Senior 2

Appendix 2: Chemistry in Action

BLM Blackline Master

Lewis Dot Diagrams (Electron Dot Diagrams)

A Lewis dot diagram is a convenient, shorthand method to represent an element and its valence electrons. The arrangement of elements on the periodic table yields more information about the electronic structures of atoms, and how those structures can help predict the properties of many elements. The arrangement of valence electrons is key to understanding how atoms behave in chemical reactions.

Lewis dot diagrams are diagrams in which dots or other small symbols are placed around the chemical symbol of an element to illustrate the valence electrons.

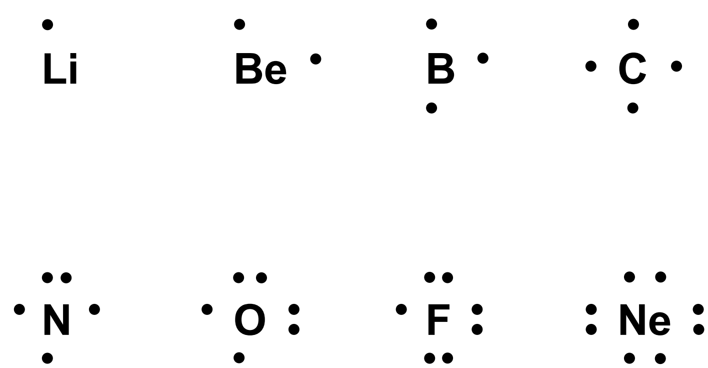
Note:

i. Each dot represents one valence electron.

ii. In the dot diagram, the element's symbol represents the core of the atom ﬁ nucleus plus all the inner electrons.

iii. Atoms in the same family (column) will have similar Lewis dot diagrams, except for helium (He) which has only two valence electrons.

The Lewis dot diagrams for the elements in the second period are as follows:



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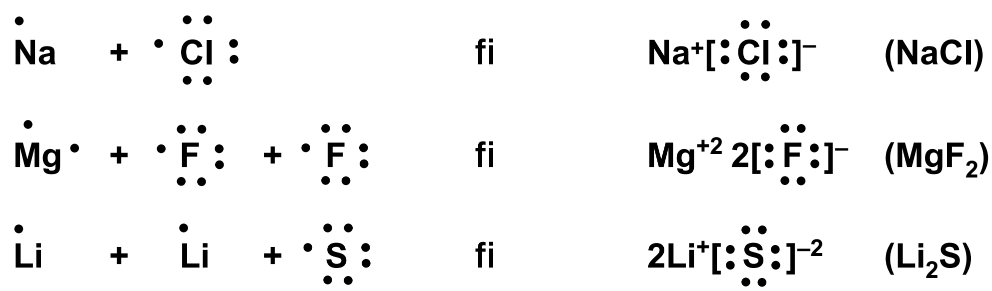
Chemical Bonds and Lewis Dot Diagrams

Atoms gain, lose, or share electrons to obtain full valence shells and become stable. The number of electrons present remains the same, but their arrangement changes when compounds form. Metal atoms tend to lose electrons, while non-metal atoms tend to gain or share electrons. Members of the noble gas family are chemically inert; they do not react.

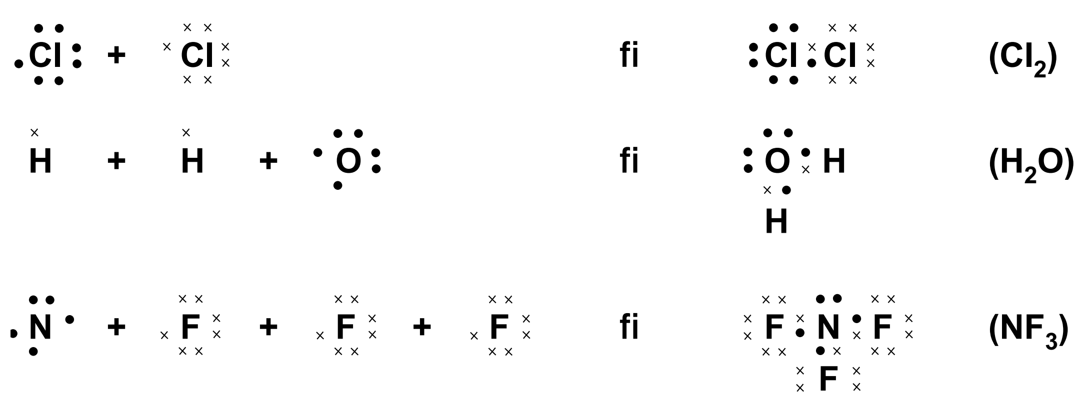
Ionic Bonds

Ionic bonds result when electrons are transferred from metal atoms to non-metal atoms. The metal atoms lose electrons to become positive ions, while the non-metal atoms gain electrons to become negative ions. The ions are then held together by the attraction of opposite charges in an ionic bond.



