unreacted ions are often called spectator ions. Chemical reactions in aqueous solution may be divided into four types: precipitation, acid–base, complexation, and redox reactions. A brief discussion of the main characteristics of each of these reaction types follows.
Precipitation Reactions

A precipitation reaction is defined as a reaction that produces a new compound that is not soluble in an aqueous solution. Precipitation can be regarded as the reverse of dissolution. The factors mentioned earlier — i.e., hydration energy and lattice energy — play a major role in determining whether an ionic compound will precipitate from solution. Ionic compounds (salts) in which the cation and anion have approximately the same size and the same charge tend to form especially stable crystal structures and precipitate from aqueous solution. Acidic cations — e.g., Al\(^{3+}\) — and nonbasic anions — e.g., NO\(_3^-\) — give rise to soluble salts because these ions are quite different in size and have much smaller lattice energy than hydration energy. Although these generalizations are useful, it is often difficult to make exact predictions, and it is worthwhile to remember a few simple aqueous solubility rules:

- Most alkali metal compounds are soluble.
- Most nitrates are soluble.
- All common inorganic acids are soluble.
- All ammonium salts are soluble.

You will certainly come across other solubility generalizations as you carry out the chemical reactions in this series of experiments.

Acid–Base Reactions

In aqueous solution an acid reacts with a base to give a salt and water. An acid–base reaction, sometimes called neutralization, is characterized by the formation of covalent, neutral water molecules from hydrated hydrogen ions (H\(_3\)O\(^+\)) and hydroxide ions (OH\(^-\)).

Complexation Reactions

Complexation reactions are closely related to acid–base reactions. Earlier, it was pointed out that some metal cations that have small size and high charge can act as acids — in fact they are called acidic cations. These acidic cations can react with electron–rich species called ligands to form complexes A complex may be defined as a chemical compound in which there is one or more coordinate–covalent bonds. A coordinate–covalent bond is a covalent bond in which the shared pair of electrons is provided by the ligand. A complexation reaction is therefore defined as a reaction in which one or more coordinate covalent bonds are produced during the formation of product.

Redox Reactions

Redox (an abbreviation of reduction–oxidation) reactions are defined as chemical reactions in which electrons are transferred from one species to another. Oxidation is defined as loss of electrons; reduction is defined as gain of electrons. A transfer of electrons means that the oxidation number of some of the atoms involved in the redox reaction must have changed. The oxidation number is defined by a simple set of rules based on arbitrarily counting electrons:
• The sum of the oxidation numbers of all the atoms in a molecule is zero.
• The sum in an ion equals the charge on the ion.
• Oxidation numbers are on a per-atom basis.
• Oxygen usually has an oxidation number of 2 (except in peroxides), and hydrogen has an oxidation number of 1.
• Monatomic ions (e.g., Na\(^+\)) have an oxidation number equal to the charge on the ion.

**Writing and Interpreting Chemical Equations**

Chemical reactions in aqueous solution are described very efficiently by writing a chemical equation in which the component ions of dissolved ionic compounds are written as separate ions — e.g., K\(^+\) and I\(^-\) rather than KI. These chemical equations are called ionic equations. Reactions that involve spectator ions can be written in a form in which the spectator ions are deleted (do not appear). These equations are called net ionic equations. The following simple rules should enable you to write most net ionic equations for a reaction:

• Soluble ionic salts are written as separate ions.
• Insoluble compounds are written as complete formula units with the subscript (s).
• Covalent compounds (e.g., CO\(_2\), H\(_2\)O) are written as molecules.

The example that follows develops the net ionic equation for a precipitation reaction between silver nitrate solution and hydrochloric acid solution.

