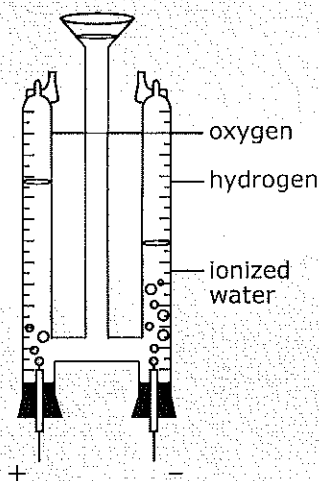


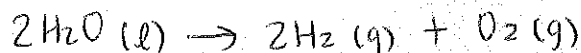
4.1 Writing and Balancing Chemical Equations — The Magic of Chemistry

Warm Up (p. 162)

Water may be broken down into its component elements, oxygen and hydrogen, by the application of electric current. This process is called electrolysis.



1. Use your knowledge of chemical nomenclature to write an equation showing the process occurring in the diagram. (Be sure to indicate hydrogen and oxygen as diatomic molecules.)



2. Count the total number of atoms of hydrogen and oxygen before and after the change. Make any necessary changes to ensure that these are equal.
3. Make any necessary changes to indicate the states of the substances in your equation.
4. Include anything else you feel is necessary to help your equation fully describe what is happening during this process.

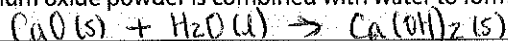
Quick Check (p. 163)

1. List three pieces of evidence of chemical change in the decomposition of mercury(II) oxide.

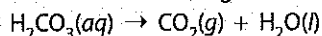
color change (red → silver), evolution of gas (O_2) and ↑ solid mass.
 $\text{CaO}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{OH})_2(\text{s})$

2. Convert the following from a word equation to a formula equation:

Calcium oxide powder is combined with water to form calcium hydroxide solid.



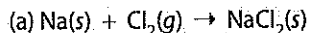
3. Convert the following from a formula equation to a word equation:



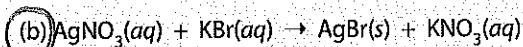
carbonic acid decomposes to carbon dioxide gas and water

Quick Check (p. 165)

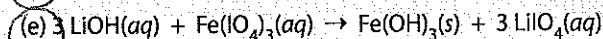
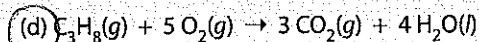
Circle the letter of the properly balanced equations:



(incorrect formula created)



(missing coefficient)

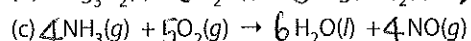
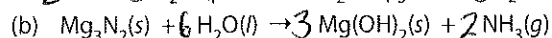
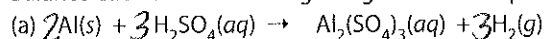


Practice Problems — Balancing Formula Equations (p. 168)

1. Complete the following table to check the balancing of the sample equation above.

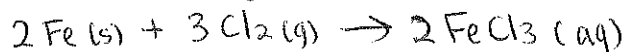
Atom	Reactant Side	Product Side
Na	8	8
O	20	20
H	8	8
Cr	4	4

2. Balance each of the following using the smallest possible whole number coefficients.

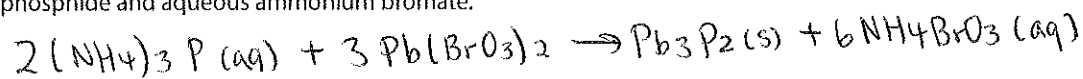
**Practice Problems — Writing Balanced Formula Equations from Word Equations** (p. 169)

Write balanced equations for each of the following, using the smallest possible whole number coefficients. Include appropriate phase indicators.

1. Solid iron and chlorine gas combine to form iron(III) chloride.



2. Aqueous solutions of ammonium phosphide and lead(II) bromate react to give a precipitate of lead(II) phosphide and aqueous ammonium bromate.



3. Calcium metal reacts in an aqueous solution of nickel(III) sulphate to form calcium sulfate solution and nickel metal.



4.1 Activity: Building a Balanced Equation (p. 170)

Question

How can you represent the chemical reaction between nitrogen dioxide gas and water to form nitric acid and nitrogen monoxide?

Materials

You may use any of the following materials:

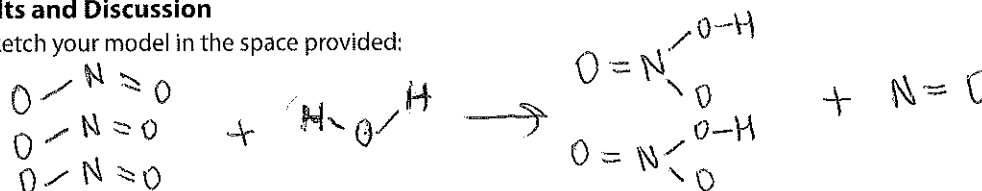
- a series of coins of different sizes representing nitrogen, oxygen, and hydrogen atoms
- modelling clay in different colours and sizes representing the required atoms
- a chemical model kit
- a series of drawings using different colours and sizes of images to represent the required atoms

Procedure

1. Use the materials you've selected to build a model of the balanced equation that represents the reaction in the question.
2. Compare your model with models done by one or more of your classmates.

Results and Discussion

1. Sketch your model in the space provided:



2. Make a table to check the balancing in your model equation:



3. How many chemical bonds were broken during your reaction? Describe each of these bonds.

3 O-N single bonds broken
2 H-O single bonds "
3 N=O double bonds "
} 8 bonds broken, endothermic

4. How many chemical bonds were formed during your reaction? Describe each.

2 N=O bonds, 4 N-O bonds, 1 N=O bond, 2 O-H bonds $\Rightarrow$ 9 bonds formed, exothermic

5. Were more bonds formed or broken? What might this mean for your reaction?

More bonds formed, exothermic rxn.

(Note: Circles with single bonds everywhere is certainly acceptable at this point)

4.1 Review Questions (p. 171)

1. State two examples of chemical change from everyday life. What evidence indicates that these are chemical changes?

For Example: Rusting — colour change
Burning wood — exothermic (heat released)

2. What chemical law requires us to balance chemical equations?

Law of Conservation of Mass

Who was responsible for formulating this law?

Antoine Lavoisier

3. Write a word equation for each formula equation below:

(a) $C(s) + O_2(g) \rightarrow CO_2(g)$ Solid carbon combines with oxygen gas to form carbon dioxide gas

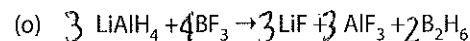
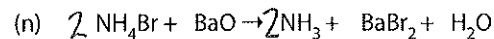
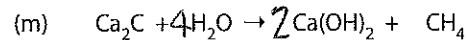
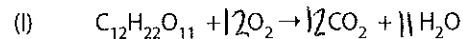
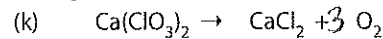
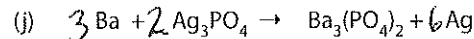
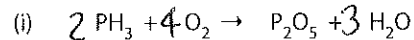
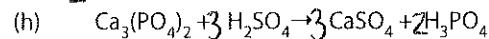
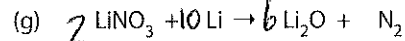
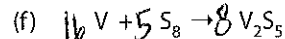
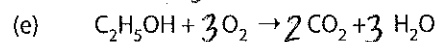
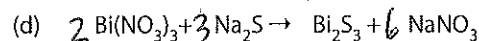
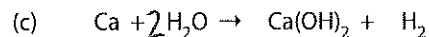
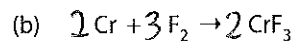
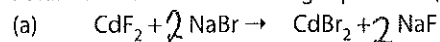
(b) $CH_4(g) + 2 O_2(g) \rightarrow CO_2(g) + 2 H_2O(g)$ Methane gas combusts in oxygen gas to produce carbon dioxide gas and water vapor

(c) $Cl_2(g) + 2 KI(s) \rightarrow I_2(s) + 2 KCl(s)$ Chlorine gas reacts with solid potassium iodide to form solid iodine and potassium chloride solid.