Covalent Bonds

Covalent bonds result when non-metal atoms share electrons. By overlapping their valence electron shells, the atoms share pairs of electrons. This increases the number of electrons in each atom's valence shell, so that the atoms appear to have full shells.



SLA
Student Learning Activity

Experiment: Properties of Acids and Bases

Purpose

1. To classify substances as acids or bases using their characteristic properties.
2. To determine the pH values of the acids and bases used.
3. To examine the reactivity of acids with metals.

Equipment/Materials

- Safety goggles
- Red and blue litmus paper
- Test tubes and test-tube racks
- Indicators (universal indicator, phenolphthalein, bromothymol blue)
- A pH meter or pH test paper
- Eyedroppers
- Microtray
- Samples of acids and bases (milk of magnesia, ammonia window cleaner, vinegar, lemon juice, tomato juice, drain cleaner, carbonated drinks, shampoo, black coffee, laundry detergent, milk, salt water, tap water, baking soda, apples)
- Samples of metals (copper, zinc, iron, magnesium)
- Hydrochloric acid (6M), acetic acid (6M), sodium hydroxide (0.5M), calcium hydroxide (0.5M)

Procedure

Part A: Effects of Acids and Bases on Indicators

1. Place five drops of each of the following into the wells of a microtray: 6M Hydrochloric acid, 6M acetic acid, 0.5M sodium hydroxide, and 0.5M calcium hydroxide.
2. Using a different piece of clean, dry, red litmus paper for each of the solutions, dip the end of a piece of red litmus paper into each solution. Record results in a data table.
3. Repeat step 2 using blue litmus paper. Record results in a data table.

Part B: Determine the pH Range of a Substance

1. Use the bromothymol, phenolphthalein, and universal indicators to measure the pH of the samples.
2. Add two drops of the bromothymol indicator to each sample from part A. Record your observations in a data table. Wash microtray and repeat for phenolphthalein and universal indicator.
3. Confirm your results by retesting each sample with a pH meter or pH test paper.

Part C: Determine the Reactivity of Acids with Metals

1. Place a small sample of each metal to be tested in the different wells of a clean, dry microtray.
2. Use an eyedropper to add five drops of the hydrochloric acid onto each sample of metal. Note any signs of chemical change and record your observations in a data table. Repeat using the acetic acid.

Part D: Determine the Acidity and Basicity of Household Substances

1. In different wells in your microtray, add five drops of any six of the following: vinegar, lemon juice, tomato juice, milk, household ammonia, cola, milk of magnesia, window cleaner, drain cleaner, shampoo, salt water, tap water.
2. Test each substance as you did in part A, using red and blue litmus paper.
3. Test each substance as you did in part B, using indicators and pH paper.
4. Record results in data table.

> **SAFETY**
>
> **Many household cleaners and hydrochloric acid solution(s) are corrosive or caustic. Any spills on the skin, in the eyes, or on clothing should be washed immediately with cold water. Report any spills to the teacher.**

Data Collection

Part A

Sample	Reaction with **red** litmus paper	Reaction with **blue** litmus paper
Hydrochloric acid		
Acetic acid		
Sodium hydroxide		
Calcium hydroxide		

Part B

Sample	Bromothymol Blue	Phenolphthalein	Universal indicator
Hydrochloric acid			
Acetic acid			
Sodium hydroxide			
Calcium hydroxide			

Part C

Metal	Reaction with hydrochloric acid	Reaction with acetic acid
Zinc		
Magnesium		
Iron		
Copper		

Part D

Household substance	Red litmus paper	Blue litmus paper	pH paper	Bromothymol Blue	Phenolphthalein	Universal indicator

Data Analysis

Tests	Acid properties	Base properties
Red litmus paper		
Blue litmus paper		
pH paper		
Bromothymol Blue		
Phenolphthalein		
Universal indicator		
Reaction with a metal		

Household sample	pH value	Acid or base?

Questions (respond on a separate sheet)

1. Can either red or blue litmus paper be used to identify acids? Explain.
2. How accurate are indicators for measuring pH?
3. What signs of chemical change were observed when acids were placed on metals?
4. Did all metals react similarly? Explain.
5. List the general properties of acids and bases.
6. Which of the household substances are acidic? Which are almost neutral? Which are basic?

