Reactions in Aqueous Solutions

9.2

When you mix two aqueous ionic compounds together, there are two possible outcomes. Either the compounds will remain in solution without reacting, or one aqueous ionic compound will chemically react with the other. How can you predict which outcome will occur? Figure 9.4 shows what happens when an aqueous solution of lead(II) nitrate is added to an aqueous solution of potassium iodide. As you can see, a yellow solid—a precipitate—is forming. This is a double displacement reaction. Recall, from Chapter 4, that a double displacement reaction is a chemical reaction that involves the exchange of ions to form two new compounds. It has the general equation

$$WX + YZ \rightarrow WZ + YX$$

In a double displacement reaction, the cations exchange anions. In the reaction shown in Figure 9.4, for example, the lead cation is exchanged with the iodide anion.

You can usually recognize a double displacement reaction by observing one of these possible results:

- the formation of a precipitate (so that ions are removed from solution as an insoluble solid)
- the formation of a gas (so that ions are removed from solution in the form of a gaseous product)
- the formation of water (so that H^+ and OH^- ions are removed from solution as water)

In this section, you will examine each of these results. At the same time, you will learn how to represent a double displacement reaction using a special kind of chemical equation: an ionic equation.

Figure 9.4 **Lead(II) nitrate and potassium iodide are clear, colourless aqueous solutions. Mixing them causes a double displacement reaction. An insoluble yellow precipitate (lead(II) iodide) and a soluble salt (potassium nitrate) are produced.**

Section Preview/ Specific Expectations

In this section, you will

- **describe** combinations of aqueous solutions that result in the formation of precipitates
- **perform** a qualitative analysis of ions in solutions
- **represent** double displacement reactions by their net ionic equations
- **write** balanced chemical equations and net ionic equations for double displacement reactions
- **communicate** your understanding of the following terms: *spectator ions, total ionic equation, net ionic equation, qualitative analysis*

Language LINK

Double displacement reactions are also called *metathesis reactions*. The word "metathesis" (pronounced with the stress on the second syllable: meh-TATH-e-sis) means "interchange." In chemistry, a metathesis reaction occurs when ions or atoms are exchanged between different compounds. Chemists are not the only people who use this word. Use a dictionary or encyclopedia to find non-chemistry examples of metathesis.

Double Displacement Reactions That Produce a Precipitate

A double displacement reaction that results in the formation of an insoluble substance is often called a *precipitation reaction*. Figure 9.4 is a clear example of a precipitation reaction. What if you did not have this photograph, however, and you were unable to do an experiment? Could you have predicted that mixing $Pb(NO_3)_{(aq)}$ and $2KI_{(aq)}$ would result in an insoluble compound? Yes. When you are given (on paper) a pair of solutions to be mixed together, start by thinking about the exchange of ions that may occur. Then use the general solubility guidelines (Table 9.1) to predict which compounds, if any, are insoluble.

For example, consider lead(II) nitrate, $Pb(NO_3)_2$, and potassium iodide, KI. Lead(II) nitrate contains Pb^{2+} cations and NO_3^- anions. Potassium iodide contains K^+ cations and I^- anions. Exchanging positive ions results in lead(II) iodide, PbI_2, and potassium nitrate, KNO_3. From the solubility guidelines, you know that all potassium salts and nitrates are soluble. Thus, potassium nitrate is soluble. The Pb^{2+} ion is listed in guideline 2 as an insoluble cation. The I^- ion is listed in guideline 3 as a soluble anion. Remember that a higher guideline number takes precedence over a lower guideline number. Thus, you can predict that lead(II) iodide is insoluble. It will form a precipitate when the solutions are mixed. The balanced chemical equation for this reaction is

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

Sample Problem

Predicting the Formation of a Precipitate

Problem

Which of the following pairs of aqueous solutions produce a precipitate when mixed together? Write the balanced chemical equation if you predict a precipitate. Write "NR" if you predict that no reaction takes place.

(a) potassium carbonate and copper(II) sulfate

(b) ammonium chloride and zinc sulfate

What Is Required?