Since the two solutes AgNO\(_3\) and HCl are obviously soluble and therefore dissolved, we can write the reactants as

\[ \text{Ag}^+ + \text{NO}_3^- + \text{H}^+ + \text{Cl}^- \]

understanding that all the ions are hydrated. In this reaction an off-white precipitate comes out of solution. The precipitate must be either AgCl\(_{(s)}\) or HNO\(_3\)(aq). In this instance it is easy to make the decision because we know that all common inorganic acids such as HNO\(_3\) are soluble in water. The precipitate must be AgCl\(_{(s)}\), and the spectator ions are H\(^+\) and NO\(_3^-\). The ionic equation is

\[ \text{Ag}^+ + \text{NO}_3^- + \text{H}^+ + \text{Cl}^- \rightleftharpoons \text{AgCl}_{(s)} + \text{H}^+ + \text{NO}_3^- \]

The net ionic equation is

\[ \text{Ag}^+ + \text{Cl}^- \rightleftharpoons \text{AgCl}_{(s)} \]

The long form of this equation would be

\[ \text{AgNO}_3 + \text{HCl} \rightleftharpoons \text{AgCl}_{(s)} + \text{HNO}_3 \]

Not only is the net ionic equation more concise, but it also suggests that any soluble Ag\(^+\) salt and any soluble Cl\(^-\) salt would give the same reaction. Hence, many chemical reactions are summarized in this single net ionic equation.
Laboratory Experiments

In this two-laboratory-period sequence of experiments, you can investigate a) the chemical factors that control solubility; b) the properties of solutions; c) chemical reactions in aqueous solutions; d) the correlation of chemical phenomenology, nomenclature, and reaction writing; and e) the relationships between chemical reactivity and the periodic chart. You will have ample opportunity to test your understanding of the chemical principles by analyzing the data from several experiments that involve identifying unknown solutions.

Section A  Naming and Making Solutions

Objectives  ► To write the chemical formula for fourteen solutes.

Discussion  ► Most of the equipment and the solutions which are required for this series of experiments is set out at your table.

Equipment List
Straws
Microstirrer
Wash bottle filled with distilled water
Microtowels
Waste disposal cup
Hand lens
Scissors
Tweezers
Plastic reaction surface

Other chemicals, etc., will be passed around the class or are available at Reagent Central.

Procedure

1. Have your hand lens, scissors, and tweezers with you.

2. Remove the reaction matrix chart labelled Figure 7, located after Section H, at the end of this section. In order to keep the sheet flat, cut off the serrated edges with your scissors. Insert the chart between the outer sheets of the plastic reaction surface.

3. Obtain a 24-well tray containing 14 empty and 2 filled microburets.

Each empty microburet is labelled with the formula for one of the 14 solutions that you will be using in this chapter’s experiments. The 2 filled microburets contain test solutions — universal acid-base indicator (labelled UI) in one, and starch/KI solution (labelled starch) in the other.

4. Use good transfer technique to fill each empty microburet with the appropriate solution.
Caution: Make sure you fill each microburet with the correct solution indicated by the label. A mistake in filling the microburets will cause all the chemical reactions to be incorrect.

5. Now, let’s see how many chemical names you know. Complete the table below, matching the formula on the label of the microburet to a name in the table.

Note: The test solutions are not listed in the table.

<table>
<thead>
<tr>
<th>Name</th>
<th>Chemical Formula of Solute</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sulfuric acid</td>
<td>1 M</td>
</tr>
<tr>
<td>Sodium phosphate</td>
<td>0.2 M</td>
</tr>
<tr>
<td>Sodium hydroxide</td>
<td>1 M</td>
</tr>
<tr>
<td>Sodium carbonate</td>
<td>0.2 M</td>
</tr>
<tr>
<td>Silver nitrate</td>
<td>0.1 M</td>
</tr>
<tr>
<td>Potassium iodide</td>
<td>0.2 M</td>
</tr>
<tr>
<td>Nitric acid</td>
<td>1 M</td>
</tr>
<tr>
<td>Lead nitrate</td>
<td>0.2 M</td>
</tr>
<tr>
<td>Hydrochloric acid</td>
<td>1 M</td>
</tr>
<tr>
<td>Iron (III) chloride</td>
<td>0.1 M</td>
</tr>
<tr>
<td>Copper (II) sulfate</td>
<td>0.2 M</td>
</tr>
<tr>
<td>Barium chloride</td>
<td>0.1 M</td>
</tr>
<tr>
<td>Ammonia</td>
<td>1 M</td>
</tr>
<tr>
<td>Aluminum chloride</td>
<td>0.3 M</td>
</tr>
</tbody>
</table>