TSM
Teacher Support Material

Experiment: Acids and Bases *(Teacher Notes)*

1. Review appropriate safety precautions.
2. As only binary acids have been discussed in class, students will not be expected to write chemical formulas for complex acids and bases. However, a wide range of acids and bases such as H_2SO_4, HNO_3, CH_3COOH, HCl, $Ca(OH)_2$, $NaOH$, NH_3 and $Ba(OH)_2$ can be used. The samples could be placed in dropper bottles for safer handling.
3. Instead of 6M acids, 3M acids could be used.
4. Some indicators that may be used are phenolphthalein, bromothymol blue, universal indicator, and methyl orange. Indicators that provide definite pH values will work best.
5. The metals used may include Cu, Zn, Fe, Al, Mg. Clean pieces or strips will work best.

Data Collection

Part A

Sample	Reaction with **red** litmus paper	Reaction with **blue** litmus paper
Hydrochloric acid	No change	Turns red
Acetic acid	No change	Turns red
Sodium hydroxide	Turns blue	No change
Calcium hydroxide	Turns blue	No change

Part B

Sample	Bromothymol Blue	Phenolphthalein	Universal indicator
Hydrochloric acid	Turns yellow	No change	Turns red
Acetic acid	Turns yellow	No change	Turns orange-red
Sodium hydroxide	No change	Turns pink	Turns violet
Calcium hydroxide	No change	Turns pink	Turns blue-violet

Part C

Metal	Reaction with hydrochloric acid	Reaction with acetic acid
Zinc	Strong reaction	Weaker reaction
Magnesium	Strong reaction	Weaker reaction
Iron	Weak reaction	Very weak or no reaction
Copper	No reaction	No reaction

Part D

Household substance	Red litmus paper	Blue litmus paper	pH paper	Bromothymol Blue	Phenolphthlein	Universal indicator
Lemon juice			Red			
Apple			Red-orange			
Vinegar			Red-orange			
Carbonated drink			Orange			
Tomato juice			Orange			
Black coffee			Orange			
Milk			Yellow-green			
Water			Green			
Salt water			Green-blue			
Baking soda			Green-blue			
Laundry detergent			Turquoise			
Milk of magnesia			Blue			
Household ammonia			Blue			
Bleach			Blue			
Drain cleaner			Dark blue			

Note: pH paper colours may differ from the chart and should be tested ahead of time, as different brands can be more acidic or basic.

Data Analysis

Tests	Acid properties	Base properties
Red litmus paper	No change	Turns blue
Blue litmus paper	Turns red	No change
pH paper	Red – Orange	Green – dark blue
Bromothymol Blue	Turns yellow	No change
Phenolphthalein	No change	Turns pink
Universal indicator	Red-orange	Blue-purple
Reaction with a metal	Reaction	No reaction

Household sample	pH value	Acid or base?
Lemon juice	2	
Apple	3	
Vinegar	2.5 – 3.5	
Carbonated drink	4	
Tomato juice	4	
Black coffee	5	
Milk	6.5	
Water	7	
Salt water	8	
Baking soda	8	
Laundry detergent	9	
Milk of magnesia	10	
Household ammonia	11–12	
Bleach	12.5	
Drain cleaner	13 – 14	

Note: pH paper colours may differ from the chart and should be tested ahead of time, as different brands can be more acidic or basic.

Answers to Questions

1. Either type of litmus paper can be used. Red litmus paper will not change. Blue litmus paper will turn red.
2. Indicators usually give a range of pH values, not a precise pH value.
3. When acids were placed on most metals, a reaction could be observed. The metal disappeared and a gas and heat were produced.
4. Not all metals reacted similarly. The iron reacted less than the magnesium or zinc, and the copper did not react at all.
5. Acids: pH value: less than 7. Stronger acids have lower pH values, react with metals, turn litmus paper red, turn bromothymol blue indicator to yellow, turn universal indicator solution from red to green, and do not react with a phenolphthalein solution. Bases: pH value: greater than 7, stronger bases have higher pH values, do not react with metals, turn litmus paper blue, do not react with bromothymol blue, turn universal indicator solution from green to purple, and turn phenolphthalein solution pink.
6. See data analysis table. Students should arrive at the following conclusions:
    1. Acids have a pH of less than 7.
    2. Bases have a pH greater than 7.
    3. Acids react with metals, but not all react with the same intensity. The strength of the reaction varies with each acid and each metal.

SLA
Student Learning Activity

Balancing Chemical Equations

1. Translate the following word equations to balanced chemical equations. Identify the type of reaction. Write your responses on a separate sheet.
   a. Sodium metal combines with chlorine gas to produce sodium chloride crystals.
   b. Solid magnesium reacts with hydrogen chloride to produce a magnesium chloride solution and hydrogen gas.
   c. Potassium iodide reacts with calcium sulfide to produce potassium sulfide and calcium iodide.
   d. Silver oxide decomposes to produce silver metal and oxygen gas.
   e. Dicarbon hexahydride (ethane) reacts with oxygen gas to produce carbon dioxide gas and water vapour.

