Topic 1: Reactions in Aqueous Solutions

Appendices

Appendix 1.1A: Developing a Set of Solubility Rules: Lab Activity

Purpose

In this lab activity, you will develop your own procedure to create a set of solubility rules. You will be provided with 0.1 mol/L solutions of various anions and cations so that you can observe whether precipitates are formed.

Solutions

The solutions the class will use include the following:

- **Set A**
  silver ions (Ag^{+})
  barium ions (Ba^{2+})
  sodium ions (Na^{+})
  ammonium ions (NH_4^{+})
  calcium ions (Ca^{2+})
  chloride ions (Cl^{-})
  carbonate ions (CO_3^{2-})
  sulphate ions (SO_4^{2-})
  nitrate ions (NO_3^{-})
  phosphate ions (PO_4^{3-})
- **Set B**
  zinc ions (Zn^{2+})
  iron ions (Fe^{3+})
  sodium ions (Na^{+})
  magnesium ions (Mg^{2+})
  potassium ions (K^{+})
  chloride ions (Cl^{-})
  hydroxide ions (OH^{-})
  bromide ions (Br^{-})
  carbonate ions (CO_3^{2-})
  acetate ions ($C_2H_3O_2^{-}$)

Before you begin mixing solutions, set up a grid to organize your observations.

Follow-up Questions

1. Scientists have developed a set of solubility rules with respect to the solubility of anions with numerous cations.
   a) List the cations that did not form any precipitates.
   b) For each anion, list the cations with which it was insoluble (formed a precipitate).
2. List the set of solubility rules that you have developed.

Appendix 1.1B: Developing a Set of Solubility Rules: Lab Activity (Teacher Notes)

Introduction

Have student groups perform the lab activity using the solutions in either Set A or Set B below and then share their observations.

Where appropriate, 1.0 mol/L solutions can be prepared instead of 0.1 mol/L solutions. Involving students in the preparation of solutions is desireable. It may be clearer for students if the ions that participate in the reactions come from separate solutions. For instance, in Set A, a solution of 0.1 mol/L NaCl could be the source of Na^+ ions, and 0.1 mol/L Na_2CO_3 acts as the source of CO_3^{2-} ions. These solutions would replace the following solution in Set A below: 2 × 0.1 mol/L solutions of sodium carbonate (Na_2CO_3) labelled Na^+ and CO_3^{2-}. The NH_4Cl is used for the NH_4^+ ions and the $(NH_4)_2SO_4$ is used as the source for SO_4^{2-} ions. For Set B, NaCl can be used as a source for sodium ions, and KCl can be used as a source for potassium ions.

If this strategy is not followed, students will no doubt observe "anomalous" precipitates (discrepant events) that were unexpected, and may be difficult to explain. To avoid confusion, teachers are encouraged to proceed according to the level of difficulty desired for students' explanations of results.

Solutions

Prepare solution sets of 25 mL dropper bottles.

Set A

1 × 0.1 mol/L solution of silver nitrate ($AgNO_3$) labelled Ag^+

2 × 0.1 mol/L solutions of barium chloride ($BaCl_2$) labelled Ba^{2+} and Cl^-

2 × 0.1 mol/L solutions of sodium carbonate (Na_2CO_3) labelled Na^+ and CO_3^{2-}

2 × 0.1 mol/L solutions of ammonium sulphate ($(NH_4)_2SO_4$) labelled NH_4^+ and SO_4^{2-}

2 × 0.1 mol/L solutions of calcium nitrate ($Ca(NO_3)_2$) labelled Ca^{2+} and NO_3^-

1 × 0.1 mol/L solution of potassium phosphate (K_3PO_4) labelled PO_4^{3-}

Set B

1 × 0.1 mol/L solution of zinc acetate ($Zn(C2H_3O_2)_2$) labelled Zn^{2+}

2 × 0.1 mol/L solutions of iron(III) chloride ($FeCl_3$) labelled Fe^{3+} and Cl^-

2 × 0.1 mol/L solutions of sodium hydroxide (NaOH) labelled Na+ and OH^-

1 × 0.1 mol/L solution of magnesium bromide ($MgBr_2$) labelled Mg^{2+}

1 × 0.1 mol/L solution of sodium bromide (NaBr) labelled Br^-

2 × 0.1 mol/L solutions of potassium carbonate (K_2CO_3) labelled K^+ and CO_3^{2-}

1 × 0.1 mol/L solution of sodium acetate ($NaC_2H_3O_2$) labelled $C_2H_3O_2^-$

Appendix 1.1A: Developing a Set of Solubility Rules: Lab Activity (Teacher Notes)

(continued)

Probable Results

Set A					
	Cl^-	CO_3^{2-}	SO_4^{2-}	NO_3^-	PO_4^{3-}
Ag^+	PPT	PPT	PPT	NP	PPT
Ba^{2+}	NP	PPT	PPT	NP	PPT
Na^+	NP	NP	NP	NP	NP
NH_4^+	NP	NP	NP	NP	NP
Ca^{2+}	NP	PPT	PPT	NP	PPT

PPT = precipitate; NP = no precipitate

1. a) The cations that did not form any precipitates were Na^+ and NH_4^+.

   b) Cl^- formed a precipitate with Ag^+.

   CO_3^{2-} formed a precipitate with Ag^+, Ba^{2+}, and Ca^{2+}.

   SO_4^{2-} formed a precipitate with Ag^+, Ba^{2+}, and Ca^{2+}.

   Note: Ag_2SO_4 is sparingly soluble, so students may or may not see a precipitate.

   NO_3^- did not form a precipitate with any of the cations.

   PO_4^{3-} formed a precipitate with Ag^+, Ba^{2+}, and Ca^{2+}.