(d) $HCl(aq) + NaOH(s) \rightarrow NaCl(aq) + H_2O(l)$ hydrochloric acid neutralizes solid sodium hydroxide to form sodium chloride solution and water

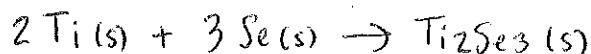
(e) $KF(s) \rightarrow K(s) + F_2(g)$ Solid potassium fluoride decomposes to form solid potassium and fluorine gas

4. Balance each of the following equations. (State indicators are not required.)



5. Write a balanced formula equation for each of the following (phase indicators should be included if possible):

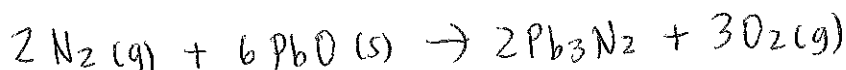
(a) Titanium metal reacts with selenium to produce crystals of titanium(III) selenide.



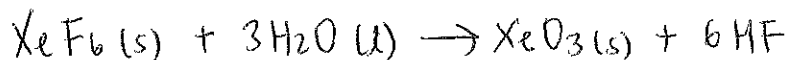
(b) Phosphoric acid is neutralized with barium hydroxide to produce a precipitate of barium phosphate in water.



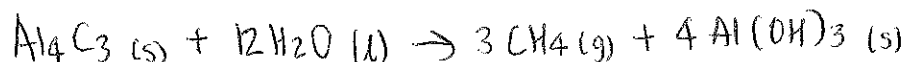
(c) Nitrogen gas reacts with lead(II) oxide powder to yield lead(II) nitride and oxygen gas.



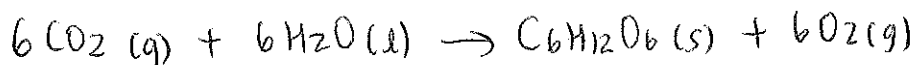
(d) Xenon hexafluoride crystals react with water to produce xenon trioxide powder and hydrofluoric acid.



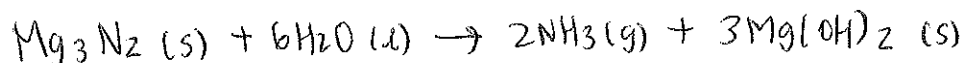
(e) Aluminum carbide is reacted with water in the synthesis of methane gas. Aluminum hydroxide precipitate is also formed.



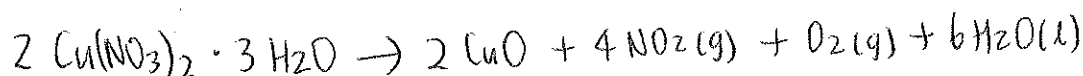
(f) Plants produce the simple sugar $\text{C}_6\text{H}_{12}\text{O}_6$ and oxygen gas from carbon dioxide and water during photosynthesis.



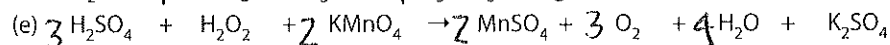
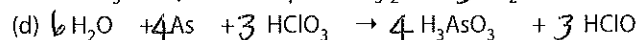
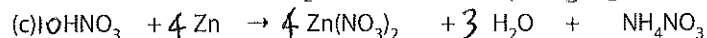
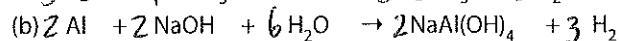
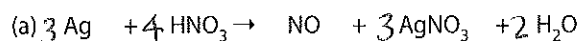
(g) Ammonia gas (NH_3) is formed along with a precipitate of magnesium hydroxide from the reaction of magnesium nitride powder with water.



(h) Strong heating of copper(II) nitrate trihydrate produces copper(II) oxide, nitrogen dioxide, oxygen gas, and water.



6. Balancing Bonkers: Some equations are extremely difficult to balance even with the application of the balancing hints! Later on in your chemistry career, you will learn to balance some equations using the concepts of oxidation and reduction. This will be a major stress reducer when it comes to balancing. In the meantime, try to balance a few of these exhauster equations:



Hint: You may want to come back to these once you've finished section 4.3.

4.2 Classifying Chemical Changes and Predicting Products

Warm Up (p. 173)

Reaction Type	Reactants	Products
Synthesis (combination)	two substances	one substance
Decomposition	one substance	two substances
Single replacement	element + compound	new element + compound
Double replacement	two compounds	two new compounds
Neutralization	acid + base	salt + water
Combustion	organic compound + oxygen	carbon dioxide + water

Balance the following equations. Then use the table above to classify each as one of the major reaction types listed:

- $2 \text{Na(s)} + 2 \text{H}_2\text{O(l)} \rightarrow 2 \text{NaOH(aq)} + \text{H}_2\text{(g)}$ Single Replacement
- $\text{Li}_2\text{O(s)} + \text{H}_2\text{O(l)} \rightarrow 2 \text{LiOH(aq)}$ combination (synthesis)
- $2 \text{C}_6\text{H}_{14}\text{(l)} + 19 \text{O}_2\text{(g)} \rightarrow 12 \text{CO}_2\text{(g)} + 14 \text{H}_2\text{O(g)}$ Combustion
- $2 \text{HCl(aq)} + \text{Sr(OH)}_2\text{(aq)} \rightarrow \text{SrCl}_2\text{(aq)} + 2 \text{H}_2\text{O(l)}$ Neutralization
- $2 \text{AlBr}_3\text{(s)} \rightarrow 2 \text{Al(s)} + 3 \text{Br}_2\text{(l)}$ Decomposition

Quick Check (p. 176)

Balance each of the following reactions and classify them as synthesis, decomposition or combustion.

- $\text{P}_2\text{O}_5\text{(g)} + 3 \text{H}_2\text{O(l)} \rightarrow 2 \text{H}_3\text{PO}_4\text{(aq)}$ Combination (synthesis)
- $\text{CH}_4\text{(g)} + 2 \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + 2 \text{H}_2\text{O(g)}$ Combustion
- $\text{H}_2\text{CO}_3\text{(aq)} \rightarrow \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$ (bal) Decomposition

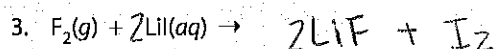
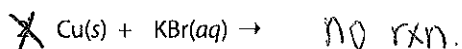
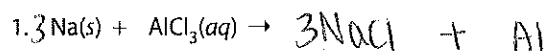
Practice Problems — Predicting Products of Reactions (p. 178)

First classify the equations, then determine the products and write a balanced formula equation for each of the following:

- $2 \text{Fe}_2\text{O}_3\text{(s)} \rightarrow 4 \text{Fe} + 3 \text{O}_2$ Class: Decomposition
- $2 \text{C}_3\text{H}_7\text{OH(l)} + 9 \text{O}_2\text{(g)} \rightarrow 6 \text{CO}_2 + 8 \text{H}_2\text{O}$ Class: Combustion
- $\text{Ag}_2\text{O(s)} + \text{H}_2\text{O(l)} \rightarrow 2 \text{AgOH}$ Class: Combination (syn)
- $\text{Pb(OH)}_4\text{(s)} \rightarrow \text{PbO}_2 + \text{H}_2\text{O (bal)}$ Class: Decomposition

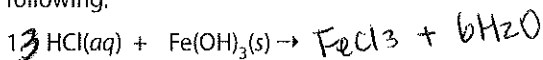
Quick Check (p. 179)

Use the series of chemical reactivity in Table 4.2.2 to determine whether each of the following single replacement reactions would proceed. Predict the products for those that do proceed and balance the equations.