2. Translate the following chemical equations to word equations. Identify the type of reaction.
   a. $Fe_{(s)} + CuS_{(aq)}$ fi $FeS_{(aq)} + Cu_{(s)}$
   b. $4Fe_{(s)} + 3O_{2(g)}$ fi $2Fe_2O_{3(s)}$
   c. $BaF_{2(aq)} + 2LiBr_{(aq)}$ fi $BaBr_{2(aq)} + 2LiF_{(aq)}$
   d. $CH_{4(g)} + 2O_{2(g)}$ fi $CO_{2(g)} + 2H_2O_{(g)}$
   e. $2MgO_{(s)}$ fi $2Mg_{(s)} + O_{2(g)}$

3. Complete and balance the following equations.
   a. $C_8H_{18(l)} + O_{2(g)}$ fi (combustion)
   b. $Al_{(s)} + I_{2(g)}$ fi (synthesis)
   c. $BeF_{2(aq)} + K_2O_{(aq)}$ fi (double displacement)
   d. $Cl_{2(g)} + NaBr_{(aq)}$ fi (single displacement)
   e. $NaCl_{(l)}$ fi (decomposition)

TSM
Teacher
Support
Material

Balancing Chemical Equations *(Teacher Notes)*

1. Translate the following word equations to balanced chemical equations. Identify the type of reaction.
   a. $2Na_{(s)} + Cl_{2(g)}$ fi $2NaCl_{(s)}$ (synthesis)
   b. $Mg_{(s)} + 2HCl_{(aq)}$ fi $MgCl_{2(aq)} + H_{2(g)}$ (single displacement)
   c. $2KI_{(aq)} + CaS_{(s)}$ fi $K_2S + CaI_2$ (double displacement)
   d. $2Ag_2O_{(s)}$ fi $4Ag_{(s)} + O_{2(g)}$ (decomposition)
   e. $2C_2H_{6(g)} + 7O_{2(g)}$ fi $4CO_{2(g)} + 6H_2O_{(g)}$ (combustion)

2. Translate the following chemical equations to word equations. Identify the type of reaction.
   a. Iron metal reacts in a solution of copper (II) sulfide to produce iron sulfide in solution and a solid copper precipitate (single displacement).
   b. Iron metal reacts with oxygen gas in the air to produce iron (III) oxide as rust (synthesis).
   c. Solutions of barium fluoride and lithium bromide combine to produce a new solution containing barium bromide and lithium fluoride (double displacement).
   d. Carbon tetrahydride (methane gas) burns in the presence of oxygen gas to produce carbon dioxide gas and water vapour (combustion).
   e. Solid magnesium oxide decomposes into pure magnesium metal and gives off oxygen gas (decomposition).

3. Complete and balance the following equations.
   a. $2C_8H_{18(l)} + 25O_{2(g)}$ fi $16CO_{2(g)} + 18H_2O_{(g)}$
   b. $6Al_{(s)} + 9I_{2(g)}$ fi $6AlI_{3(s)}$
   c. $BeF_{2(aq)} + K_2O_{(aq)}$ fi $BeO_{(s)} + 2KF_{(aq)}$
   d. $Cl_{2(g)} + 2NaBr_{(aq)}$ fi $2NaCl_{(aq)} + Br_{2(g)}$
   e. $2NaCl_{(l)}$ fi $Na_{2(g)} + Cl_{2(g)}$

SLA
Student Learning Activity

Experiment: Law of Conservation of Mass

Purpose

To determine experimentally whether mass is conserved during a chemical reaction.

Equipment/Materials

- Balance
- Erlenmeyer flask (125 mL)
- Antacid tablet
- Balloon
- Warm water (25 mL)

Procedure

Part A: Open Reaction

1. Before the reaction, find the combined mass of the Erlenmeyer flask with 25 mL of water and one whole antacid tablet. Record results.
2. Drop the antacid tablet into the water, and agitate gently until the reaction has stopped.
3. After the reaction, find the combined mass of the Erlenmeyer flask with water and antacid tablet. Record results.

Part B: Closed Reaction

1. Before the reaction, find the combined mass of the Erlenmeyer flask with 25 mL of water, one whole antacid tablet, and the balloon. Record results.
2. Drop the antacid tablet into the water and quickly place the balloon over the mouth of the Erlenmeyer Flask. Agitate gently until the balloon is inflated and the reaction has stopped.
3. After the reaction has stopped, find the combined mass of the Erlenmeyer flask with water, the antacid tablet, and balloon still attached. Record results.