Appendix 1.1B: Developing a Set of Solubility Rules: Lab Activity (Teacher Notes)

(continued)

Set B					
	Cl^-	OH^-	Br^-	CO_3^{2-}	$C_2H_3O_2^-$
Zn^{2+}	NP	PPT	NP	PPT	NP
Fe^{3+}	NP	PPT	NP	NP	NP
Na^+	NP	NP	NP	NP	NP
Mg^{2+}	NP	PPT	NP	PPT	NP
K^+	NP	NP	NP	NP	NP

1. a) The cations that did not form any precipitates were Na^+ and K^+.

   b) Cl^- did not form a precipitate with any of the cations.

   OH^- formed a precipitate with Zn^{2+}, Fe^{3+} and Mg^{2+}.

   Br^- did not form a precipitate with any of the cations.

   CO_3^{2-} formed a precipitate with Zn^{2+} and Mg^{2+}.

   $C_2H_3O_2^-$ did not form a precipitate with any of the cations.

2. Solubility Rules

   a) Most nitrate (NO_3^-) salts are soluble.

   b) Most salts containing the alkali metal ions (Li^+, Na^+, K^+, Rb^+, Cs^+) and the ammonium ion (NH_4^+) are soluble.

   c) Most chloride (Cl^-), bromide (Br^-), and iodide (I^-) salts are soluble. Notable exceptions are salts containing the ions Ag^+, Pb^{2+}, and Hg_2^{2+}.

   d) Most sulphate (SO_4^{2-}) salts are soluble. Notable exceptions are $BaSO_4$, $PbSO_4$, $HgSO_4$, and $CaSO_4$.

   e) Most hydroxide (OH^-) salts are only slightly soluble. The important soluble hydroxides are NaOH and KOH. The compounds $Ba(OH)_2$, $Sr(OH)_2$, and $Ca(OH)_2$ are marginally soluble.

   f) Most sulphide (S^{2-}), carbonate (CO_3^{2-}), chromate (CrO_4^{2-}), and phosphate (PO_4^{3-}) salts are only slightly soluble.

Appendix 1.2: Solubility Rules

Negative Ions	Positive Ions	Solubility
Essentially all	Alkali ions (Li^+, Na^+, K^+, Rb^+, Cs^+)	Soluble
Essentially all	Hydrogen ion $H^+_{(aq)}$	Soluble
Essentially all	Ammonium ion (NH_4^+)	Soluble
Nitrate, NO_3^-	Essentially all	Soluble
Acetate, CH_3COO^-	Essentially all (except Ag^+)	Soluble
Chloride, Cl^- Bromide, Br^- Iodide, I^-	Ag^+, Pb^{2+}, Hg_2^{2+}, Cu^+, Tl^+	Low solubility
	All others	Soluble
Sulphate, SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Ra^{2+}	Low solubility
	All others	Soluble
Sulphide, S^{2-}	Alkali ions, $H^+_{(aq)}$, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+}	Soluble
	All others	Low solubility
Hydroxide, OH^-	Alkali ions, $H^+_{(aq)}$, NH_4^+, Sr^{2+}, Ba^{2+}, Ra^{2+}, Tl^{2+}	Soluble
	All others	Low solubility
Phosphate, PO_4^{3-} Carbonate, CO_3^{2-} Sulphite, SO_3^{2-}	Alkali ions, $H^+_{(aq)}$, NH_4^+	Soluble
	All others	Low solubility
Chromate, CrO_4^{2-}	Ba^{2+}, Sr^{2+}, Pb^{2+}, Ag^+	Low solubility
	All others	Soluble

Appendix 1.3: Predicting Precipitation Reactions

Use a solubility rules table to predict precipitation reactions.

a) Predict the products of the reactions in the following examples.

b) Write a balanced molecular equation and check the table for the solubility of the products.

c) Write a total ionic equation.

d) Write a net ionic equation.

Example 1

$AlCl_3$ reacts with KOH

a) Al^{3+} combines with OH^- to form $Al(OH)_3$, and K^+ combines with Cl^- to form KCl.

b) The balanced molecular equation will be

$$AlCl_{3(aq)} + 3KOH_{(aq)} \longrightarrow Al(OH)_{3(s)} + 3KCl_{(aq)}$$

Notice from the solubility table that the Al^{3+} ion is insoluble with the OH^- ion, thus forming a precipitate.

c) Compounds that are written as aqueous are broken down to their respective cations and anions. Solids are written in molecular form.

$$Al^{3+}{}_{(aq)} + 3Cl^-{}_{(aq)} + 3K^+{}_{(aq)} + 3OH^-{}_{(aq)} \longrightarrow Al(OH)_{3(s)} + 3K^+{}_{(aq)} + 3Cl^-{}_{(aq)}$$

d) Ions that are common to both sides of the reaction are called spectator ions. These ions are cancelled when writing the net ionic equation.

$$Al^{3+}{}_{(aq)} + \cancel{3Cl^-{}_{(aq)}} + \cancel{3K^+{}_{(aq)}} + 3OH^-{}_{(aq)} \longrightarrow Al(OH)_{3(s)} + \cancel{3K^+{}_{(aq)}} + \cancel{3Cl^-{}_{(aq)}}$$

The net ionic equation would be

$$Al^{3+}{}_{(aq)} + 3OH^-{}_{(aq)} \longrightarrow Al(OH)_{3(s)}$$

Appendix 1.3: Predicting Precipitation Reactions *(continued)*

Example 2

$AgNO_3$ reacts with CaI_2

a) Ag^+ combines with I^- to form AgI, and Ca^{2+} combines with NO_3^- to form $Ca(NO_3)_2$.

b) The balanced molecular equation will be

$$2AgNO_{3(aq)} + CaI_{2(aq)} \longrightarrow 2AgI_{(s)} + Ca(NO_3)_{2(aq)}$$

Notice from the solubility table that the Ag^+ ion is insoluble with the I^- ion, thus forming a precipitate.

c) Compounds that are written as aqueous are broken down to their respective cations and anions. Solids are written in molecular form.