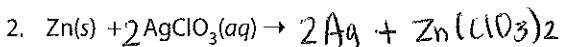


Practice Problems — Predicting Products of Replacement Reactions (p. 182)

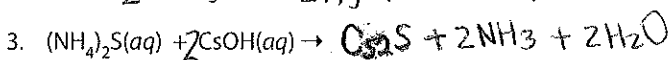
First classify the equations, then determine the products and write a balanced formula equation for each of the following:



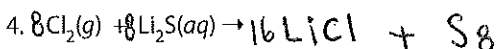
Class: Neutralization



Class: Single Replacement



Class: Double Replacement



Class: Single Replacement

4.2 Activity: Identifying an Unknown Substance (p. 182–183)

Results and Discussion

1. Use the table below to record all information and answer the questions below.

Identity of Unknown	Balanced Equation	Explanation
1) Cl_2	$\text{Cl}_2(aq) + 2\text{NaBr}(aq) \rightarrow 2\text{NaCl}(aq) + \text{Br}_2(aq)$	Chlorine replaces bromine, but iodine does NOT. (Bromine is orange in elemental form.)
2) MgCO_3	$2\text{HNO}_3(aq) + \text{MgCO}_3(s) \rightarrow \text{Mg}(\text{NO}_3)_2(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)$	bubbles indicate carbonic acid formed and decomposed $\rightarrow \text{CO}_2(g)$
3) $\text{Pb}(\text{NO}_3)_2$	$\text{Pb}(\text{NO}_3)_2(aq) + 2\text{NaI}(aq) \rightarrow \text{PbI}_2(s) + 2\text{NaNO}_3(aq)$	PbI_2 precipitates. The other possible iodide (NaI) is soluble. (all IA containing cmpds are)
4) ZnCl_2	$2\text{Al}(s) + 3\text{ZnCl}_2(aq) \rightarrow 3\text{Zn}(s) + 2\text{AlCl}_3(aq)$	Al replaces Zn^{2+} in compounds. It will not replace Ca^{2+} . (see activity series)
5)		

2. Write a balanced chemical equation, including phase indicators for each of the confirming tests performed.

See Above

3. Explain why you decided on the unknown in each equation.

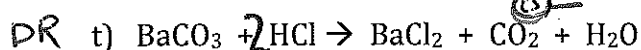
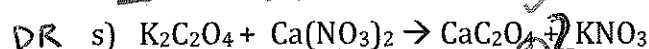
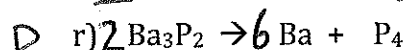
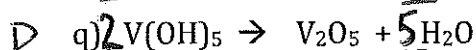
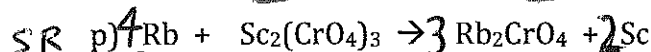
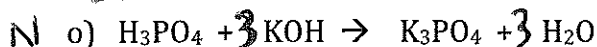
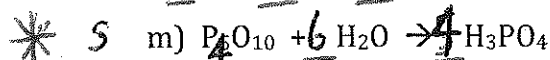
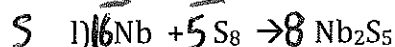
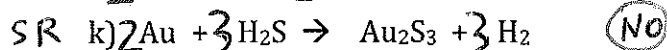
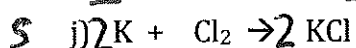
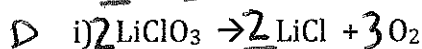
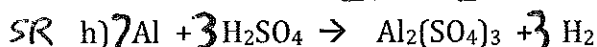
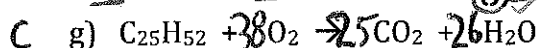
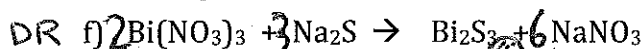
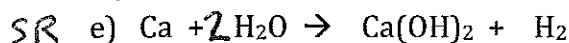
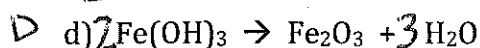
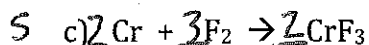
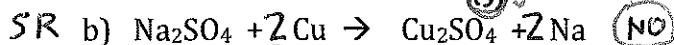
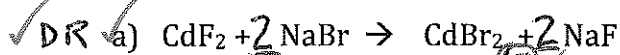
See Above

4. Design a test to distinguish between a fifth pair of unknown substances. Give the substances, the test and the expected observations in the provided spaces in the table above. Complete row #5 for your identified substance in the second table above.

Many possible.

4.2 Review Questions (p. 184)

2 each 1. Balance each of the following reactions:



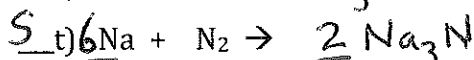
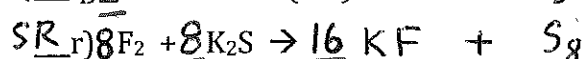
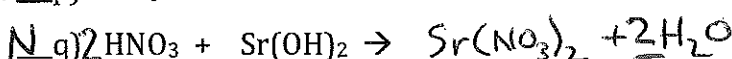
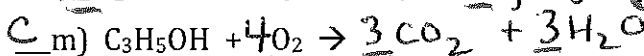
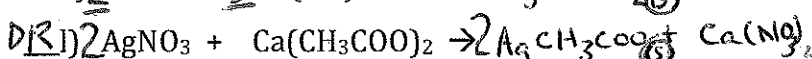
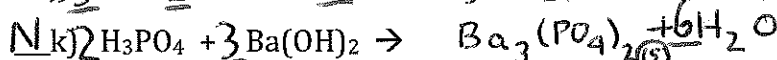
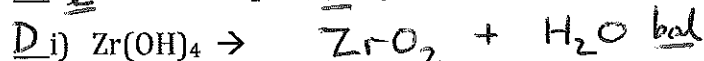
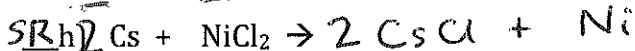
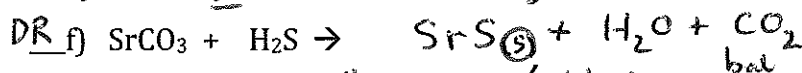
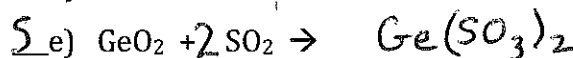
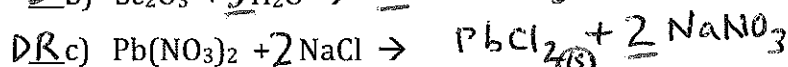
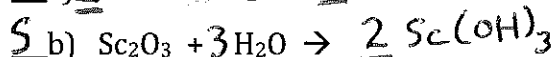
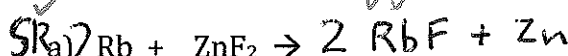
2. a) Classify each of the reactions in question 1 as synthesis, decomposition, combustion, single replacement, double replacement or neutralization.

b) Which of the single replacement reactions would *not* proceed spontaneously? *b, k*

c) Which of the double replacement reactions involve precipitate formation? *a, f, s*

d) Indicate the precipitates with a (s).

3. Classify each of the following reactions, using the following key: S = Synthesis, D = Decomposition, C = Combustion, SR = Single Replacement, DR = Double Replacement, N = Neutralization. Complete the equations and balance them. Indicate any precipitates that form with a (s).

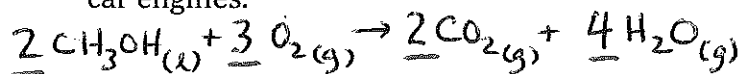
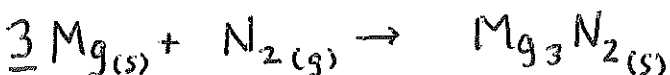


4. Classify each of the following chemical changes using the key from question 3. Write

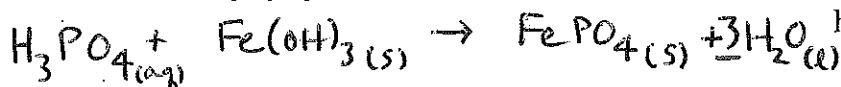


balanced formula equations for each, including ^{DR} Potassium chromate solution indicates the endpoint in a potato chip analysis with a standard silver nitrate solution.

Sa) A piece of magnesium ribbon on a stock shelf reacts with nitrogen gas in the air to form a black coating over time. Cj) Methanol (CH₃OH) is combusted in race car engines.

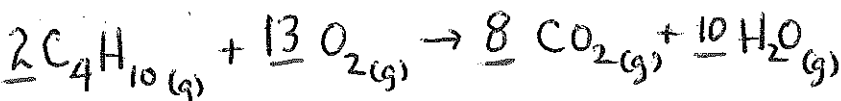


Nb) Phosphoric acid solution removes iron(III) hydroxide stains from an old bath tub. DRk) Baking soda (sodium hydrogen carbonate) ~~may~~ is used to neutralize a spill of hydrochloric acid.

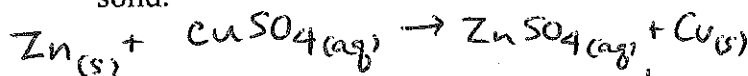
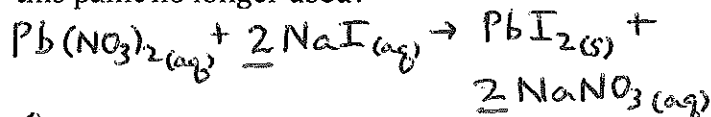


Cc) Butane gas is combusted in a disposable lighter.

DRl) A bright yellow pigment once used in paints is formed from the reaction of lead(II) nitrate and sodium iodide solutions. Why is this paint no longer used?