You need to predict whether or not each pair of aqueous solutions forms an insoluble product (a precipitate). If it does, you need to write a balanced chemical equation.

What Is Given?

You know the names of the compounds in each solution.

Plan Your Strategy

Start by identifying the ions in each pair of compounds. Then exchange the positive ions in the two compounds. Compare the resulting compounds against the solubility guidelines, and make your prediction.

Continued ...

Continued ...
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Act on Your Strategy

(a) Potassium carbonate contains K^+ and CO_3^{2-} ions. Copper(II) sulfate contains Cu^{2+} and SO_4^{2-} ions. Exchanging positive ions results in potassium sulfate, K_2SO_4, and copper(II) carbonate, $CuCO_3$.

All potassium salts are soluble, so these ions remain dissolved in solution.

The copper(II) ion is listed in guideline 5 as a soluble cation. The carbonate anion is listed in guideline 2 as insoluble. Because guideline 2 is higher, copper(II) carbonate should be insoluble. So you can predict that a precipitate forms. The balanced chemical equation for this reaction is

$$K_2CO_{3(aq)} + CuSO_{4(aq)} \rightarrow K_2SO_{4(aq)} + CuCO_{3(s)}$$

(b) Ammonium chloride contains NH_4^+ and Cl^- ions. Zinc sulfate consists of Zn^{2+} and SO_4^{2-} ions. Exchanging positive ions results in ammonium sulfate, $(NH_4)_2SO_4$, and zinc chloride, $ZnCl_2$.

Since all ammonium salts are soluble, the ammonium sulfate stays dissolved in solution.

Zinc chloride consists of a guideline 5 soluble cation and a guideline 3 soluble anion. Because guideline 3 is higher, zinc chloride should be soluble. Thus, you can predict that no precipitate forms.

$$NH_4Cl_{(aq)} + ZnSO_{4(aq)} \rightarrow NR$$

Check Your Solution

An experiment is always the best way to check a prediction. If possible, obtain samples of these solutions from your teacher, and mix them together.

Practice Problems

4. Predict the result of mixing each pair of aqueous solutions. Write a balanced chemical equation if you predict that a precipitate forms. Write “NR” if you predict that no reaction takes place.
   - **(a)** sodium sulfide and iron(II) sulfate
   - **(b)** sodium hydroxide and barium nitrate
   - **(c)** cesium phosphate and calcium bromide
   - **(d)** sodium carbonate and sulfuric acid
   - **(e)** sodium nitrate and copper(II) sulfate
   - **(f)** ammonium iodide and silver nitrate
   - **(g)** potassium carbonate and iron(II) nitrate
   - **(h)** aluminum nitrate and sodium phosphate
   - **(i)** potassium chloride and iron(II) nitrate
   - **(j)** ammonium sulfate and barium chloride
   - **(k)** sodium sulfide and nickel(II) sulfate
   - **(l)** lead(II) nitrate and potassium bromide

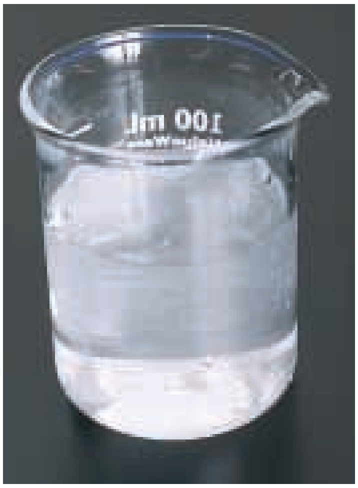

Figure 9.5 You can easily identify limestone and marble by their reaction with hydrochloric acid. The gas that is produced by this double displacement reaction is carbon dioxide.