Data Collection

	Mass before reaction	Mass after reaction
Open system		
Closed system		

Data Analysis

Interpret your data collection results and draw conclusions about whether mass is conserved during all chemical reactions. Construct your response in the space below.

Questions

1. In part A, how did the mass of the reactants and glassware before the reaction compare to the mass of the reactants and glassware after the reaction? Explain.

2. In part B, how did the mass of the reactants and glassware before the reaction compare to the mass of the reactants and glassware after the reaction? Explain.

3. Did your results support the Law of Conservation of Mass? Explain. If not, what sources of error might have affected your results?

4. A chemical reaction occurs when one or more substances react to produce one or more new substances. What happened to the reactants? Did they disappear? Were new atoms created when products were formed?

5. Can matter be created? Destroyed? Changed?

6. Do gases have a mass? If so, how can their mass be measured?

TSM
Teacher Support Material

Experiment: Law of Conservation of Mass
(Teacher Notes)

1. Caution should be exercised, as it is possible that enough gas will be produced from the reaction to propel the balloon off the mouth of the flask.
2. Avoid using hot water, as it will cause the reaction to happen too fast, and students may not have enough time to properly place the balloon on the mouth of the Erlenmeyer flask.
3. The amount of water or tablet size can be left up to the discretion of the teacher. Perform the lab ahead of time to ensure expected results.
4. Often student measurements may not be accurate or equipment may not be calibrated precisely. This can lead to incorrect assumptions about conservation of mass. Explain possible sources of error to students to avoid erroneous conclusions.

Data Collection

Answers may vary, depending upon the amount of water used, amount of antacid tablet, size of balloon, and size of flask. However, in the open system, the mass before the reaction should be higher than the mass after the reaction. In the closed system, both masses should be the same.

Data Analysis

Students should conclude that in the open system, masses differed due to lost gas. In the closed system, masses were the same, because the gas could not escape. However, students should recognize that in any chemical reaction, mass of reactants is always equal to mass of products and therefore conserved.

Answers to Questions

1. The mass before the reaction was greater than the mass after the reaction. Because the Erlenmeyer flask was open, a gas escaped.
2. The mass before the reaction was equal to the mass after the reaction. The balloon that was placed on the mouth of the Erlenmeyer prevented the gas from escaping.
3. Answers may vary. Possible sources of error could include an improperly calibrated balance, incorrect balance readings, poor seal with the balloon on the flask, too much time taken to place the balloon on the mouth of the flask, balloon propelled off the mouth of the flask.
4. The reactants were transformed into the products. They did not disappear, but atoms that made up the reactants were rearranged to form the new substances.
5. The Law of Conservation of Mass states that matter cannot be created or destroyed, only changed to form new substances with the original atoms.
6. Gases do have a mass, but their mass is very small and difficult to measure. Possible ways to answer this question include measuring the mass of an empty container (balloon), re-massing the container with the gas, and then subtracting the results. Using a gas collection tube is another method of obtaining the mass of gas.

SLA
Student Learning Activity

Experiment: Reaction Types

Purpose

1. To investigate and identify five types of chemical reactions
2. Balance chemical reactions

Equipment/Materials

Bunsen burner	magnesium ribbon
crucible tongs	hydrochloric acid (6M)
test tubes	mossy zinc
test-tube holder	copper wire
test-tube rack	silver nitrate (0.5M)
wooden splints	cobalt chloride paper
safety goggles	limewater solution
candle	potassium iodide (0.1M)
Erlenmeyer flasks	lead (II) nitrate (0.1M)
500 mL beaker	sodium hydroxide (0.1M)
evaporating dish	copper (II) sulfate (0.1M)/copper (II) chloride (0.1M)

SAFETY

In this investigation you will be working with open flames, handling acids, heating chemicals, and producing gaseous products.

Wear safety goggles and review safety procedures before performing any experiments.

Do not look directly at the magnesium ribbon flame! Hold the burning magnesium away from you and over an evaporating dish.

Handle hydrochloric acid with care as it can cause painful burns. Do not inhale any of the HCl fumes.

Procedure

Part A: Synthesis

1. Before heating the magnesium ribbon, examine its properties.
2. Using crucible tongs, hold the magnesium ribbon in the blue flame of a Bunsen burner for one to two minutes, or until the magnesium starts to burn.
3. When the ribbon stops burning, put the residue in an evaporating dish.
4. Examine the magnesium and note any changes in its appearance caused by heating.

Part B: Single Displacement

1. Place about 5 mL of 6M HCl in a test tube. Add a small piece of mossy zinc to the solution. Observe and record what happens.
2. Using a test-tube holder, invert a second test tube over the mouth of the first test tube where the reaction is occurring. Remove the inverted test tube after about 30 seconds and quickly insert a burning wooden splint. Note the appearance of the substance in the reaction test tube.