$$2Ag^+_{(aq)} + 2NO_3{}^-_{(aq)} + Ca^{2+}_{(aq)} + 2I^-_{(aq)} \longrightarrow 2AgI_{(s)} + Ca^{2+}_{(aq)} + 2NO_3{}^-_{(aq)}$$

d) Ions that are common to both sides of the reaction are called spectator ions. These ions are cancelled when writing the net ionic equation.

$$2Ag^+_{(aq)} + 2\cancel{NO_3{}^-_{(aq)}} + \cancel{Ca^{2+}_{(aq)}} + 2I^-_{(aq)} \longrightarrow 2AgI_{(s)} + \cancel{Ca^{2+}_{(aq)}} + 2\cancel{NO_3{}^-_{(aq)}}$$

The net ionic equation would be

$$2Ag^+_{(aq)} + 2I^-_{(aq)} \longrightarrow 2AgI_{(s)}$$

Appendix 1.4: Colour Chart for Ions in Aqueous Solutions

Ion	Symbol	Colour
Chrome(II)	Cr^{2+}	Blue
Chrome(III)	Cr^{3+}	Green
Cobalt(II)	Co^{2+}	Pink
Chromate	CrO_4^{2-}	Yellow
Dichromate	$Cr_2O_7^{2-}$	Orange
Copper(I)	Cu^{+}	Green
Copper(II)	Cu^{2+}	Blue
Iron(II)	Fe^{2+}	Green
Iron(III)	Fe^{3+}	Pale yellow
Manganese(II)	Mn^{2+}	Pink
Permanganate	MnO_4^{-}	Purple
Nickel(II)	Ni^{2+}	Green

Appendix 1.5: Identifying Unknown Solutions (Teacher Notes and Preparation Guide)

Purpose

Present student groups with four unknown solutions. Their job will be to identify each unknown solution using only a spot plate, a stir stick, a chart showing the colour of ions in aqueous solutions, a table of solubility rules, and the solutions themselves.

Solutions

The sets of solutions that students will use could include 0.1 mol/L solutions of the following:

Set 1: $Ba(NO_3)_2$, $NaOH$, Na_2CO_3, $CuSO_4$

Set 2: $Co(NO_3)_2$, Na_3PO_4, Na_2SO_4, $AgNO_3$

Set 3: $Cr_2(SO_4)_3$, $MnSO_4$, $Ba(NO_3)_2$, $Zn(NO_3)_2$

Set 4: $Fe(NO_3)_3$, KI, $Pb(NO_3)_2$, $NaOH$

Set 5: $NiSO_4$, Na_2CO_3, $MnSO_4$, $NaCl$

Set 6: $CuSO_4$, $NaCl$, Na_3PO_4, $Zn(NO_3)_2$

Questions

Students must correctly identify the four solutions and explain how they identified each of the solutions using the solubility rules.

1. Using a chart that shows the colour of common ions in aqueous solutions, can you identify any of your unknowns based on this information? Explain.
2. Which solutions that you mixed formed a precipitate? Can you identify any of the unknown solutions based on this result? Explain.
3. Are there any reactions that have no precipitate formation? Can you identify any of the unknown solutions based on this result? Explain.

Preparation Guide

Prepare 0.1 mol/L solutions of each of the following.

Set 1

Solution 1: 2.613 g of $Ba(NO_3)_2$ in 100 mL of solution

Solution 2: 0.40 g of $NaOH$ in 100 mL of solution

Solution 3: 1.06 g of Na_2CO_3 in 100 mL of solution

Solution 4: 2.50 g of $CuSO_4 \cdot 5H_2O$ in 100 mL of solution

Appendix 1.5: Identifying Unknown Solutions (Teacher Notes and Preparation Guide) *(continued)*

Set 2

Solution 1: 2.91 g of $Co(NO_3)_2 \cdot 6H_2O$ in 100 mL of solution

Solution 2: 2.90 g of $Na_3PO_4 \cdot 7H_2O$ in 100 mL of solution

Solution 3: 1.421 g of Na_2SO_4 in 100 mL of solution

Solution 4: 1.699 g of $AgNO_3$ in 100 mL of solution

Set 3

Solution 1: 3.60 g of $Cr_2(SO_4)_3$ in 100 mL of solution

Solution 2: 1.69 g of $MnSO_4 \cdot H_2O$ in 100 mL of solution

Solution 3: 2.613 g of $Ba(NO_3)_2$ in 100 mL of solution

Solution 4: 2.97 g of $Zn(NO_3)_2 \cdot 6H_2O$ in 100 mL of solution

Set 4

Solution 1: 4.04 g of $Fe(NO_3)_3 \cdot 9H_2O$ in 100 mL of solution

Solution 2: 1.66 g of KI in 100 mL of solution

Solution 3: 3.312 g of $Pb(NO_3)_2$ in 100 mL of solution

Solution 4: 0.40 g of NaOH in 100 mL of solution

Set 5

Solution 1: 2.63 g of $NiSO_4 \cdot 6H_2O$ in 100 mL of solution

Solution 2: 1.06 g of Na_2CO_3 in 100 mL of solution

Solution 3: 1.69 g of $MnSO_4 \cdot H_2O$ in 100 mL of solution

Solution 4: 0.584 g of NaCl in 100 mL of solution

Set 6

Solution 1: 2.50 g of $CuSO_4 \cdot 5H_2O$ in 100 mL of solution

Solution 2: 0.584 g of NaCl in 100 mL of solution

Solution 3: 2.90 g of $Na_3PO_4 \cdot 7H_2O$ in 100 mL of solution

Solution 4: 2.97 g of $Zn(NO_3)_2 \cdot 6H_2O$ in 100 mL of solution

Appendix 1.6A: Process Notes for Writing Net Ionic Equations (Teacher Notes)