Double Displacement Reactions That Produce a Gas

Double displacement reactions are responsible for producing a number of gases. (See Figure 9.5.) These gases include

- hydrogen
- hydrogen sulfide (a poisonous gas that smells like rotten eggs)
- sulfur dioxide (a reactant in forming acid rain)
- carbon dioxide
- ammonia

A Reaction that Produces Hydrogen Gas

The alkali metals form bonds with hydrogen to produce compounds called hydrides. Hydrides react readily with water to produce hydrogen gas. Examine the following equation for the reaction of lithium hydride, LiH, with water. If you have difficulty visualizing the ion exchange that takes place, rewrite the equation for yourself using HOH instead of H_2O.

$$LiH_{(s)} + H_2O_{(\ell)} \rightarrow LiOH_{(aq)} + H_{2(g)}$$

A Reaction that Produces Hydrogen Sulfide Gas

Sulfides react with certain acids, such as hydrochloric acid, to produce hydrogen sulfide gas.

$$K_2S_{(aq)} + 2HCl_{(aq)} \rightarrow 2KCl_{(aq)} + H_2S_{(g)}$$

A Reaction that Produces Sulfur Dioxide Gas

Some reactions produce a compound that, afterward, decomposes into a gas and water. Sodium sulfite is used in photography as a preservative. It reacts with hydrochloric acid to form sulfurous acid. The sulfurous acid then breaks down into sulfur dioxide gas and water. The net reaction is the sum of both changes. If the same compound appears on both sides of an equation (as sulfurous acid, H_2SO_3, does here), it can be eliminated. This is just like eliminating terms from an equation in mathematics.

$$Na_2SO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + \cancel{H_2SO_{3(aq)}}$$
$$\cancel{H_2SO_{3(aq)}} \rightarrow SO_{2(g)} + H_2O_{(\ell)}$$

Therefore, the net reaction is

$$Na_2SO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + SO_{2(g)} + H_2O_{(\ell)}$$

A Reaction that Produces Carbon Dioxide Gas

The reaction of a carbonate with an acid produces carbonic acid. Carbonic acid decomposes rapidly into carbon dioxide and water.

$$Na_2CO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + \cancel{H_2CO_{3(aq)}}$$
$$\cancel{H_2CO_{3(aq)}} \rightarrow CO_{2(g)} + H_2O_{(\ell)}$$

The net reaction is

$$Na_2CO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(\ell)}$$

A Reaction that Produces Ammonia Gas

Ammonia gas is very soluble in water. You can detect it easily, however, by its sharp, pungent smell. Ammonia gas can be prepared by the reaction of an ammonium salt with a base.

$$NH_4Cl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + NH_{3(aq)} + H_2O_{(\ell)}$$

Double Displacement Reactions That Produce Water

The neutralization reaction between an acid and a base is a very important double displacement reaction. In a neutralization reaction, water results when an H^+ ion from the acid bonds with an OH^- ion from the base.

$$H_2SO_{4(aq)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(\ell)}$$

Most metal oxides are bases. Therefore, a metal oxide will react with an acid in a neutralization reaction to form a salt and water.

$$2HNO_{3(aq)} + MgO_{(s)} \rightarrow Mg(NO_3)_{2(aq)} + H_2O_{(\ell)}$$

Non-metal oxides are acidic. Therefore, a non-metal oxide will react with a base. This type of reaction is used in the space shuttle. Cabin air is circulated through canisters of lithium hydroxide (a base) to remove the carbon dioxide before it can reach dangerous levels.

$$2LiOH_{(s)} + CO_{2(g)} \rightarrow Li_2CO_{3(aq)} + H_2O_{(\ell)}$$

Representing Aqueous Ionic Reactions with Net Ionic Equations

Mixing a solution that contains silver ions with a solution that contains chloride ions produces a white precipitate of silver chloride. There must have been other ions present in each solution, as well. You know this because it is impossible to have a solution of just a cation or just an anion. Perhaps the solution that contained silver ions was prepared using silver nitrate or silver acetate. Similarly, the solution that contained chloride ions might have been prepared by dissolving NaCl in water, or perhaps NH_4Cl or another soluble chloride. Any solution that contains $Ag^+_{(aq)}$ will react with any other solution that contains $Cl^-_{(aq)}$ to form a precipitate of $AgCl_{(s)}$. The other ions in the solutions are not important to the net result. These ions are like passive onlookers. They are called **spectator ions**.