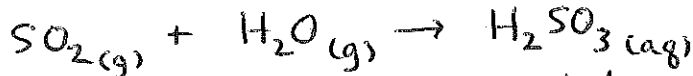
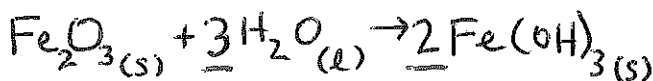


SRd) A zinc strip placed in a solution of copper (II) sulphate becomes coated with brownish solid.



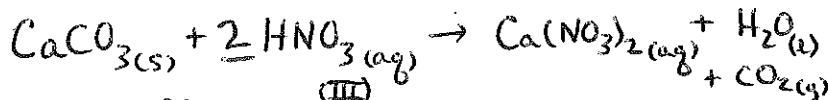
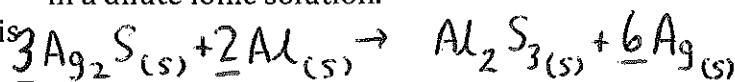
Sm) Iron(III) oxide and water combine to form a basic compound often called rust.

Se) Sulphur dioxide emitted from industrial plants combines with water vapour to form acid rain.



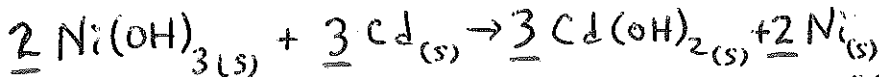
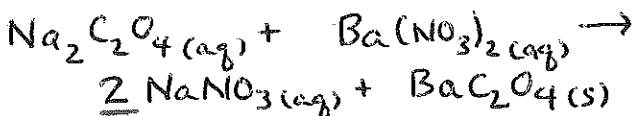
SRn) Dark silver sulphide tarnish may be removed from knives and forks by placing them in contact with a piece of aluminum foil in a dilute ionic solution.

DRf) Calcium carbonate in marble structures is eroded over time by nitric acid in acid rain.



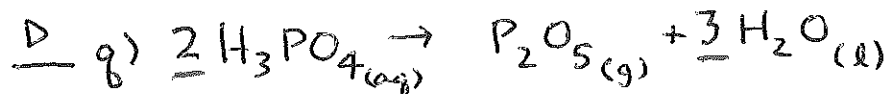
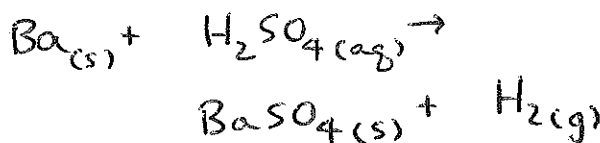
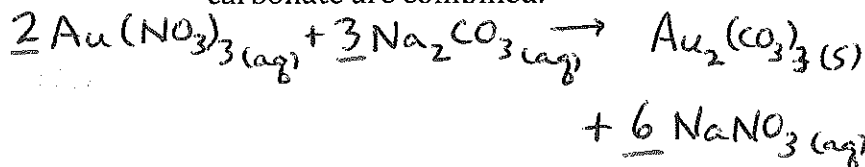
DRo) A precipitate of barium oxalate forms in a solution of sodium nitrate following the combination of two solutions.

SRg) Nickel(III) hydroxide reacts with a cadmium anode in a prototype rechargeable battery.



SRp) A precipitate of barium sulphate and hydrogen gas are formed from the combination of a metal and an acid.

DRh) Solutions of gold(III) nitrate and sodium carbonate are combined.

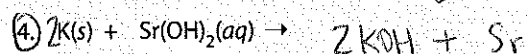
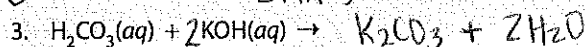
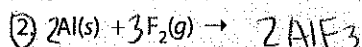
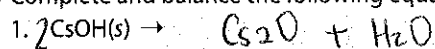


bal

4.3 Another Way to Classify — Identifying Electron Transfer (Extension)

Warm Up (p. 187)

Complete and balance the following equations.



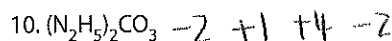
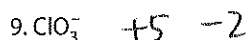
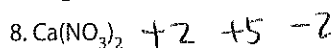
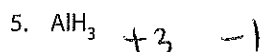
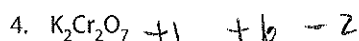
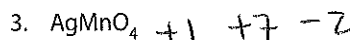
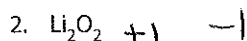
Circle the numbers of the equations that involve a change in the oxidation number (charge) of one or more atoms during the reaction.

Quick Check (p. 187)

1. Decomposition
2. Synthesis (combination)
3. Neutralization
4. Single replacement

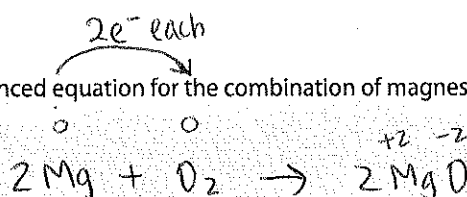
Practice Problems — Assigning Oxidation States (p. 190)

Assign the oxidation state of each atom in the following species:



(p. 190) Quick Check

1. Write a balanced equation for the combination of magnesium metal with oxygen gas to synthesize a metal oxide.



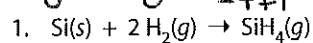
2. Indicate the oxidation state of each element over top of its symbol.
3. Use arrows to indicate the direction of electron transfer between the reactants.

Quick Check (p. 191)

1. (a) How many electrons did each magnesium atom lose in the previous Quick Check? 2
 (b) How many electrons did each oxygen atom gain? 2
 (c) Which species was oxidized? Mg
 Which was reduced? O₂
2. (a) In Figure 4.3.1, how many electrons did each manganese(IV) ion gain? 2
 (b) How many electrons did each chloride ion lose? 1
 Did all the chloride ions lose electrons? no (2 of the Cl⁻ ions are spectators)

Practice Problems — Recognizing Oxidizing and Reducing Agents (p. 193)

Determine the oxidizing and reducing agent in each of the following reactions. Begin by clearly indicating the oxidation states of all elements in each equation.

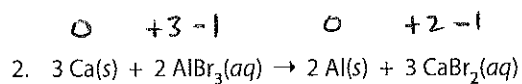


O.A.:

RA.:

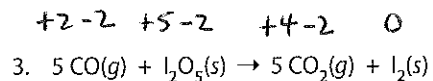
Si

H₂



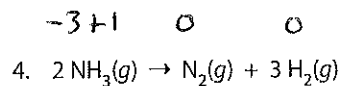
Al³⁺

Ca



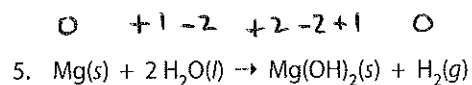
I₂O₅

CO



NH₃

NH₃



H₂O

Mg

4.3 Activity: Building a Table of Reduction Strengths (p. 194)

Question

Can you build a table of reduction strengths?

Background

It is possible to separate out the reduction and oxidation portions of a redox reaction and to represent them as half reactions. For example, the sample reaction from the previous page: $\text{Zn(s)} + 2\text{HCl(aq)} \rightarrow \text{ZnCl}_2\text{(aq)} + \text{H}_2\text{(g)}$ may be broken into the following half reactions: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ (oxidation) and $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ (reduction). Notice that the chloride ion is not included in either reaction, as its oxidation state does not change.