Solve the problem, showing all steps.	**Use words to describe each step of the solution.**
$Na_2S + FeSO_4 \rightarrow Na_2SO_4 + FeS$	Step 1: Predict the products of the double displacement reaction and ensure that the equation is balanced.
$Na_2S_{(aq)} + FeSO_{4(aq)} \rightarrow$ $Na_2SO_{4(aq)} + FeS_{(s)}$	Step 2: Use (aq) and (s) to identify each species as being soluble or slightly soluble (i.e., write the molecular equation).
$2Na^+{}_{(aq)} + S^{2-}{}_{(aq)} + Fe^{2+}{}_{(aq)} + SO_4{}^{-}{}_{(aq)} \rightarrow$ $2Na^+{}_{(aq)} + SO_4{}^{2-}{}_{(aq)} + FeS_{(s)}$	Step 3: Write the ionic equation by breaking up soluble species into their ions.
$\cancel{2Na^+{}_{(aq)}} + S^{2-}{}_{(aq)} + Fe^{2+}{}_{(aq)} + \cancel{SO_4{}^{-}{}_{(aq)}} \rightarrow$ $\cancel{2Na^+{}_{(aq)}} + \cancel{SO_4{}^{2-}{}_{(aq)}} + FeS_{(s)}$	Step 4: Cancel out all spectator ions and rewrite the equation.
$S^{2-}{}_{(aq)} + Fe^{2+}{}_{(aq)} \rightarrow FeS(s)$	This gives the net ionic equation.

Appendix 1.6B: Process Notes for Writing Net Ionic Equations (BLM)

Solve the problem, showing all steps	**Use words to describe each step of the solution process.**
$BaCl_2 + Na(PO_4)_3 \rightarrow$	Step 1: Predict the products of the double displacement reaction and ensure that the equation is balanced. Step 2: Use (aq) and (s) to identify each species as being soluble or slightly soluble (i.e., write the molecular equation). Step 3: Write the ionic equation by breaking up soluble species into their ions. Step 4: Cancel out all spectator ions and rewrite the equation. This gives the net ionic equation.

Appendix 1.7A: Titration: Lab Activity

Purpose

Titrations are procedures that are usually used to determine the unknown concentrations of substances. In this lab activity, you will add drops of a known concentration of sodium hydroxide to a beaker containing a known concentration of sulphuric acid until neutralization occurs. The number of moles of each reactant can then be calculated from the volumes present, so that their ratio can be compared to the ratio of coefficients in the balanced equation.

Materials

50 mL beaker

three micropipettes

phenolphthalein indicator

10 mL graduated cylinder

distilled water

0.1 mol/L sodium hydroxide (NaOH)

0.1 mol/L sulphuric acid (H_2SO_4)

Procedure

1. Using the 10 mL graduated cylinder and a micropipette, count and record the number of drops required to obtain 1.0 mL of distilled water. Perform this process a total of three times.

   Note: For the best, most reproducible results, hold the micropipette vertically, and squeeze the bulb slowly and gently. Avoid introducing air bubbles into the stem of the pipette, as they will result in half or quarter drops.

2. Add 5 mL of distilled water and one drop of phenolphthalein indicator to a 50 mL beaker. Swirl the beaker well.

3. Using a second micropipette (to avoid contamination of the solutions), add 20 drops of 0.1 mol/L H_2SO_4 to the beaker. Swirl the solution carefully.

4. Using a third micropipette, add the 0.1 mol/L NaOH drop by drop, until the addition of one drop of the base permanently changes the colour of the solution. Be sure to swirl the beaker gently after each drop is added. Record the number of drops required to reach the endpoint of the titration.

   Note: The endpoint of the titration occurs when one drop of an acid or a base permanently changes the colour of the indicator used in the titration.

5. Rinse the contents of the beaker down the sink with plenty of water (the final rinse should be with distilled water), and perform steps 2 through 4 a total of three times.

Appendix 1.7A: Titration: Lab Activity *(continued)*

Note: The trials should agree with one another to within one drop. If you make a mistake, miss the endpoint, or lose count of the drops, perform another trial. Do not erase the results, but make note of what went wrong.

Qualitative Observations

- Describe each solution before reaction occurs.
- Describe the solution after adding the drops of phenolphthalein.

Quantitative Data Tables

Trial	Drops of Water in 1.0 mL
1	
2	
3	
Average	

Volume of Water Used (mL)	Drops of Sulphuric Acid	Volume of Sulphuric Acid (mL)	Drops of Sodium Hydroxide	Volume of Sodium Hydroxide (mL)
5	20			
5	20			
5	20			
Average				

Appendix 1.7A: Titration: Lab Activity *(continued)*

Calculations

1. Write a balanced molecular equation for the reaction.
2. Draw a particulate representation of the balanced reaction.
3. Calculate the average number of drops required to obtain 1.0 mL of distilled water.
4. Using the data obtained in step 2 of the procedure, calculate the volume of NaOH added in each trial.
5. Calculate the average number of moles of NaOH required to neutralize the sample of H_2SO_4.
6. Using the data obtained in step 2, calculate the volume of H_2SO_4 added in each trial.
7. Using your balanced equation, determine the average number of moles present in the sample of H_2SO_4.
8. Use the coefficients in the balanced equation to determine the ratio of moles between the sodium hydroxide and the sulphuric acid.
9. Use the number of moles obtained in steps 4 and 5 of the procedure to determine the ratio of moles between the sodium hydroxide and the sulphuric acid.

Conclusion

State the stoichiometric relationship between the sodium hydroxide and the sulphuric acid.

Questions

1. a) Write a balanced molecular equation for the reaction between barium hydroxide and sulphuric acid.
   b) Use the coefficients in the balanced equation to calculate the volume of barium hydroxide required to react with 20 mL of sulphuric acid.
2. a) Write a balanced molecular equation for the reaction between aluminum hydroxide and sulphuric acid.
   b) Use the coefficients in the balanced equation to calculate the volume of aluminum hydroxide required to react with 30 mL of sulphuric acid.

Sources of Error

What possible errors could have occurred in your lab activity?

Appendix 1.7B: Titration: Lab Activity (Teacher Notes)

Purpose

To demonstrate the stoichiometry of a neutralization reaction between a strong acid and a strong base.