The reaction between silver nitrate and sodium chloride can be represented by the following chemical equation:

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow NaNO_{3(aq)} + AgCl_{(s)}$$

This equation does not show the change that occurs, however. It shows the reactants and products as intact compounds. In reality, soluble ionic compounds dissociate into their respective ions in solution. So chemists often use a **total ionic equation** to show the dissociated ions of the soluble ionic compounds.

$$Ag^+_{(aq)} + NO_3{}^-_{(aq)} + Na^+_{(aq)} + Cl^-_{(aq)} \rightarrow Na^+_{(aq)} + NO_3{}^-_{(aq)} + AgCl_{(s)}$$

Notice that the precipitate, AgCl, is still written as an ionic formula. This makes sense because precipitates are insoluble, so they do not dissociate into ions. Also notice that the spectator ions appear on both sides of the equation. Here is the total ionic equation again, with slashes through the spectator ions.

$$Ag^+_{(aq)} + \cancel{NO_3{}^-_{(aq)}} + \cancel{Na^+_{(aq)}} + Cl^-_{(aq)} \rightarrow \cancel{Na^+_{(aq)}} + \cancel{NO_3{}^-_{(aq)}} + AgCl_{(s)}$$

If you eliminate the spectator ions, the equation becomes

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

An ionic equation that is written this way, without the spectator ions, is called a **net ionic equation**. Before you try writing your own net ionic equations, examine the guidelines in Table 9.2 below.

Table 9.2 Guidelines for Writing a Net Ionic Equation

1. Include only ions and compounds that have reacted. Do not include spectator ions.
2. Write the soluble ionic compounds as ions. For example, write $NH_4{}^+{}_{(aq)}$ and $Cl^-{}_{(aq)}$, instead of $NH_4Cl_{(aq)}$.
3. Write insoluble ionic compounds as formulas, not ions. For example, zinc sulfide is insoluble, so you write it as $ZnS_{(s)}$, not Zn^{2+} and S^{2-}.
4. Since covalent compounds do not produce ions in aqueous solution, write their molecular formulas. Water is a common example, because it dissociates only very slightly into ions. When a reaction involves a gas, always include the gas in the net ionic equation.
5. Write strong acids (discussed in the next chapter) in their ionic form. There are six strong acids:
   - hydrochloric acid (write as $H^+{}_{(aq)}$ and $Cl^-{}_{(aq)}$, not $HCl_{(aq)}$)
   - hydrobromic acid (write as $H^+{}_{(aq)}$ and $Br^-{}_{(aq)}$)
   - hydroiodic acid (write as $H^+{}_{(aq)}$ and $I^-{}_{(aq)}$)
   - sulfuric acid (write as $H^+{}_{(aq)}$ and $SO_{4(aq)}$)
   - nitric acid (write as $H^+{}_{(aq)}$ and $NO_3{}^-{}_{(aq)}$)
   - perchloric acid (write as $H^+{}_{(aq)}$ and $ClO_4{}^-{}_{(aq)}$)

   All other acids are weak and form few ions. Therefore, write them in their molecular form.
6. Finally, check that the net ionic equation is balanced for charges as well as for atoms.

Sample Problem

Writing Net Ionic Equations

Problem

A chemical reaction occurs when the following aqueous solutions are mixed: sodium sulfide and iron(II) sulfate. Identify the spectator ions. Then write the balanced net ionic equation.

What Is Required?

You need to identify the spectator ions and write a balanced net ionic equation for the reaction between sodium sulfide and iron(II) sulfate.

What Is Given?

You know the chemical names of the compounds.

Continued ...

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Plan Your Strategy

Step 1 Start by writing the chemical formulas of the given compounds.

Step 2 Then write the complete chemical equation for the reaction, using your experience in predicting the formation of a precipitate.

Step 3 Once you have the chemical equation, you can replace the chemical formulas of the soluble ionic compounds with their dissociated ions.

Step 4 This will give you the total ionic equation. Next you can identify the spectator ions (the ions that appear on both sides of the equation).