OR

1. Loosely coil a 10-cm piece of copper wire around a pen or pencil.
2. Place the coiled wire into a test tube that is two-thirds full of 0.5M silver nitrate solution. Allow solution to stand overnight.
3. Observe and record changes to both the copper wire and the silver nitrate solution.

Part C: Combustion

1. Observe the appearance of both a strip of cobalt chloride paper and a limewater solution. Describe and record their appearance.
2. Using beaker tongs, place a large beaker (500 mL) over a lit candle. Hold the beaker over the candle until the candle is extinguished.
3. When the flame goes out, turn the beaker over, examine, and record its contents.
4. Test for any moisture with a strip of cobalt chloride paper. Record any observable changes.
5. Relight the candle and place an Erlenmeyer flask over the candle until the candle goes out.
6. Quickly overturn the flask and add 25 mL of limewater solution. Record any observable changes.

Part D: Double Displacement

1. Add 5 mL of a 0.1M potassium iodide solution to a test tube. Next, add 5 mL of a 0.1M lead (II) nitrate solution. Observe what happens and note any changes in the mixture. Record results.

OR

1. Add 5 mL of a 0.1M sodium hydroxide solution to a test tube. Next, add 5 mL of 0.1M copper sulfate or 5 mL of 0.1M copper chloride solution. Observe what happens and note any changes in the mixture. Record results.

Part E: Decomposition

1. Place approximately 2g of copper (II) carbonate in a clean, dry test tube. Note the appearance of the sample.
2. Using a test-tube holder, heat the copper (II) carbonate strongly for about two minutes.
3. Extinguish the burner flame and insert a burning wooden splint into the test tube.
4. Note: If carbon dioxide gas is present, it will put the flame out.
5. Observe any changes in the appearance of the chemical remaining in the test tube. Record results.

Data Collection

Sample	Before reaction	After reaction
A. Synthesis		
B. Single displacement		
C. Combustion		
D. Double displacement		
E. Decomposition		

Data Analysis

In this section, students should interpret observations and explain reaction types by the products they form. Include chemical equations for each of the reactions. For example, combustion reactions always form carbon dioxide and water as their products.

Questions

1. In this experiment, what method was used to test the presence of carbon dioxide gas? Hydrogen gas? Water?

2. Balance each of the following equations by inserting the proper coefficients where needed.

   a. $Mg_{(s)}$ + $O_{2(g)}$ fi $MgO_{(s)}$

   b. $HCl_{(aq)}$ + $Zn_{(s)}$ fi $H_{2(g)}$ + $ZnCl_{2(s)}$

   c. $C_{25}H_{52(s)}$ + $O_{2(g)}$ fi $CO_{2(g)}$ + $H_2O_{(g)}$

   d. $Cu_{(s)}$ + $AgNO_{3(aq)}$ fi $Ag_{(s)}$ + $Cu(NO_3)_{2(aq)}$

   e. $KI_{(aq)}$ + $Pb(NO_3)_{2(aq)}$ fi $KNO_{3(aq)}$ + $PbI_{2(s)}$

   f. $NaOH_{(aq)}$ + $CuSO_{4(aq)}$ fi $Na_2SO_{4(aq)}$ + $Cu(OH)_{2(s)}$

   g. $CuCO_{3(s)}$ fi $CuO_{(s)}$ + $CO_{2(g)}$

3. Classify each of the following reactions

   a. $2Fe_{(s)}$ + $O_{2(g)}$ fi $2FeO_{(s)}$ ____________________

   b. $BaCl_{2(aq)}$ + $Na_2SO_{4(aq)}$ fi $BaSO_{4(s)}$ + $2NaCl_{(aq)}$ ____________________

   c. $Cl_{2(aq)}$ + $2KBr_{(aq)}$ fi $2KCl_{(aq)}$ + $Br_{2(aq)}$ ____________________

   d. $2Ag_2O_{(s)}$ fi $4Ag_{(s)}$ + $O_{2(g)}$ ____________________

   e. $CH_{4(g)}$ + $2O_{2(g)}$ fi $CO_{2(g)}$ + $2H_2O_{(g)}$ ____________________

TSM
Teacher Support Material

Experiment: Reaction Types
(Teacher Notes)

1. Review appropriate safety precautions.
2. As only binary compounds have been discussed in class, the formulas of any reactants or products that contain polyatomic ions must be provided.