Qualitative Observations

Distilled water:	clear, colourless liquid
Sulphuric acid:	clear, colourless liquid
Sodium hydroxide:	clear, colourless liquid
Phenolphthalein:	clear, colourless liquid

Quantitative Data Tables

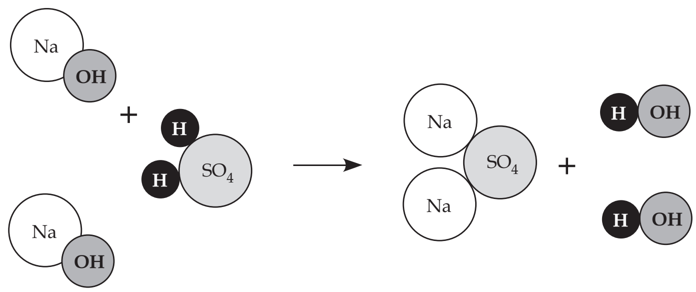
Trial	Drops of Water in 1.0 mL
1	23
2	24
3	23
Average	23

Volume of Water Used (mL)	Drops of Sulphuric Acid	Volume of Sulphuric Acid (mL)	Drops of Sodium Hydroxide	Volume of Sodium Hydroxide (mL)
5	20	0.858	69	2.96
5	20	0.858	68	2.92
5	20	0.858	70	3.00
Average	20	0.858	69	2.96

Appendix 1.7B: Titration: Lab Activity (Teacher Notes) *(continued)*

Calculations

1. $2NaOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$
2. Molecule size is not a true representation of the actual size of the compound.



3. (23 + 24 + 23)/3 = 23.3 drops
4. volume NaOH = (1 mL/23.3 drops) × 69 drops = 2.96 mL of NaOH
5. moles NaOH = 0.10 mole/L × 2.96 × 10^{-3} L = 0.000296 moles NaOH
6. volume H_2SO_4 = (1 mL/23.3 drops) × 20 drops = 0.858 mL of H_2SO_4
7. moles H_2SO_4 = 0.10 mole/L × 0.858 × 10^{-3} L = 0.0000858 moles H_2SO_4
8. coefficient NaOH/coefficient H_2SO_4 = 2/1 = 2
9. moles NaOH/moles H_2SO_4 = 0.000296/0.0000858 = 3.45

Conclusion

Answers will vary. For example, the stoichiometric relationship between the sodium hydroxide and the sulphuric acid in the balanced equation is 2 to 1, while the experimental relationship is 3.45 to 1.

Questions

1. a) $Ba(OH)_{2(aq)} + H_2SO_{4(aq)} \longrightarrow BaSO_{4(aq)} + 2H_2O_{(l)}$
   b) The volume of barium hydroxide required to react with 20 mL of sulphuric acid is 20 mL.
2. a) $2Al(OH)_{3(aq)} + 3H_2SO_{4(aq)} \longrightarrow Al_2(SO_4)_{3(aq)} + 6H_2O_{(l)}$
   b) The volume of aluminum hydroxide required to react with 30 mL of sulphuric acid is 20 mL.

Sources of Error

Sources of error could include calibration of the micropipette and graduated cylinder, as well as the accuracy of the concentrations of the solutions used.

Appendix 1.8: Process Notes for Balancing Neutralization Reactions

Solve the problem, showing all the steps.	**Use words to describe each step of the solution process.**
$H_2SO_4 + NaOH \rightarrow Na_2SO_4 + H_2O$	Step 1: Predict the products of the neutralization reaction. Remember that a salt and water are formed.
$H_2SO_{4(aq)} + 2NaOH(aq) \rightarrow$ $Na_2SO_{4(aq)} + 2H_2O(l)$	Step 2: Ensure that the equation is balanced. Use (aq) and (l) to identify each species as being soluble or slightly soluble (i.e., write the molecular equation).
$2H^+_{(aq)} + SO_4{}^{2-}{}_{(aq)} + 2Na^+_{(aq)} + 2OH^-{}_{(aq)}$ $\rightarrow 2Na^+_{(aq)} + SO_4{}^{2-}{}_{(aq)} + 2H_2O_{(l)}$	Step 3: Write a total ionic equation, showing all ions that are in solution.
$\cancel{2}H^+_{(aq)} + \cancel{SO_4{}^{2-}{}_{(aq)}} + \cancel{2Na^+_{(aq)}} + \cancel{2}OH^-{}_{(aq)}$ $\rightarrow \cancel{2Na^+_{(aq)}} + \cancel{SO_4{}^{2-}{}_{(aq)}} + 2H_2O_{(l)}$	Step 4: Cancel the spectator ions.
$H^+_{(aq)} + OH^-{}_{(aq)} \rightarrow H_2O_{(l)}$	Step 5: Write the net ionic equation.

Appendix 1.9A: Test Tube Mystery: Lab Activity (Guidelines)

Purpose

Chemists, like detectives, attempt to identify unknowns through a process of careful and creative analysis. This usually involves observing the colours, odours, and reactions of unknown substances and comparing them with those of known substances. In this experiment, you will try to identify 12 different chemical compounds by reacting them with each other, observing the results, and comparing the results with the known characteristics of some common chemicals.

Chemical Compounds

The 12 chemicals used in this experiment are listed below (in no particular order):

potassium chromate (K_2CrO_4)

aluminum chloride ($AlCl_3$)

sodium carbonate (Na_2CO_3)

sodium acetate ($NaCH_3COO$)

hydrochloric acid (HCl)

sodium hydroxide (NaOH)

ammonium hydroxide (NH_4OH)

iron(III) nitrate ($Fe(NO_3)_3$)

silver nitrate ($AgNO_3$)

copper(II) sulphate ($CuSO_4$)

nickel(II) chloride ($NiCl_2$)

lead(II) nitrate ($Pb(NO_3)_2$)

Research and Plan

Before starting the lab activity, you will have to do extensive research on the characteristic colours of the solutions, any distinguishing odours, their flame-test colours, and the colours of any precipitates that may be created through the combination of each different species. Your written plan must include a data table grid that includes each species, the solution and flame-test colours, the colours of potential precipitates, and any other information that you think will help to identify your unknowns.