An attempt to replace zinc with hydrogen in the following single replacement reaction: $\text{H}_2\text{(g)} + \text{ZnCl}_2\text{(aq)} \rightarrow 2\text{HCl(aq)} + \text{Zn(s)}$ is *unsuccessful* because H_2 gas is too weak as a reducing agent to be oxidized to H^+ . Similarly Zn^{2+} is too weak as an oxidizing agent to be reduced to Zn. Comparing these two competing reactions shows us that H^+ ion is a stronger oxidizing agent than Zn^{2+} . Consider the following data:

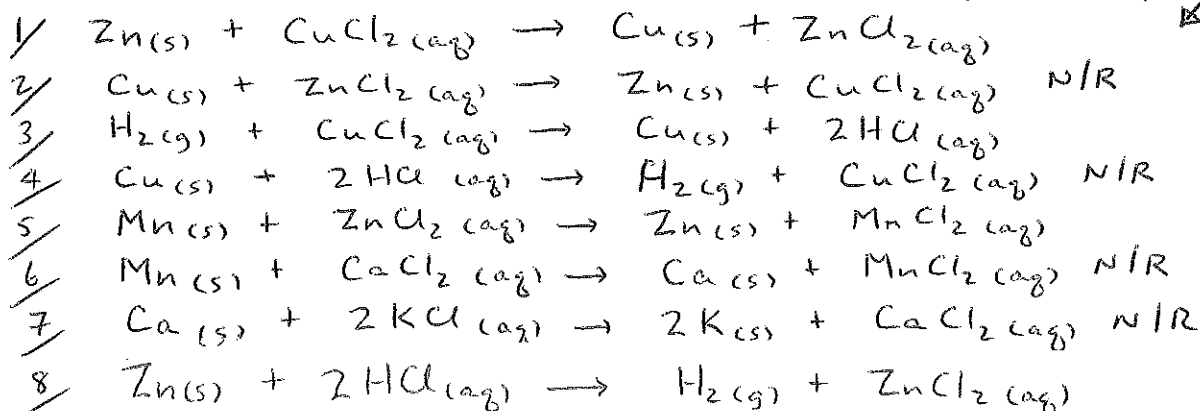
Attempted Reaction	Result
1. A piece of zinc metal is placed in a solution of CuCl_2 .	Brown coating results on the zinc.
2. A piece of copper metal is placed in a ZnCl_2 solution.	No reaction occurs.
3. Hydrogen gas is bubbled through a solution of CuCl_2 .	Bits of brown metal form in an acid.
4. Copper metal is placed in a solution of HCl.	No reaction occurs.
5. Manganese metal is placed in a solution of ZnCl_2 .	Black solid forms in pink $\text{Mn}^{2+}\text{(aq)}$.
6. Manganese metal is placed in a solution of CaCl_2 .	No reaction occurs.
7. Calcium metal is placed in a solution of KCl.	No reaction occurs.
8. A coil of zinc is placed in HCl(aq) .	Bubbles form as zinc disintegrates.

Procedure

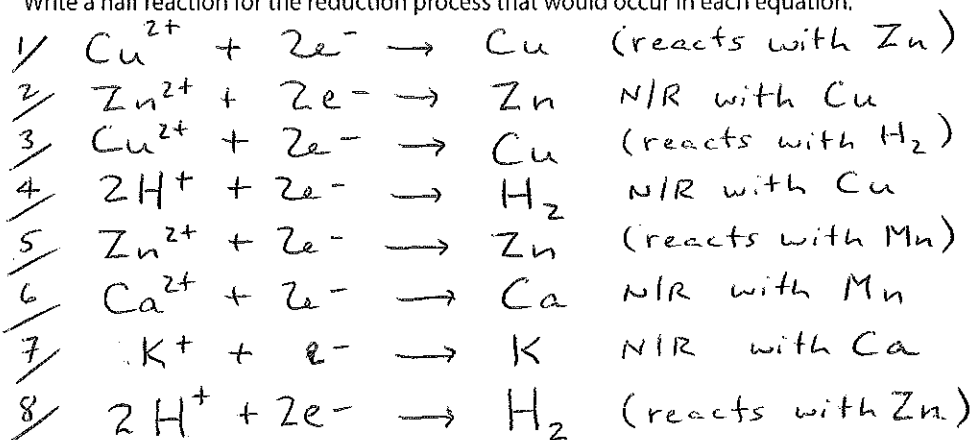
- Work with a partner or by yourself to design a table of reduction half reactions. Your table should be organized with the strongest oxidizing agent (the species most likely to be reduced) at the top. See the Results and Discussion section below for more ideas.
- Compare your completed table with those of other groups.

Results and Discussion

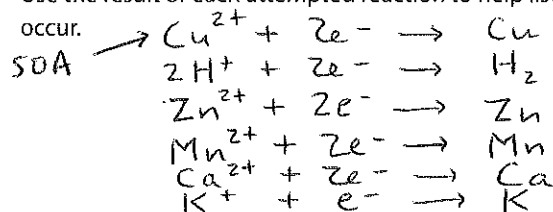
- Write a balanced equation for each attempted reaction (successful and *unsuccessful*) in the (data table above.)



2. Write a half reaction for the reduction process that would occur in each equation.



3. Use the result of each attempted reaction to help list the reductions from the most likely to the least likely to occur.



4. Label the species at the top of your list as the "strongest oxidizing agent."

5. What role do metals always play in these reactions (oxidizing or reducing agent)?

Reducing Agent

6. Where would the strongest reducing agent be located?

Bottom, Right

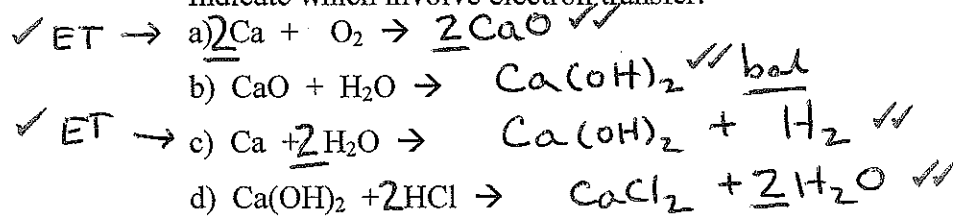
4.3 Review Questions (p. 196)

1. Give the oxidation state for the underlined element in each of the following species:



2. Complete and balance each of the following reactions. (States are not required.)

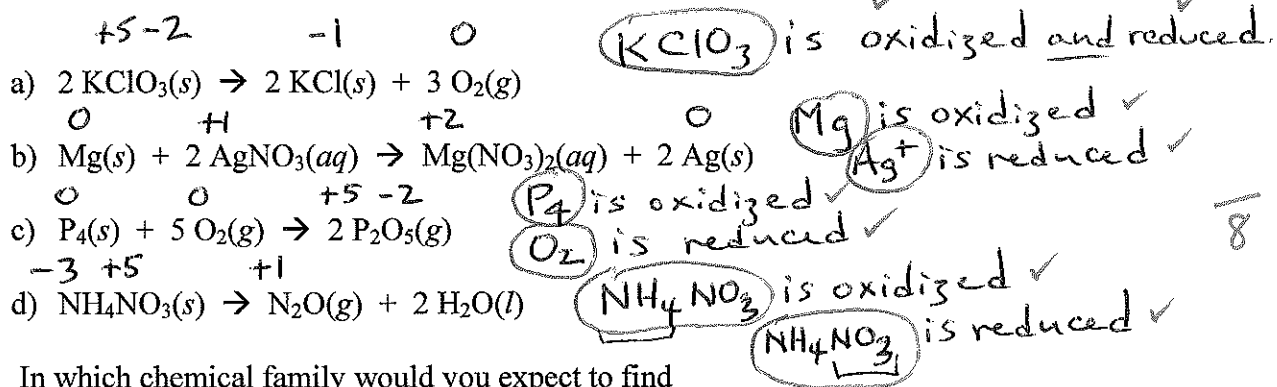
Indicate which involve electron transfer.



3. a) What is an oxidizing agent? Species that gets reduced. (O.N. ↓)
 (causes another species to be oxidized)
 b) What is a reducing agent? Species that gets oxidized. (O.N. ↑)
 (causes another to be reduced)
 c) How would you expect electronegativity to be related to the strength of each?

$\uparrow \text{EN} \rightarrow \text{Stronger O.A. / Weaker R.A.}$

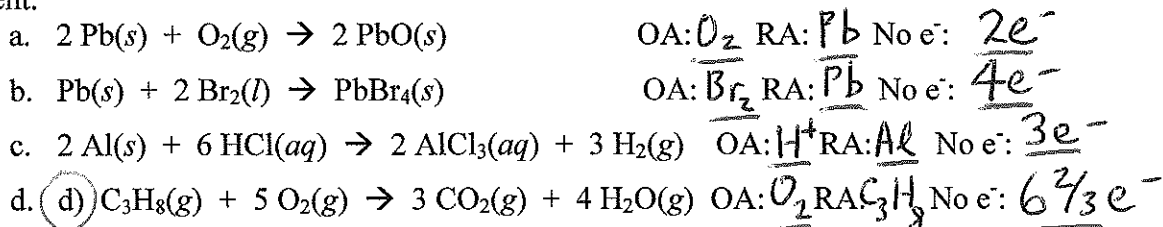
4. For each of the following reactions, indicate the species being oxidized and reduced and show the oxidation states above their symbols.