Step 5 Finally, by rewriting the total ionic equation *without* the spectator ions, you will have the net ionic equation.

Act on Your Strategy

Steps 1 and 2 The chemical equation for the reaction is

$Na_2S_{(aq)} + FeSO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + FeS_{(s)}$

Step 3 The total ionic equation is

$\cancel{2Na^+_{(aq)}} + S^{2-}_{(aq)} + Fe^{2+}_{(aq)} + \cancel{SO_4{}^{2-}_{(aq)}} \rightarrow \cancel{2Na^+_{(aq)}} + \cancel{SO_4{}^{2-}_{(aq)}} + FeS_{(s)}$

Step 4 Therefore, the spectator ions are $Na^+_{(aq)}$ and $SO_4{}^{2-}_{(aq)}$.

Step 5 The net ionic equation is

$Fe^{2+}_{(aq)} + S^{2-}_{(aq)} \rightarrow FeS_{(s)}$

Check Your Solution

Take a final look at your net ionic equation to make sure that no ions are on both sides of the equation.

Practice Problems

5. Mixing each pair of aqueous solutions results in a chemical reaction. Identify the spectator ions. Then write the balanced net ionic equation.

(a) sodium carbonate and hydrochloric acid

(b) sulfuric acid and sodium hydroxide

6. Identify the spectator ions for the reaction that takes place when each pair of aqueous solutions is mixed. Then write the balanced net ionic equation.

(a) ammonium phosphate and zinc sulfate

(b) lithium carbonate and nitric acid

(c) sulfuric acid and barium hydroxide

Identifying Ions in Aqueous Solution

Suppose that you have a sample of water. You want to know what, if any, ions are dissolved in it. Today technological devices, such as the mass spectrometer, make this investigative work fairly simple. Before such devices, however, chemists relied on *wet chemical techniques*: experimental tests, such as submitting a sample to a series of double displacement reactions. Chemists still use wet chemical techniques. With each reaction, insoluble compounds precipitate out of the solution. (See Figure 9.6.) This enables the chemist to determine, eventually, the identity of one or several ions in the solution. This ion-identification process is an example of **qualitative analysis**.

Chemists use a range of techniques for qualitative analysis. For example, the colour of an aqueous solution can help to identify one of the ions that it contains. Examine Table 9.3. However, the intensity of ion colour varies with its concentration in the solution. Also keep in mind that many ions are colourless in aqueous solution. For example, the cations of elements from Groups 1 (IA) and 2 (IIA), as well as aluminum, zinc, and most anions, are colourless. So there are limits to the inferences you can make if you rely on solution colour alone.

Another qualitative analysis technique is a flame test. A dissolved ionic compound is placed in a flame. Table 9.4 lists the flame colours associated with several ions. Notice that all the ions are metallic. The flame test is only useful for identifying metallic ions in aqueous solution.

Qualitative analysis challenges a chemist's creative imagination and chemical understanding. Discover this for yourself in Investigation 9-B.

Table 9.3 The Colour of Some Common Ions in Aqueous Solution

	Ions	Symbol	Colour
Cations	chromium (II) copper(II)	Cr^{2+} Cu^{2+}	blue
Cations	chromium(III) copper(I) iron(II) nickel(II)	Cr^{3+} Cu^{+} Fe^{2+} Ni^{2+}	green
Cations	iron(III)	Fe^{3+}	pale yellow
Cations	cobalt(II) manganese(II)	Co^{2+} Mn^{2+}	pink
Anions	chromate	CrO_4^{2-}	yellow
Anions	dichromate	$Cr_2O_7^{2-}$	orange
Anions	permanganate	MnO_4^{-}	purple

Table 9.4 The Flame Colour of Selected Metallic Ions

Ion	Symbol	Colour
lithium	Li^{+}	red
sodium	Na^{+}	yellow
potassium	K^{+}	violet
cesium	Cs^{+}	violet
calcium	Ca^{2+}	red
strontium	Sr^{2+}	red
barium	Ba^{2+}	yellowish-green
copper	Cu^{2+}	bluish-green
boron	B^{2+}	green
lead	Pb^{2+}	bluish-white