Appendix 1.9A: Test Tube Mystery: Lab Activity (Guidelines) *(continued)*

Materials

On the day of the lab activity, you will be provided with the following materials:

- 12 test tubes containing 8 mL each of different solutions
- well plates
- stir sticks
- cotton swabs/flame-test wires/moist wooden splints
- Bunsen burners
- matches
- litmus paper
- 10 micropipettes
- gloves
- distilled water

Avoid running out of your samples, as you will not be provided with any more. Do not assume that solution sets other groups are using are numbered in the same way—they are not!

Lab Write-up

After recording all your observations in the lab activity, you will attempt to identify each of the unknowns. A formal lab write-up must include a logical explanation of how you determined the identity of each test tube. This will include net ionic equations for any precipitates you saw.

Caution

All solutions must be treated as if they were poisonous and corrosive. Avoid inhaling any fumes. Some reactions may occur very quickly, while others will occur more slowly. Observe each reaction for at least two minutes before disposing of the products. Gas evolution (bubbling) will be immediate. Rinse off your stir stick after each use. As time will be limited, use your time wisely.

Appendix 1.9B: Test Tube Mystery: Lab Activity (Preparation Guide)

Teachers can prepare the solutions for this lab activity in advance or have students prepare them. Prepare a solution, given the amount of solute (in grams) and the volume of solution (in millilitres), and determine the concentration in moles/litre.

Materials

well plates
stir sticks
cotton swabs/flame-test wires/moist wooden splints
Bunsen burners
matches
litmus paper
micropipettes (10 per group)
gloves
distilled water
test tube rack
test tubes (12 × 10/group = 120 test tubes)
test tube stoppers or plastic wrap to cover the test tubes
100 mL solutions of the following 12 solutions:

- 0.2 mol/L K_2CrO_4—to prepare, dissolve 3.88 g of K_2CrO_4 in 100 mL of distilled water
- 1.0 mol/L $AlCl_3 \cdot 6H_2O$—dissolve 24.14 g of $AlCl_3$ in 100 mL of distilled water
- 1.0 mol/L Na_2CO_3—dissolve 10.6 g of Na2CO3 in 100 mL of distilled water
- 1.0 mol/L $NaCH_3COO \cdot 3H_2O$—dissolve 13.61 g of $NaCH_3COO$ in 100 mL of distilled water
- 6.0 mol/L HCl—mix 49.6 mL in 100 mL of distilled water
- 6.0 mol/L NaOH—dissolve 24.0 g of NaOH in 100 mL of distilled water
- 6.0 mol/L NH_4OH—mix 40.5 mL in 100 mL of distilled water
- 0.1 mol/L $Fe(NO3)_3 \cdot 9H_2O$—dissolve 4.04 g of $Fe(NO_3)_3 \cdot 9H_2O$ in 100 mL of distilled water
- 0.1 mol/L $AgNO_3$—dissolve 1.70 g of $AgNO_3$ in 100 mL of distilled water
- 0.1 mol/L $CuSO_4$—dissolve 2.50 g of $CuSO_4 \cdot 5H_2O$ in 100 mL of distilled water
- 0.1 mol/L $NiCl_2 \cdot 6H_2O$—dissolve 2.38 g of $NiCl_2$ in 100 mL of distilled water
- 0.1 mol/L $Pb(NO_3)_2$—dissolve 3.31 g of $Pb(NO_3)_2$ in 100 mL of distilled water

Appendix 1.9B: Test Tube Mystery: Lab Activity (Preparation Guide) *(continued)*

100 mL solutions should be prepared in advance of the lab activity.

Test tubes can be pre-labelled with the information, set 1, test tube 1, and so on.

Students are given an 8 to 10 mL sample of each solution (12 different test tubes) that are contained in a test tube rack.

A suggested teacher key is given for setting up each set of test tubes.

Solution Set Key

Substance	Group 1 and Group 9	Group 2 and Group 10	Group 3 and Group 6	Group 4 and Group 7	Group 5 and Group 8
K_2CrO_4	3	1	2	3	4
$AlCl_3$	5	5	6	7	8
Na_2CO_3	6	9	10	11	12
$NaCH_3COO$	1	4	1	2	3
HCl	11	8	5	6	7
NaOH	4	12	9	10	11
NH_4OH	7	3	4	1	2
$Fe(NO_3)_3$	10	7	8	5	6
$AgNO_3$	2	11	12	9	10
$CuSO_4$	8	2	3	4	1
$NiCl_2$	12	6	7	8	5
$Pb(NO_3)_2$	9	10	11	12	9

Appendix 1.9C: Test Tube Mystery: Lab Activity (Teacher Key 1)

Example of what students could have prepared before doing the lab activity.

	Kr_2CrO_4	$AlCl_3$	Na_2CO_3	$NaCH_3COO$	HCl	$NaOH$	NH_4OH	$Fe(NO_3)_3$	$AgNO_3$	$CuSO_4$	$NiCl_2$	$Pb(NO_3)_2$
Kr_2CrO_4		NP	NP	NP	NP	NP	NP	NP	Ag_2CrO_4	NP	NP	$PbCrO_4$
$AlCl_3$			$Al_2(CO_3)_3$	NP	NP	$Al(OH)_3$	$Al(OH)_3$	NP	$AgCl$	NP	N	$PbCl_2$
Na_2CO_3				NP	gas	NP	NP	$Fe_2(CO_3)_3$	$AgCO_3$	$CuCO_3$	$NiCO_3$	$PbCO_3$
$NaCH_3COO$					NP	NP	NP	NP	$AgCH_3COO$	NP	NP	NP
HCl						NP	NP	NP	$AgCl$	NP	NP	$PbCl_2$
$NaOH$							NP	$Fe_2(OH)_3$	$AgOH$	$Cu(OH)_2$	$Ni(OH)_2$	$Pb(OH)_2$
NH_4OH								$Fe_2(OH)_3$	$AgOH$	$Cu(OH)_2$	$Ni(OH)_2$	$Pb(OH)_2$
$Fe(NO_3)_3$									NP	NP	NP	NP
$AgNO_3$										Ag_2SO_4	$AgCl$	NP
$CuSO_4$											NP	$PbSO_4$
$NiCl_2$												$PbCl_2$
$Pb(NO_3)_2$												