5. In which chemical family would you expect to find

- a) the most readily oxidized elements? Explain
Alkali metals (IA) - lose e^- easily (low EN)
- b) the most readily reduced elements? Explain
Halogens (VII A) or 17 - gain e^- easily, (high EN)

6. Determine the oxidizing and reducing agent in each of the following reactions. Then indicate the number of electrons transferred by one atom of the reducing agent.



7. Predict the products and write balanced equations for each of the following reactions. Include phase subscripts. Classify each equation according to the system learned in section 6.2. Then classify each again as redox or not.

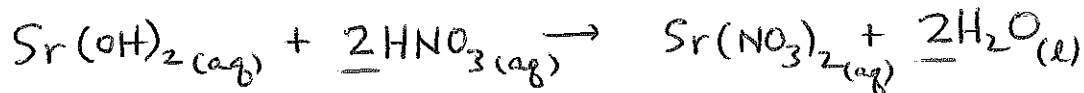
✓ S / ✓ Not → a) Sulphur dioxide gas reacts with solid calcium oxide.



D / Redox → b) Solid potassium nitrate is heated to form solid potassium nitrite and a gas.



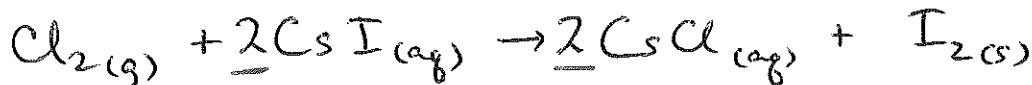
N / Not → c) Strontium hydroxide solution reacts with nitric acid.



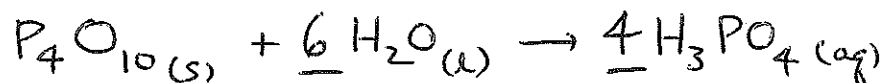
D / Not → d) Aluminum hydroxide solid is decomposed.



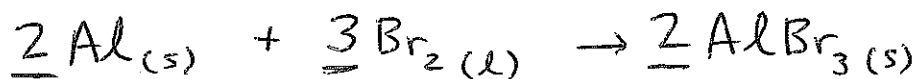
SR / Redox → e) Chlorine gas is bubbled through a solution of cesium iodide.



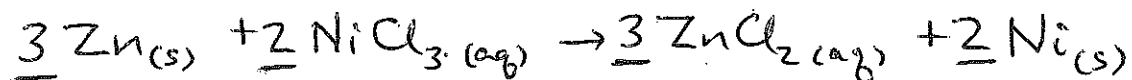
S / Not $\rightarrow$ f) Tetraphosphorus decaoxide solid is reacted with water.



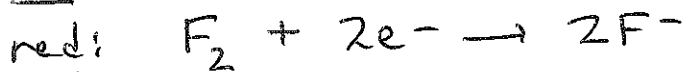
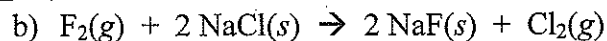
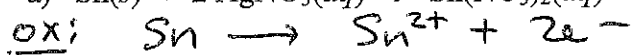
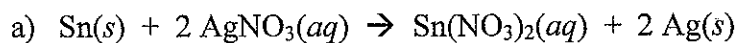
S / Redox $\rightarrow$ g) Aluminum metal reacts with liquid bromine.



SR / Redox $\rightarrow$ h) Zinc metal is placed in a solution of nickel(III) chloride.



8. Give an oxidation and a reduction half reaction for each of the following redox reactions:



4.4 Energy Changes Associated with Chemical Change — Endothermicity and Exothermicity

Warm Up (p. 198)

Complete the following table by placing a checkmark in the appropriate energy-change column for each of the classifications of chemical change listed.

Reaction Type	Energy is Released	Energy is Absorbed
Most Synthesis (Combination) Reactions		✓
Most Decomposition Reactions		✓
Neutralization Reactions	✓	
Combustion Reactions	✓	

Quick Check (p. 199)

Examine each of the equations listed on the previous page. Consider the sign of the ΔH value and determine whether the reactants or products have more stored energy in each case. Put a checkmark under the side with more enthalpy.

Reaction	Reactants	Products
Dissolving potassium hydroxide	✓	
Combustion of propane	✓	
Melting ice		✓
Replacement of iron by aluminum	✓	
Formation of calcium hydroxide	✓	

Quick Check (p. 201)

Complete the right column in the following table by labelling each of the introductory reactions in this section as endothermic or exothermic.

Reaction	Endothermic or Exothermic
Dissolving potassium hydroxide	exothermic
Combustion of propane	exothermic
Melting ice	endothermic
Replacement of iron by aluminum	exothermic
Formation of calcium hydroxide	exothermic

Practice Problems — Representing Exothermic and Endothermic Changes (p. 203)

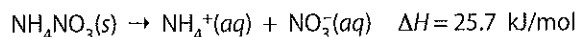
Given the following ΔH values, write a balanced thermochemical equation and an equation using ΔH notation with the smallest possible whole number coefficients for each of the following chemical changes:

1. $\Delta H_{\text{combustion}}$ of $\text{C}_2\text{H}_6(\text{g}) = -1428.5 \text{ kJ/mol}$
 $2\text{C}_2\text{H}_6(\text{g}) + 7\text{O}_2(\text{g}) \rightarrow 6\text{H}_2\text{O}(\text{l}) + 4\text{CO}_2(\text{g}) + 2857 \frac{\text{kJ}}{\text{mol}}$
 $2\text{C}_2\text{H}_6(\text{g}) + 7\text{O}_2(\text{g}) \rightarrow 6\text{H}_2\text{O}(\text{l}) + 4\text{CO}_2(\text{g}) \Delta H = -2857 \frac{\text{kJ}}{\text{mol}}$
2. $\Delta H_{\text{decomposition}}$ of $\text{NH}_3(\text{g}) = 46.1 \text{ kJ/mol}$
 $2\text{NH}_3(\text{g}) + 92.2 \frac{\text{kJ}}{\text{mol}} \rightarrow \text{N}_2(\text{g}) + 3\text{H}_2(\text{g})$
 $2\text{NH}_3(\text{g}) \rightarrow \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \Delta H = 92.2 \frac{\text{kJ}}{\text{mol}}$
3. $\Delta H_{\text{formation}}$ of $\text{HBr}(\text{g}) = -36.1 \text{ kJ/mol}$
 $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr}(\text{g}) + 72.2 \frac{\text{kJ}}{\text{mol}}$
 $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr}(\text{g}) \Delta H = -72.2 \frac{\text{kJ}}{\text{mol}}$

4.4 Activity: Designing a Household Product (p. 204)

The formation of iron(III) oxide, from iron powder and oxygen gas from the air is very exothermic and produces 826 kJ/mol of iron(III) oxide ore.

The dissolving of ammonium nitrate in water produces a solution of separated ammonium and nitrate ions. This physical change can be shown as

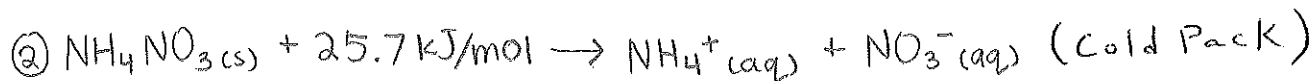
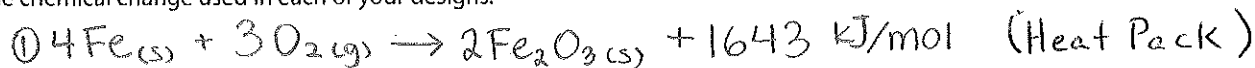


Procedure

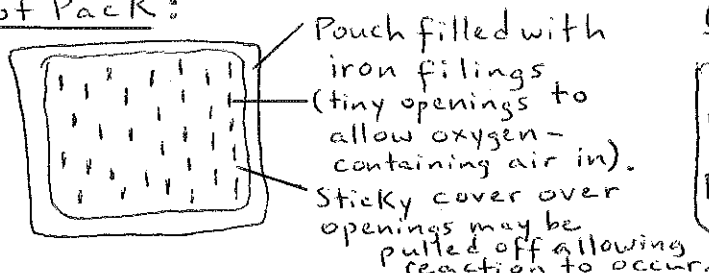
1. Use the theoretical information provided above to design two useful products for treating athletic injuries. Present your designs in each of the following formats:
 - Describe each design in point form.
 - Provide a labeled diagram of each design.
2. Compare your designs with those of a classmate. Use your comparison to improve your designs.