6. How well did you do? Check with your instructor for the answers.

These solutions, which were made by the laboratory staff, are aqueous solutions. The correct amount of solute was dissolved in distilled water (the solvent), and the resulting solution was diluted with water to a known volume to produce a solution of exactly known concentration (which the table identifies). Before you start exploring chemical reactions, it is necessary to know some of the chemical properties of the individual solutions. In order to make a solution in the manner described earlier, you must know something about the solubility of the solute in water. Some general principles about the factors controlling solubility and about the nature of solutions were presented in the Background Chemistry section of this section.
Section B  Solubility and Solutions

Objectives  ► To examine some of the general principles governing the solubility of various solute in water and other solvents.

Discussion  ► Unless special techniques are used (they weren’t), solutions which are made in the presence of gases (e.g., the atmosphere) will contain dissolved gases.

- Which of the following two gases, ammonia or carbon dioxide, do you think would have the greatest solubility in distilled water? Explain your answer with a simple molecular picture.

\[
\begin{align*}
(a) & \quad \text{NH}_3 \\
(b) & \quad \text{CO}_2
\end{align*}
\]

ammonia
(nonpolar covalent molecule)

The solubility of liquids in water follows the same chemical principles. Here is some interesting data for the solubility of different alcohols in water.

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Solubility (moles L(^{-1}) at 25°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ethanol</td>
<td>C(_2)H(_5)OH</td>
<td>Infinite</td>
</tr>
<tr>
<td>1-pentanol</td>
<td>C(<em>5)H(</em>{11})OH</td>
<td>0.25</td>
</tr>
<tr>
<td>1-decanol</td>
<td>C(<em>{10})H(</em>{21})OH</td>
<td>0.00024</td>
</tr>
</tbody>
</table>

- Explain the data on the basis of a simple molecular picture.
- Speaking of the solubility of liquids, how can you remove ball point pen ink from the best Grateful Dead T-shirt?

Now, more to the point of the solutions in front of you. Of these solutions, 10 were made by dissolving ionically bonded solids in water and 4 by diluting pure liquid reagents (HCl, HNO\(_3\), H\(_2\)SO\(_4\), and NH\(_3\)) with water. The only solute that is not appreciably dissociated upon dissolution in water is ammonia.

- Give a simple explanation (with a picture) of the dissolution of solid potassium iodide in water to give a solution.
- Three of the solutions are going to be acidic. They are made by dissolving acids in water. Which are they?
Procedure

1. Use the plastic surface with the inserted reaction matrix chart, Figure 7, to carry out the following tests: Drop 1 drop of each of the acids (about 1 cm apart) onto a white background in the test area (indicated on the reaction matrix chart) of the plastic sheet.

   Note: Use the test area to mix solutions until indicated otherwise, in Section E.

2. To each drop add 1 drop of UI (universal indicator).
   - Record color changes. What happens with distilled water and UI?

3. Now test 1 drop of the ammonia solution with UI.
   - Note changes.

4. Try this one; put 1 drop UI next to a NH$_3$ drop.
   - What do you conclude?

   Note: Recall that ammonia is a base.

5. Try testing sodium hydroxide solution.
   To clean up the waste, add water from a wash bottle to dilute and mix all the drops. Suck the waste liquid up with a large-drop microburet and eject to waste cup. Wipe with a damp microtowel. This is a general procedure for disposing of all tests.

   Note: If the waste cup is full, transfer the solution to the appropriately labelled container at the side of the room.

The UI test solution gives you a practical way to distinguish between acids and bases. A simple way of writing the dissociation of an acid (e.g., HCl) in aqueous solution is

$$\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$$

Note: The test solution is responding to the H$^+$ ions.