Data Collection

Sample	Before reaction	After reaction
A. Synthesis	Magnesium is shiny, malleable, and ductile.	The product is a white powder, called magnesium oxide.
B. Single displacement	HCl is a clear, colourless, room-temperature solution. Zinc is a dark gray malleable solid.	A combustible gas is produced. The solution is still clear and colourless, but its temperature has increased.
C. Combustion	The candle is an opaque solid composed of paraffin wax.	The candle reacts with oxygen to produce a colourless gas that extinguishes a wooden splint. Residue along the inside of the test tube turns cobalt chloride paper pink.
D. Double displacement	Both reactants (potassium iodide and lead [II] nitrate) are odourless, colourless solutions. The HCl solution is colourless. The copper chloride solution has a pale blue colour.	A yellow precipitate is formed that will settle out of the colourless solution. A bluish gel-like precipitate is formed and settles out of the colourless solution.
E. Decomposition	The copper carbonate is a granular blue powder.	After heating, the powder becomes white. A colourless gas is released, which will extinguish a burning splint.

Data Analysis

Synthesis reactions: element + element fi compound

$Mg_{(s)} + O_{2(g)}$ fi $MgO_{(s)}$

Single displacement: element + compound fi element + compound

$HCl_{(aq)} + Zn_{(s)}$ fi $H_{2(g)} + ZnCl_{2(s)}$

$Cu_{(s)} + AgNO_{3(aq)}$ fi $Ag_{(s)} + Cu(NO_3)_{2(aq)}$

Combustion reactions: hydrocarbon + oxygen $\rightarrow$ carbon dioxide + water vapour

$C_{25}H_{52(s)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

Double displacement: compound + compound $\rightarrow$ compound + compound

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

$NaOH_{(aq)} + CuSO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + Cu(OH)_{2(s)}$

Decomposition reactions: compound $\rightarrow$ compound + compound

$CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2(g)}$

Note: In decomposition reactions, elements or compounds can be produced as products.

Answers to Questions

1. In this experiment, what method was used to test the presence of carbon dioxide gas? Hydrogen gas? Water?

   If carbon dioxide gas is present, it will put the flame of a wooden splint out and turn limewater milky. If hydrogen gas is present, a popping sound will be heard as the hydrogen gas explodes. If water is present, the cobalt chloride paper will turn from blue to pink.

2. Balance each of the following equations by inserting the proper coefficients where needed.

   a. $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

   b. $2HCl_{(aq)} + Zn_{(s)} \rightarrow H_{2(g)} + ZnCl_{2(s)}$

   c. $Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow 2Ag_{(s)} + Cu(NO_3)_{2(aq)}$

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

   e. $2NaOH_{(aq)} + CuSO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + Cu(OH)_{2(s)}$

3. Classify each of the following reactions

   a. $2Fe_{(s)} + O_{2(g)} \rightarrow 2FeO_{(s)}$ *(synthesis)*

   b. $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaCl_{(aq)}$ *(double displacement)*

   c. $Cl_{2(aq)} + 2KBr_{(aq)} \rightarrow 2KCl_{(aq)} + Br_{2(aq)}$ *(single displacement)*

   d. $2Ag_2O_{(s)} \rightarrow 4Ag_{(s)} + O_{2(g)}$ *(decomposition)*

   e. $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$ *(combustion)*

TSM
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Support
Material

Ionic Compounds

Ionic compounds form when electrons are transferred from metal atoms to non-metal atoms so that the atoms obtain the stable electron arrangements of the nearest noble gases.

Rules for naming binary ionic compounds

1. Name the positive ion first by writing the full name of the metallic element.
2. Name the non-metal ion next by dropping the last syllable(s) of the name of the element and adding the suffix “ide.”
   Example: reaction between sodium and chlorine forms fi sodium chloride
   Example: given the chemical formula SrS fi strontium sulfide

Rules for writing chemical formulas for binary ionic compounds

1. Using the name of the binary compound, write the symbols and charge for the ions involved. Write the ion charge as a superscript.
   Example: aluminum oxide — Al^{3+} O^{2-}

2. Determine subscripts that will produce a balance of charge.
   Example: Al_2^{3+} O_3^{2-}

3. As a check to ensure the formula is written correctly, multiply the charge for each ion by the subscript for the same ion. The total positive charge should equal the total negative charge and the net charge per ionic formula should be zero.
   Example: $Al^{3+} \times 2 = 6^+$
   $O^{2-} \times 3 = 6^-$
   $6^+ + 6^- = 0$ (zero net charge)

4. Write the final chemical formula without the charges.
   Final Answer: Al_2O_3

Note:

1. The subscript “1” is not used.
   Example: lithium bromide — Li^{1+} Br^{1-}
   Li_1^{1+} Br_1^{1-}
   Final Answer: LiBr

2. Reduce and simplify the ratio.
   Example: magnesium sulfide — Mg^{2+} S^{2-}
   Mg_2^{2+} S_2^{2-}
   Mg_2S_2
   Final Answer: MgS

Stock System

The Stock System is used only if the metal element in the compound may have more than one charge. Example: Iron can form ions that have a charge of 2+ or 3+.