NP = no precipitate

Appendix 1.9D: Test Tube Mystery: Lab Activity (Teacher Key 2)

Substance	Identifying Colour/ Colour of Solution	Colour in Litmus Paper	Flame-Test Colour	Reacts with	To Make	Colour of Precipitate
K_2CrO_4	Yellow	Blue	Violet	$AgNO_3$ $Pb(NO_3)_2$	Ag_2CrO_4 $PbCrO_4$	Brick red Yellow
$AlCl_3$		Neutral		Na_2CO_3 NaOH $AgNO_3$ $Pb(NO_3)_2$	$Al_2(CO_3)_3$ $Al(OH)_3$ AgCl $PbCl_2$	White White White* Yellow
Na_2CO_3		Blue	Yellow	HCl $Fe(NO_3)_3$ $AgNO_3$ $CuSO_4$ $NiCl_2$ $PB(NO_3)_2$ $AlCl_3$	bubbles $Fe(CO_3)_3$ $AgCO_3$ $CuCO_3$ $NiCO_3$ $PbCO_3$ $Al_2(CO_3)_3$	Bubbles White White* White White White White
$NaCH_3COO$		Blue	Yellow	$AgNO_3$	$AgCH_3COO$	White*
HCl		Pink		$AgNO_3$ $Pb(NO_3)_2$ Na_2CO_3	AgCl $PbCl_2$ Bubbles	White* Yellow Bubbles
NaOH		Blue		$Fe(NO_3)_3$ $AgNO_3$ $CuSO_4$ $NiCl_2$ $Pb(NO_3)_2$ $AlCl_3$	$Fe(OH)_3$ AgOH $Cu(OH)_2$ $Ni(OH)_2$ $Pb(OH)_2$ $Al(OH)_3$	White Brown* White White White White
NH_4OH	Strong odour	Blue		$Fe(NO_3)_3$ $AgNO_3$ $CuSO_4$ $NiCl_2$ $Pb(NO_3)_2$ $AlCl_3$	$Fe(OH)_3$ AgOH $Cu(OH)_2$ $Ni(OH)_2$ $Pb(OH)_2$ $Al(OH)_3$	White Brown* White White White White
$Fe(NO_3)_3$	Pale yelow	Neutral		Na_2CO_3 NaOH	$Fe_2(CO_3)_3$ $Fe(OH_3)$	White White/Brown

* Most Ag precipitates start out a white to greyish-white colour, but turn purple/brown/black over time.

continued

Appendix 1.9D: Test Tube Mystery: Lab Activity (Teacher Key 2) *(continued)*

Substance	Identifying Colour/ Colour of Solution	Colour in Litmus Paper	Flame-Test Colour	Reacts with	To Make	Colour of Precipitate
$AgNO_3$		Neutral		$NiCl_2$ K_2CrO_4 $AlCl_3$ Na_2CO_3 $NaCH_3COO$ HCl NaOH $CuSO_4$	AgCl Ag_2CrO4 AgCl Ag_2CO3 $AgCH_3COO$ AgCl AgOH Ag_2SO_4	White* Brick red White* White* White* White* Brown White*
$CuSO_4$	Blue	Neutral	Bluish-green	$Pb(NO_3)_2$ Na_2CO_3 NaOH	$PbSO_4$ $CuCO_3$ $Cu(OH)_2$	White White White
$NiCl_2$	Green/ Green-blue	Neutral		$Pb(NO_3)_2$ Na_2CO_3 NaOH $AgNO_3$	$PbCl_2$ $NiCO_3$ $Ni(OH)_2$ AgCl	Yellow White White White*
$Pb(NO_3)_2$		Neutral	Bluish-white	K_2CrO_4 $AlCl_3$ Na_2CO_3 HCl NaOH $CuSO_4$ $NiCl_2$	$PbCrO_4$ $PbCl_2$ $PbCO_3$ $PBCl_2$ $Pb(OH)_2$ $PbSO_4$ $PbCl_2$	Yellow Yellow White Yellow White White Yellow

* Most Ag precipitates start out a white to greyish-white colour, but turn purple/brown/black over time.

Appendix 1.10A: Compare and Contrast Oxidation and Reduction

Oxidation	Reduction
Historical Definition: *Example:*	Historical Definition: *Example:*
Present Definition: *Example:*	Present Definition: *Example:*
Mnemonic Device:	Mnemonic Device:
When Balancing a Redox Reaction: One substance is ________________ and it is also the ____________ agent. Its oxidation number _____________ .	When Balancing a Redox Reaction: One substance is ________________ and it is also the ____________ agent. Its oxidation number _____________ .

Appendix 1.10B: Compare and Contrast Oxidation and Reduction (Sample Response)

Oxidation	**Reduction**
Historical Definition: Gain of oxygen *Example:* $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ $CH_4 + 2O_2 \rightarrow CO_2 + 2H_20$	Historical Definition: Loss of oxygen *Example:* $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$
Present Definition: Loss of electrons *Example:* $Mg + S \rightarrow MgS$ (Magnesium undergoes oxidation)	Present Definition: Gain of electrons *Example:* $Mg + S \rightarrow MgS$ (Sulphur undergoes reduction)
Mnemonic Device: OIL LEO	Mnemonic Device: RIG GER
When Balancing a Redox Reaction: One substance is ___oxidized___ and it is also the ___reducing___ agent. Its oxidation number ___increases___ .	When Balancing a Redox Reaction: One substance is ___reduced___ and it is also the ___oxidizing___ agent. Its oxidation number ___decreases___ .