Similarly, the dissociation of a base, e.g. NaOH, is written

$$\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$$

and the test solution responds to OH$^-$ ions.

The table that follows is a list of ions and molecules that you will encounter during these experiments.
<table>
<thead>
<tr>
<th>Cations</th>
<th>Anions</th>
<th>Molecules</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Formula</strong></td>
<td><strong>Name</strong></td>
<td><strong>Formula</strong></td>
</tr>
<tr>
<td>H⁺</td>
<td>hydrogen ion</td>
<td>Cl⁻</td>
</tr>
<tr>
<td>NH₄⁺</td>
<td>ammonium ion</td>
<td>I⁻</td>
</tr>
<tr>
<td>Na⁺</td>
<td>sodium ion</td>
<td>OH⁻</td>
</tr>
<tr>
<td>Ag⁺</td>
<td>silver ion</td>
<td>NO₃⁻</td>
</tr>
<tr>
<td>K⁺</td>
<td>potassium ion</td>
<td>HCO₃⁻</td>
</tr>
<tr>
<td>Ba²⁺</td>
<td>barium ion</td>
<td>CO₃²⁻</td>
</tr>
<tr>
<td>Pb²⁺</td>
<td>lead ion</td>
<td>SO₄²⁻</td>
</tr>
<tr>
<td>Cu²⁺</td>
<td>copper (II) ion</td>
<td>PO₄³⁻</td>
</tr>
<tr>
<td>Fe³⁺</td>
<td>iron (III) ion</td>
<td></td>
</tr>
<tr>
<td>Al³⁺</td>
<td>aluminum ion</td>
<td></td>
</tr>
</tbody>
</table>

Notice that one of the cations (NH₄⁺) and several anions (e.g., NO₃⁻) are polyatomic ions — i.e., 2 or more atoms are covalently bonded together in the ion. Water does not break the covalent bonds between the nonmetal atoms of a polyatomic ion, and these ions will generally stay intact in chemical reactions.

6. The remaining 9 untried solutions are solutions of salts as solutes. Test 1 drop of each solution with the UI test solution.

- Note any color changes.
- Write a simple dissociation and a reaction to explain your results for AlCl₃ and for Na₂CO₃ solutions.
- Give a simple picture of the actual state of the ions present in a solution of KI.

7. Clean up and dispose of the waste as before.

- Here is a strange question for you, but one you might be able to answer if you think about the general principles involved in this last series of experiments. Skywriting from airplanes is done by spraying a fine spray of TiCl₄ into the atmosphere. The radius of the Ti⁴⁺ ion is 74 pm. Explain the chemistry of skywriting.

**Section C  Solutions and Reactions**
Objectives

To observe the results of dissolution, diffusion, and reaction of KI and Pb(NO₃)₂.

Procedure

1. Plastic containers of solid, crystalline KI and Pb(NO₃)₂ will be circulated around the room. Use a clean, dry, straw spatula to transfer 2 or 3 small crystals of each compound to the plastic surface.

2. Push 1 crystal of KI and 2 or 3 crystals of Pb(NO₃)₂ together with a straw. Observe with your hand lens. Add 1 drop of water. Stir with a plastic stirrer.
   - What happens?

3. Make a pool of water about 1.5 cm in diameter on the plastic by expelling water from a microburet.

4. Carefully push 1 crystal KI near one side of the pool, but not in it. Push a few crystals of Pb(NO₃)₂ near the other side. The position of the crystals is illustrated below.

5. Gently push the Pb(NO₃)₂ crystals onto the edge of the pool. Wait a few seconds. Now do the same with the KI. Use your hand lens to examine the reaction.
   - Explain the chemistry of what you see.
   - Predict exactly what will happen if 1 drop of KI solution is added to 1 drop of Pb(NO₃)₂ solution.
   - What would happen if they were added the other way around?
   - What is the most important reason for using solutions to carry out chemical reactions of this type?