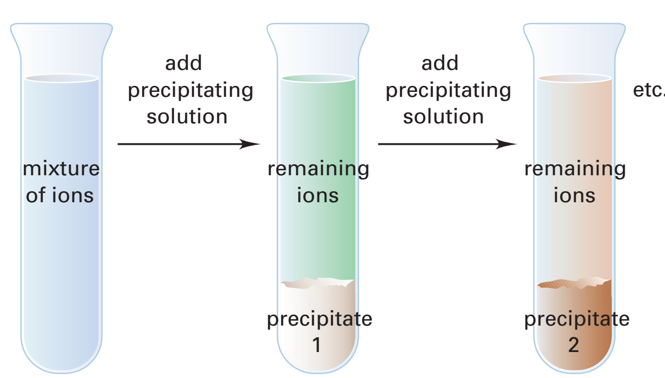

Figure 9.6 This illustration shows the basic idea behind a qualitative analysis for identifying ions in an aqueous solution. At each stage, the resulting precipitate is removed.

Investigation 9-B

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

Qualitative Analysis

In this investigation, you will apply your knowledge of chemical reactions and the general solubility guidelines to identify unknown ions.

Question

How can you identify ions in solution?

Predictions

Read the entire Procedure. Can you predict the results of any steps? Write your predictions in your notebook. Justify each prediction.

Materials

Part 1

12-well or 24-well plate, or spot plate
toothpicks
cotton swabs
unknowns: 4 dropper bottles (labelled A, B, C, and D) of solutions that include $Na^{+}{}_{(aq)}$, $Ag^{+}{}_{(aq)}$, $Ca^{2+}{}_{(aq)}$, and $Cu^{2+}{}_{(aq)}$
reactants: 2 labelled dropper bottles, containing dilute $HCl_{(aq)}$ and dilute $H_2SO_{4(aq)}$

Part 2

cotton swabs
Bunsen burner
heat-resistant pad
unknowns: 4 dropper bottles, containing the same unknowns that were used in Part 1
reactants: 4 labelled dropper bottles, containing $Na^{+}{}_{(aq)}$, $Ag^{+}{}_{(aq)}$, $Ca^{2+}{}_{(aq)}$, and $Cu^{2+}{}_{(aq)}$

Part 3

12-well or 24-well plate, or spot plate
toothpicks
cotton swabs
unknowns: 3 dropper bottles (labelled X, Y, and Z), containing solutions of $SO_4{}^{2-}{}_{(aq)}$, $CO_3{}^{2-}{}_{(aq)}$, and $I^{-}{}_{(aq)}$
reactants: 3 labelled dropper bottles, containing $Ba^{2+}{}_{(aq)}$, $Ag^{+}{}_{(aq)}$, and $HCl_{(aq)}$

Safety Precautions

- Be careful not to contaminate the dropper bottles. The tip of a dropper should not make contact with either the plate or another solution. Put the cap back on the bottle immediately after use.
- Hydrochloric acid and sulfuric acid are corrosive. Wash any spills on your skin with plenty of cool water. Inform your teacher immediately.
- Part 2 of this investigation requires an open flame. Tie back long hair, and confine any loose clothing.

Procedure

Part 1 Using Acids to Identify Cations

1. Read steps 2 and 3 below. Design a suitable table for recording your observations.
2. Place one or two drops of each unknown solution into four different wells or spots. Add one or two drops of hydrochloric acid to each unknown. Record your observations.
3. Repeat step 2. This time, test each unknown solution with one or two drops of sulfuric acid. Record your observations.
4. Answer Analysis questions 1 to 5.

Part 2 Using Flame Tests to Identify Cations

Note: Your teacher may demonstrate this part or provide you with an alternative version.

1. Design tables to record your observations.
2. Observe the appearance of each known solution. Record your observations. Repeat for each unknown solution. Some cations have a characteristic colour. (Refer to Table 9.4.) If you think that you can identify one of the unknowns, record your identification.