Appendix 1.11: Oxidation Number Rules

The rules for assigning oxidation numbers are identified below.

Rule 1: The oxidation number of any free atom (or multiple of itself) is 0.

Examples:

$C = 0$ $\quad$ $H_2 = 0$ $\quad$ $O_2 = 0$

Rule 2: An ion's oxidation number is its charge when in ionic form.

Examples:

$Na^{+} = +1$ $\quad$ $P^{3+} = +3$ $\quad$ $S^{2-} = -2$

Rule 3: In a compound or complex ion, the sum of all the oxidation numbers of each part must equal the total charge of that compound or complex ion.

Examples:

NaCl	$CaCl_2$	SO_4^{2-}
$+1 - 1 = 0$	$+2 - 1 - 1 = 0$	$+6 - 2 - 2 - 2 - 2 = -2$

Rule 4: The oxidation number of hydrogen is +1, except in metal hydrides where H is the anion (e.g., CaH_2 or LiH) and the oxidation number is –1.

Rule 5: The oxidation number of oxygen is –2, except in peroxides (e.g., H_2O_2, Na_2O_2) where it is –1, and when in combination with fluorine (O = +2).

Rule 6: The oxidation number of a Group IA (Group 1) element in a compound is +1.

Rule 7: The oxidation number of a Group IIA (Group 2) element in a compound is +2.

Rule 8: In most cases, the oxidation number of a Group VIIA (Group 17) element in a compound is –1.

Rule 9: Within a compound containing complex ions, each element's oxidation number can be determined using the charge on the complex ion.

Example:

- The compound $Ni_2(SO_4)_3$ contains the ions Ni^{3+} and SO_4^{2-}.
- Since the oxidation number of O is –2 according to rule 5 (for a total of –8), S must be +6 to result in –2 charge on the sulphate ion.
- Therefore,

$$\overset{?}{\underset{?}{Ni_2}}(\overset{+6}{\underset{+18}{S}}\overset{-2}{\underset{-24}{O_4}})_3 \quad \text{results in} \quad \overset{+3}{\underset{+6}{Ni_2}}(\overset{+6}{\underset{+18}{S}}\overset{-2}{\underset{-24}{O_4}})_3$$

Appendix 1.12A: Practical Applications of Redox Reactions (Research Report and Presentation)

To learn more about the practical applications of redox reactions taking place around you, your group (of no more than three students) will research one of the following topics, write a report on your findings, and give an oral presentation.

Topics

1. rocket fuels
2. fireworks
3. household bleach (i.e., stain removal and chlorination)
4. photography
5. metal recovery from ores
6. steelmaking
7. aluminum recycling
8. fuel cells
9. batteries
10. tarnish removal
11. fruit clocks
12. forensic blood detection using luminol
13. chemiluminescence/bioluminescence
14. electrolytic cleaning
15. electrodeposition
16. photochemical etching
17. antioxidants/preservatives

Resources

You will need access to resources such as the following:

- school, university, or public libraries
- Internet
- textbooks (see teacher)
- email communication (e.g., with an expert)
- magazines, journals, and newspapers
- interviews

Appendix 1.12A: Practical Applications of Redox Reactions (Research Report and Presentation) *(continued)*

Project Requirements

Your research, report, and presentation should include the following:

- Identify the redox application you have selected.
- Describe the redox reaction taking place, including information on the substances being oxidized and reduced, as well as the oxidizing agents and the reducing agents.
- Address the effects of the process on the environment, and the energy consumption involved in the process.

Submit your group's written project (of approximately two pages) to your teacher the day before your oral presentation. (Dates will be determined at the beginning of Topic 6: Electrochemistry, so that the written report can be copied for your classmates.)

The oral presentation should be approximately 10 minutes long and will be teacher-assessed. It will be followed by a brief question period (no longer than five minutes) in which the audience may ask clarifying questions.

Assessment

Please refer to the attached checklist and rubric for a more detailed list of the project requirements and assessment criteria for both the written report and the oral presentation.

Appendix 1.12B: Practical Applications of Redox Reactions (Sample Checklist and Assessment Rubric)

Checklist

Written Report

Have you included the following?

- ❑ Title page
- ❑ Bibliography with at least five sources
- ❑ The selected redox application
- ❑ All relevant redox reactions taking place
- ❑ All substances being oxidized and reduced, and any oxidizing agents and reducing agents
- ❑ The effects of the process on the environment
- ❑ The energy consumption involved in the process
- ❑ An introduction and a conclusion that connect the topic to redox chemistry
- ❑ Visual aids that help make the topic more easily understood
- ❑ Five possible test questions about the topic
- ❑ An answer key to the test questions

Oral Presentation

Have you included the following?

- ❑ Equal participation by all group members
- ❑ The selected redox application
- ❑ All relevant redox reactions taking place
- ❑ All substances being oxidized and reduced, and any oxidizing agents and reducing agents
- ❑ The effects of the process on the environment
- ❑ The energy consumption involved in the process
- ❑ An introduction and a conclusion that connect the topic to redox chemistry
- ❑ A logical flow and clear transitions
- ❑ Visual aids that help make the topic more easily understood

The information is

- ❑ Clear, accurate, concise
- ❑ Presented fully
- ❑ Interesting and easily understood

Appendix 1.12B: Practical Applications of Redox Reactions (Sample Checklist and Assessment Rubric) *(continued)*

Assessment Rubric

3	2	1	0
All relevant redox reactions are correct and are included in both the written and oral reports.	All relevant redox reactions are correct and are included in one of the reports.	Some relevant redox reactions are included in the written and oral reports, with some errors.	No redox reactions are included in the reports.
All substances being oxidized and reduced and all oxidizing and reducing agents are correctly listed.	Some substances being oxidized and reduced and all oxidizing and reducing agents are correctly listed.	Substances being oxidized and reduced and oxidizing and reducing agents are incorrectly listed.	Substances being oxidized and reduced and oxidizing and reducing agents are not listed.