6. Clean up as usual, remembering that lead compounds are toxic!
Section D  Four Major Types of Chemical Reaction in Aqueous Solution

Objectives  
To investigate the four major types of chemical reactions which occur in aqueous solutions.

Discussion  
In the last section you carried out a chemical reaction between KI and Pb(NO₃)₂ in an aqueous medium. The reaction between these two reactants produced a solid product. In general, products that are not soluble in water and precipitate from solution are called precipitates, and the reactions are called precipitation reactions. Here we are, back at the principle of solubility again! In order to write the net ionic reaction, you first need to identify the precipitate. There are two possible choices.
- What are they?
- Which one of these is not soluble in water?
- Perhaps it’s easier to ask which one is soluble in water.
- Now write the net ionic reaction.

Procedure

1. Use the UI test solution to investigate the reaction between HCl and NaOH solutions.
   This type of reaction is called an acid–base reaction.
   - Write the net ionic reaction, which by the way is the same for all acid–base reactions in aqueous solution.
   - Why is a test solution necessary in order to follow this reaction?

2. Now explore the reaction between CuSO₄ and NH₃ solutions by dropping 3 drops of NH₃ onto the plastic and adding 1 drop of CuSO₄. Stir.
   A change has obviously occurred. The net ionic reaction is

\[
\text{Cu}^{2+} + 4\text{NH}_3 \rightarrow \text{Cu(NH}_3)_4^{2+}
\]

This is an example of a complexation reaction. The deep–blue soluble product Cu(NH₃)₄²⁺ (illustration follows) is a complex called tetraamminecopper(II) ion. Complexes are formed between cations and molecules or ions that have lone pairs of electrons (e.g., NH₃ in the above reaction).
\[
\text{Cu}^{2+} + 4 \text{NH}_3 \rightleftharpoons \text{complexion}
\]

The bond between each NH\textsubscript{3} and Cu\textsuperscript{2+} arises from the lone pair on the N of NH\textsubscript{3} being shared with Cu\textsuperscript{2+}. Sometimes these bonds are called coordinate-covalent bonds.

3. Now add 2 more drops of CuSO\textsubscript{4} to the complex ion. Stir.

The net reaction is

\[
\text{Cu}^{2+} + 2\text{OH}^- \rightarrow \text{Cu(OH)}_{2(4)}
\]

- Now why did that happen?

**Hint** We swamped it with CuSO\textsubscript{4}!

4. One more type of reaction. React 1 drop CuSO\textsubscript{4} and 1 drop KI. Stir. Ugly.

5. Now drop 1 drop of starch test solution very close to (but not in) the reaction mixture. Wait and watch.

Starch/KI is an excellent probe for molecular iodine I\textsubscript{2}. The reaction must have produced I\textsubscript{2} as a product. Since the KI solution contains I\textsuperscript{-}, then during the reaction

\[
2\text{I}^- \rightarrow \text{I}_2 + 2 \text{electrons}
\]

electrons are produced as I\textsuperscript{+} goes to I\textsubscript{2}. Something must react with the electrons and it is Cu\textsuperscript{2+}:

\[
2 \text{Cu}^{2+} + 2 \text{electrons} \rightarrow 2\text{Cu}^+
\]

If we add these two reactions, (i) and (ii), we will get the net ionic reaction

\[
2 \text{Cu}^{2+} + 2\text{I}^- \rightarrow 2 \text{Cu}^+ + \text{I}_2
\]

This type of reaction is called a **reduction-oxidation** reaction (redox) because it is a combination of reduction [gain of electrons by Cu\textsuperscript{2+}, reaction (ii)] and oxidation [loss of electrons by I\textsuperscript{-}, reaction (i)]. Redox reactions are reactions in which electrons are transferred from one chemical species to another. In this example, from I\textsuperscript{-} to Cu\textsuperscript{2+}.