Results and Discussion

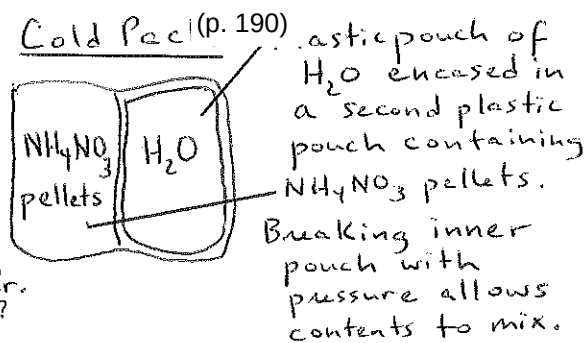
1. Write a balanced thermochemical equation using the smallest possible whole number coefficients to describe the chemical change used in each of your designs.



Hot Pack:



Cold Pack (p. 190)



2. Which of the two reactions is useful for cooling a warm, inflamed injury?

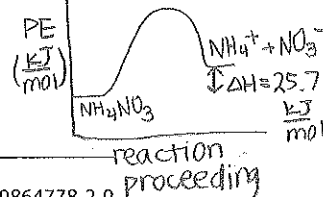
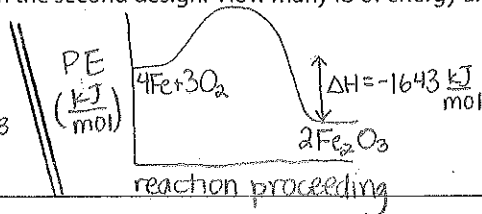
②

3. Which of the two reactions involves more bond breaking than bond forming?

②

4. Assume 240 g of ammonium nitrate are dissolved in the second design. How many kJ of energy are absorbed during this reaction?

$$240 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.0 \text{ g NH}_4\text{NO}_3} \times \frac{25.7 \text{ kJ}}{\text{mol NH}_4\text{NO}_3} = 77 \text{ kJ}$$



4.4 Review Questions (p. 205)

1. Indicate whether each of the following changes is endothermic or exothermic:

(a) Barbecuing a steak

endothermic

(b) Freezing a tray full of water to make ice

exothermic

(c) Neutralizing an acid spill with baking soda

exothermic

(d) Making a grilled cheese sandwich

endothermic

(e) Lighting a barbecue igniter

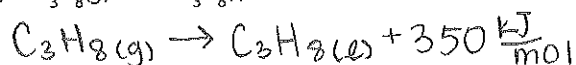
endothermic

(f) Condensing water on a mirror

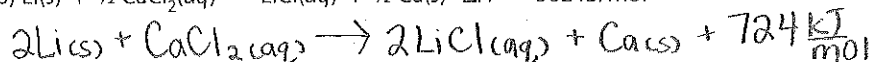
exothermic

2. Convert the following ΔH notation equations into thermochemical equations using the smallest whole number coefficients possible:

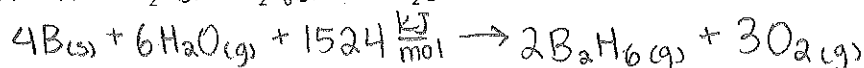
(a) $\frac{1}{2} \text{C}_3\text{H}_8(g) \rightarrow \frac{1}{2} \text{C}_3\text{H}_8(l) \quad \Delta H = -175 \text{ kJ/mol}$



(b) $\text{Li}(s) + \frac{1}{2} \text{CaCl}_2(aq) \rightarrow \text{LiCl}(aq) + \frac{1}{2} \text{Ca}(s) \quad \Delta H = -362 \text{ kJ/mol}$



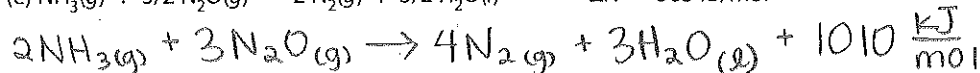
(c) $2 \text{B}(s) + 3 \text{H}_2\text{O}(g) \rightarrow \text{B}_2\text{H}_6(g) + \frac{3}{2} \text{O}_2(g) \quad \Delta H = 762 \text{ kJ/mol}$



(d) $\frac{1}{2} \text{P}_4(s) + 3 \text{Cl}_2(g) \rightarrow 2 \text{PCl}_3(s) \quad \Delta H = -613 \text{ kJ/mol}$



(e) $\text{NH}_3(g) + \frac{3}{2} \text{N}_2\text{O}(g) \rightarrow 2 \text{N}_2(g) + \frac{3}{2} \text{H}_2\text{O}(l) \quad \Delta H = -505 \text{ kJ/mol}$

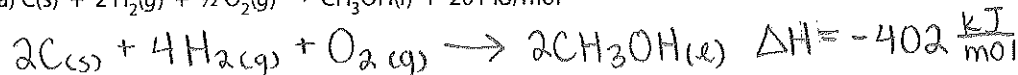


(f) $\frac{1}{2} \text{Fe}_3\text{O}_4(s) + \frac{1}{2} \text{CO}(g) \rightarrow \frac{3}{2} \text{FeO}(s) + \frac{1}{2} \text{CO}_2(g) \quad \Delta H = 9 \text{ kJ/mol}$

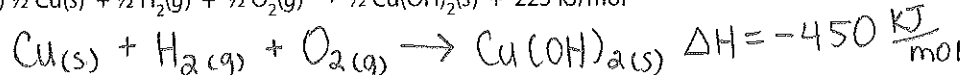


3. Convert the following thermochemical equations into ΔH notation using the smallest whole number coefficients possible.

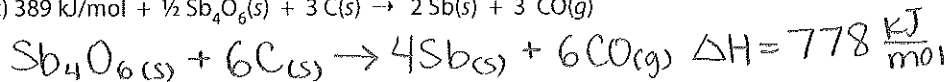
(a) $\text{C}(s) + 2 \text{H}_2(g) + \frac{1}{2} \text{O}_2(g) \rightarrow \text{CH}_3\text{OH}(l) + 201 \text{ kJ/mol}$



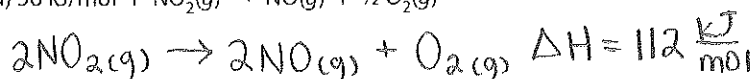
(b) $\frac{1}{2} \text{Cu}(s) + \frac{1}{2} \text{H}_2(g) + \frac{1}{2} \text{O}_2(g) \rightarrow \frac{1}{2} \text{Cu}(\text{OH})_2(s) + 225 \text{ kJ/mol}$

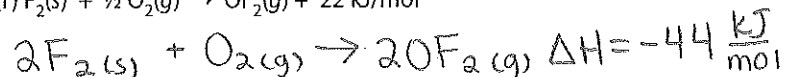
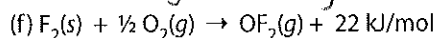
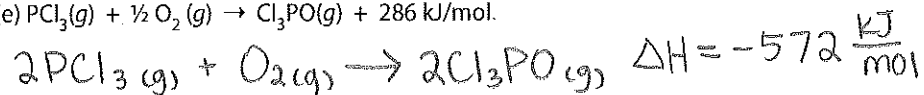
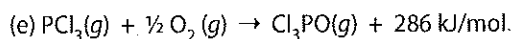


(c) $389 \text{ kJ/mol} + \frac{1}{2} \text{Sb}_4\text{O}_6(s) + 3 \text{C}(s) \rightarrow 2 \text{Sb}(s) + 3 \text{CO}(g)$



(d) $56 \text{ kJ/mol} + \text{NO}_2(g) \rightarrow \text{NO}(g) + \frac{1}{2} \text{O}_2(g)$





4. Use the equations in question 3 to answer the following questions:

(a) How much energy would be released during the formation of 4 mol of methanol?

$\frac{4 \text{ mol CH}_3\text{OH}}{\text{mol rxn}} \times \frac{1 \text{ mol rxn}}{2 \text{ mol CH}_3\text{OH}} \times \frac{-402 \text{ kJ}}{1 \text{ mol rxn}} = 802 \frac{\text{kJ}}{\text{mol rxn}}$

(b) How many moles of nitrogen dioxide could be decomposed through the use of 168 kJ of energy?