In this system, the valence of the metal element is indicated by using a Roman numeral in parenthesis following the name for the metal.

Example: Iron (II) for a valence of 2+ and Iron (III) for a valence of 3+.

Determining the Stock System name from a chemical formula may be more difficult for students. Knowing that the total charge on a compound is equal to zero, and using the known non-metal charge, students can determine the charge on the metal.

Example: Fe_2O_3

1. The non-metal ion, oxygen, has a charge of O^{2-}. As there are three oxide ions in the formula, the total negative charge in the compound is $2^- \times 3 = 6^-$.

2. The positive ions must have a charge equal to the charge of negative ions to give the compound a net charge of zero. The charge on the ion must be 6+ (? x 2 = 6+).

3. Since there are two iron ions shown, the valence on the lead is 3+. The name of the compound is lead (III) oxide.

Example: PbS_2

Total negative ion charge is $2^- \times 2 = 4^-$.

Therefore, total positive ions must equal 4^+ ($? \times 1 = 4^+$).

Since there is only one lead ion, the valence on the lead is 4.

The name of the compound is lead (IV) sulfide.

TSM
Teacher Support Material

Molecular Compounds

Binary molecular compounds contain atoms of two non-metal elements, covalently bonded by sharing electrons.

Following IUPAC guidelines, molecular compounds are named using a prefix system. A Greek prefix is used to indicate the number of each type of atom in the molecule. The prefixes are:
mono = 1
di = 2
tri = 3
tetra = 4
penta = 5
hexa = 6

Rules for naming binary molecular compounds

1. The first non-metal name is written in full.
2. The second non-metal element is named with the suffix "ide" ending.
3. Assign a prefix to each element expressing the number of atoms present in the molecule. The prefix mono is used only with the second element in the compound. When the second element is oxygen, the vowel "o" in mono is dropped and the name becomes monoxide rather than monooxide. Similarly, the "a" in tetra, penta, or hexa is also dropped.

   Example: CO_2 is written as carbon dioxide

   N_2O_5 is written as dinitrogen pentaoxide

Rules for writing binary molecular compounds

1. Both non-metal symbols are followed by a subscript indicating the number of atoms present. The number 1 is not written, as it is understood to be present.

   Example: carbon tetrachloridide = CCl_4

   silicon hexafluoride = SiF_6

2. When determining which non-metal element to place first in the compound, the general rule is to read across the periodic table from left to right. The element that appears first is usually written first. There are exceptions to this general rule.

   Example: H_2S = hydrogen sulfide

   NF_3 = nitrogen trifluoride

SLA
Student Learning Activity

Experiment: Reaction Types

Purpose

1. To investigate and identify five types of chemical reactions.
2. Balance chemical reactions.

Materials

- test tube (25 mL)
- magnesium ribbon (~5–10 cm)
- thin copper wire (~10 cm piece)
- cobalt chloride paper
- limewater solution
- Erlenmeyer flask (250 mL)
- 3.00 M HCl
- Bunsen burner

Solutions

- silver nitrate
- potassium iodide
- lead (II) nitrate
- sodium hydroxide
- copper (II) sulfate
- copper (II) chloride

Procedure

A. Burn magnesium in the blue flame of a Bunsen burner.

B. Place about 3 mL of HCl in a test tube. Add a small piece of mossy zinc to the solution. Test for any gas with a burning splint.

C. Loosely coil a 10-cm piece of copper wire around a pen or pencil. Place the coiled wire into a test tube that is two thirds full of .5M silver nitrate solution.

D. 1. Observe both a strip of cobalt chloride paper and 25 mL of limewater solution. Describe and record their appearance in your lab book or log.
   2. Light a candle. Place a large beaker (500 mL) over the candle. When the flame goes out, turn the beaker over, examine, and record its contents.
   3. Test for any moisture with a strip of cobalt chloride paper. Record any observable changes.
   4. Relight the candle and place an Erlenmeyer flask over the candle until the candle goes out.
   5. Quickly overturn the flask and add the 25 mL of limewater solution. Record any observable changes.

E. 1. Add 5 mL of a .1M potassium iodide solution to 5 mL of a .1M lead (II) nitrate solution. Observe and record results.
   2. Add 5 mL of a .1M sodium hydroxide solution to 5 mL of .1M copper sulfate or 5 mL of .1M copper chloride solution. Observe and record results.