6. Clean up as usual.
Section E  A Chemical Reaction Survey: The Ion Reaction Chart

Objectives ▶ (1) To binary-mix the solutions. (2) To describe and record chemical changes. (3) To write net ionic reactions and identify the type. (4) To name the products of chemical reactions. (5) To correlate reactivity with trends in the periodic table.

Note: Before beginning with Step 1, read Section E completely.

Procedure

1. Binary-mix the 14 solutions provided in Section A (mix one solution with one other) in a dropwise manner on the rectangles (not the test area) of the plastic reaction surface (with the inserted reaction matrix chart).

Note: Each rectangle corresponds to one mixing. The background is white and black so that you can easily see color changes and/or precipitates.

- Figure 8, located at the end of the chapter, is a reaction description chart (actually a reaction matrix chart without heavy lines), which you should use to record what you see (the phenomenology) in the binary mixing.
- Devise an accurate code that can be used to describe color changes, precipitate formation or disappearance, textures of precipitates, gas evolution, etc.
- Write the net ionic reactions in your laboratory record as you go. You may need to review the background chemistry section to do this. Identify the type (e.g., precipitation) and give the name of the product.

Note: The best way to mix is to drop 1 drop of a solution onto the rectangle. Then add 1 drop of the second solution, stir, add a second, stir.

If you are in doubt about a reaction, feel free to play in the test area. Make sure you understand the arrangement of the reaction chart before you start!

Use cotton swabs and microtowels to keep things neat. Don’t write on the plastic sheet — it will destroy your flat reaction surface and drops will roll all over the place.

2. When you have finished, leave the mixed solutions on the reaction chart.
Section F  Identification of Three Unknowns

Objectives

To identify three unknowns by comparison with the reactions of the 14 known solutions.

Procedure

1. Obtain your 3 unknowns and place them into the 24 well tray.

   Note: One unknown is one of the 14 solutions. Do this one first.

   - Write the unknown number in your lab record and on the line provided in Figure 8, the reaction description chart, in the first rectangle of the third row from the bottom.

2. Binary–mix the first unknown with the other 14 solutions in the same manner as you mixed the known solutions.

   - Compare the results of the reactions of the 14 solutions listed on the reaction matrix chart with the reaction of your unknown. Remember that some time has elapsed and some changes may have occurred — e.g., bubbles may have flown away.
   - Identify the unknown solution.

3. Binary–mix the second unknown solution with the 14 known solutions.

   Note: The second unknown contains a cation that is one of the known solution cations and an anion that it is not associated with in one of the known solutions.

   - Identify the cation in the second unknown.

4. Binary–mix the third unknown solution with the 14 known solutions.

   Note: The third unknown contains an anion that is one of the known solution anions and a cation that it is not associated with in the known solutions.

   - Identify the anion in the third unknown.
   - Record the numbers and chemical identities of your unknowns in your laboratory record.

5. Clean the entire reaction surface.
Section G  Five Unknowns: Solo

Objectives  ►To identify five unknowns by simply binary-mixing them with each other.

Procedure

1. Obtain a set of 5 unknowns (your instructor will assign them).
   The 5 unknowns are selected from the known solutions.
2. Binary-mix the 5 unknowns only with each other in the chart space labelled 5 unknowns.
3. Identify the unknown solutions.
   • Record the result in your laboratory record.

Section H  Finding General Solubility Rules

Objectives  ►To look at all of your data in order to determine a minimum set of simple solubility rules that would be operative for all the combinations of ions that you have investigated.

Note: This is not as difficult as it looks—e.g., you might begin by counting how many insoluble nitrates you have seen, and so on.
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Pre-Laboratory Quiz

What is the chemical name for bone?

Give a definition of a solution.

In an aqueous solution the solubility of the solute is dominated by 3 interactions. What are they?

What makes water have such strong hydrogen bonding?

Give a chemical equation that describes the reaction of an aluminum ion with water.

Give an example of a nonacidic cation.
What is the formula for a carbonate ion?

Write the net ionic equation for the reaction of dissolved AgNO₃ with dissolved NaCl in aqueous solution.

What is the definition of oxidation?