3. Flame tests can identify some cations. Set up the Bunsen burner and heat-resistant pad. Light the burner. Adjust the air supply to produce a hot flame with a blue cone.
4. Place a few drops of solution containing $Na^{+}_{(aq)}$ on one end of a cotton swab.

   CAUTION Carefully hold the saturated tip so it is just in the Bunsen burner flame, near the blue cone. You may need to hold it in this position for as long as 30 s to allow the solution to vaporize and mix with the flame. Record the colour of the flame.
5. Not all cations give colour to a flame. The sodium ion *does* give a distinctive colour to a flame, however. It is often present in solutions as a contaminant. For a control, repeat step 4 with water and record your observations. You can use the other end of the swab for a second test. Dispose of used swabs in the container your teacher provides.
6. Repeat the flame test for each of the other known solutions. Then test each of the unknown solutions.
7. Answer Analysis question 6.

Part 3 Identifying Anions

1. Place one or two drops of each unknown solution into three different wells or spots. Add one or two drops of $Ba^{2+}_{(aq)}$ to each unknown solution. Stir with a toothpick. Record your observations.
2. Add a drop of hydrochloric acid to any well or spot where you observed a precipitate in step 1. Stir and record your observations.
3. Repeat step 1, adding one or two drops of $Ag^{+}_{(aq)}$ to each unknown solution. Record the colour of any precipitate that forms.
4. Answer Analysis questions 7 to 9.

Analysis

1. (a) Which of the cations you tested should form a precipitate with hydrochloric acid? Write the net ionic equation.
   (b) Did your results support your predictions? Explain.
2. (a) Which cation(s) should form a precipitate when tested with sulfuric acid? Write the net ionic equation.
   (b) Did your results support your predictions? Explain.
3. Which cation(s) should form a soluble chloride and a soluble sulfate?
4. Which cation has a solution that is not colourless?
5. Based on your analysis so far, tentatively identify each unknown solution.
6. Use your observations of the flame tests to confirm or refute the identifications you made in question 5. If you are not sure, check your observations and analysis with other students. If necessary, repeat some of your tests.
7. Which anion(s) should form a precipitate with Ba^{2+}? Write the net ionic equation.
8. Which precipitate should react when hydrochloric acid is added? Give reasons for your prediction.
9. Tentatively identify each anion. Check your observations against the results you obtained when you added hydrochloric acid. Were they what you expected? If not, check your observations and analysis with other students. If necessary, repeat some of your tests.

Conclusion

10. Identify the unknown cations and anions in this investigation. Explain why you do, or do not, have confidence in your decisions. What could you do to be more confident?

Section Wrap-up

Qualitative analysis helps you identify ions that may be present in a solution. It does not, however, tell you how much of these ions are present. In other words, it does not provide any quantitative information about the quantity or concentration of ions in solution. In the next section, you will find out how to calculate this quantitative information, using techniques you learned in Unit 2.

Section Review

1 C Briefly compare the relationships among a chemical formula, a total ionic equation, and a net ionic equation. Use sentences or a graphic organizer.

2 K/U Write a net ionic equation for each double displacement reaction in aqueous solution.

(a) tin(II) chloride with potassium phosphate

(b) nickel(II) chloride with sodium carbonate

(c) chromium(III) sulfate with ammonium sulfide

3 K/U For each reaction in question 2, identify the spectator ions.

4 K/U Would you expect a qualitative analysis of a solution to give you the amount of each ion present? Explain why or why not.

5 I A solution of limewater, $Ca(OH)_{2(aq)}$, is basic. It is used to test for the presence of carbon dioxide. Carbon dioxide is weakly acidic and turns limewater milky. Use a chemical equation to explain what happens during the test. What type of reaction occurs?

6 K/U State the name and formula of the precipitate that forms when aqueous solutions of copper(II) sulfate and sodium carbonate are mixed. Write the net ionic equation for the reaction. Identify the spectator ions.

7 I All the solutions in this photograph have the same concentration: 0.1 mol/L. Use Table 9.3 to infer which ion causes the colour in each solution. How much confidence do you have in your inferences? What could you do to increase your confidence?