$168 \text{ kJ} \times \frac{1 \text{ mol rxn}}{112 \text{ kJ}} \times \frac{2 \text{ mol NO}_2}{1 \text{ mol rxn}} = 3.00 \text{ mol NO}_2$

(c) Is more energy absorbed or released during the formation of Cl_3PO gas from PCl_3 and O_2 gas?

released (Exo)

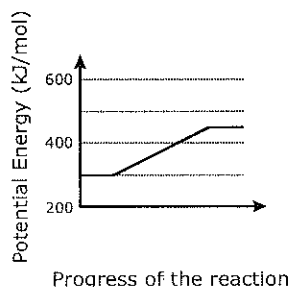
(d) What is the ΔH value for the decomposition of OF_2 gas into its elements?

$\Delta H_{\text{decomposition of OF}_2} = 44 \text{ kJ/mol OF}_2$

(e) How much energy is required to decompose 1 mol of copper(II) hydroxide?

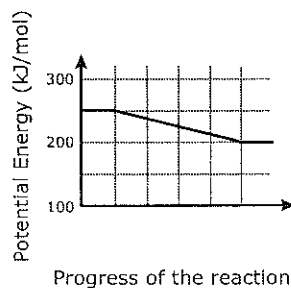
$450. \text{ kJ/mol Cu(OH)}_2$

5. Does the following potential energy diagram represent an endothermic or an exothermic reaction? What is ΔH for this reaction?



$\Delta H = 150 \frac{\text{kJ}}{\text{mol}}$
endothermic

6. What is ΔH for this reaction?

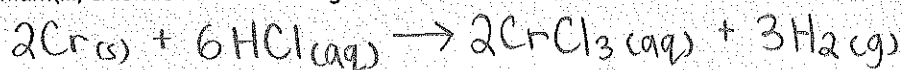


$\Delta H = -50 \frac{\text{kJ}}{\text{mol}}$
exothermic

4.5 Calculating with Chemical Change — Stoichiometry

Warm Up (p. 207)

1. Write a balanced equation for the reaction between chromium metal and hydrochloric acid to form chromium(III) chloride and a reactive gas.



2. If 2 mol of chromium metal completely reacted in an excess quantity of hydrochloric acid, how many moles of hydrogen gas would be formed?

$$2\text{ mol Cr} \times \frac{3\text{ mol H}_2}{2\text{ mol Cr}} = \boxed{3\text{ mol H}_2}$$

From equation above as well.

3. How many moles of hydrogen gas would form if 4 mol of chromium metal were completely reacted in excess hydrochloric acid?

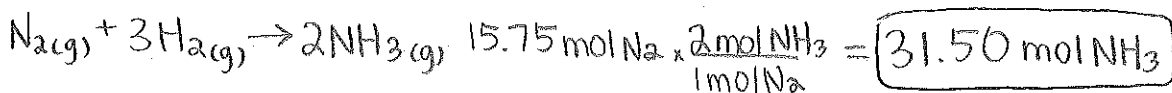
$$4\text{ mol Cr} \times \frac{3\text{ mol H}_2}{2\text{ mol Cr}} = \boxed{6\text{ mol H}_2}$$

Double last value

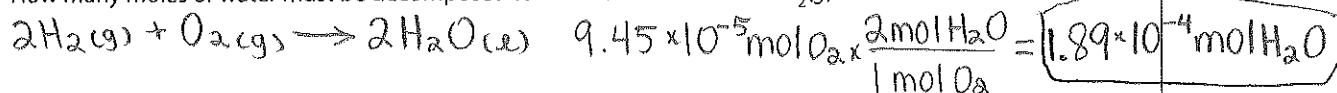
Practice Problems — Moles of One Species $\leftrightarrow$ Moles of Another Species (p. 208)

Calculate the moles of the requested species given the following information. Be sure to begin with a balanced chemical equation.

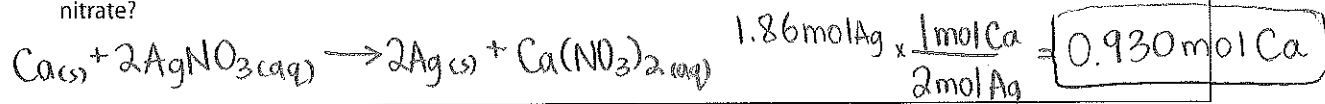
1. How many moles of ammonia ($\text{NH}_3\text{(g)}$) would be formed by the combination of 15.75 mol of nitrogen gas ($\text{N}_2\text{(g)}$) with a large excess of hydrogen gas ($\text{H}_2\text{(g)}$)?



2. How many moles of water must be decomposed to form 9.45×10^{-5} mol of $\text{O}_2\text{(g)}$?



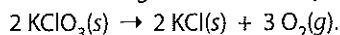
3. How many moles of calcium metal would be required to replace 1.86 mol of silver from a solution of silver nitrate?



Practice Problems — Mass of One Species $\leftrightarrow$ Mass of Another Species (p. 210)

Show all work for each of the following questions. Begin with a balanced chemical equation if one is not provided. Be sure to include the formulas as part of your units in each step. Give your final answer with the correct number of significant figures.

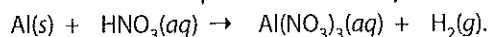
1. The following reaction is exceptional in that a single reactant species is both oxidized and reduced:



What mass of oxygen gas would be formed by the decomposition of 45.65 g of KClO_3 ? (How should your answer compare to the mass of reactant? Does your answer make sense?)

$$45.65 \text{ g KClO}_3 \times \frac{1 \text{ mol}}{122.6 \text{ g}} \times \frac{3 \text{ mol O}_2}{2 \text{ mol}} \times \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} = \boxed{17.9 \text{ g}}$$

2. This *unbalanced* equation shows one possible set of products for the reaction of aluminum and nitric acid:



Use the balanced version of this equation to determine the mass of nitric acid required to form 170 kg of hydrogen gas.

$$170 \text{ kg} \times \frac{1 \text{ mol H}_2}{0.0020 \text{ kg}} \times \frac{2 \text{ mol HNO}_3}{1 \text{ mol H}_2} \times \frac{63.0 \text{ g HNO}_3}{1 \text{ mol HNO}_3} = \boxed{1.1 \times 10^7 \text{ g}} \text{ (or } \underline{\underline{11,000 \text{ kg}}})$$

3. Hydrogen gas can be used as a reducing agent in the production of pure boron from boron trichloride. How many moles of boron would form from the reduction of 500.0 g of boron trichloride with excess hydrogen gas?



$$500.0 \text{ g BCl}_3 \times \frac{1 \text{ mol}}{117.3 \text{ g}} \times \frac{2 \text{ mol B}}{2 \text{ mol BCl}_3} = \boxed{4.263 \text{ mol}}$$

4. Write a balanced chemical equation for the reaction between sodium metal and water. How many moles of hydrogen gas would form if 2 mol of sodium reacted with excess water? What volume would this gas occupy at STP conditions?



$$2 \text{ mol Na} \times \frac{1 \text{ mol H}_2}{2 \text{ mol Na}} = \boxed{1 \text{ mol H}_2} \quad 1 \text{ mole of gas} = \boxed{22.4 \text{ L}}$$

Quick Check (p. 211)

Use the balanced equation in Figure 4.5.3 to determine how many litres of carbon dioxide would form from the combustion of 10 L of propane gas.

$$10 \text{ L C}_3\text{H}_8 \times \frac{3 \text{ L CO}_2}{1 \text{ L C}_3\text{H}_8} = \boxed{30 \text{ L}}$$

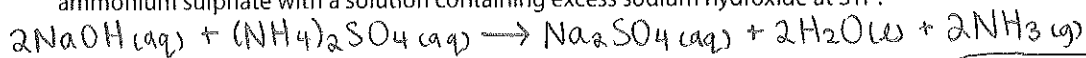
Practice Problems — Calculations Involving Gas Volume (p. 213)

Show all work for each of the following questions. Begin with a balanced chemical equation if one is not provided. Be sure to include the formulas as part of your units in each step. Give your final answer with the correct number of significant figures.

1. The reaction, $2 \text{ SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{ SO}_3(\text{g})$, occurs in the atmosphere when oxygen oxidizes the pollutant factory exhaust, sulphur dioxide. What volume of oxygen is required to oxidize the $2.41 \times 10^7 \text{ L}$ of sulphur dioxide released annually by a coal-based power plant at STP?

$$2.41 \times 10^7 \text{ L SO}_2 \times \frac{1 \text{ L O}_2}{2 \text{ L SO}_2} = \boxed{1.21 \times 10^7 \text{ L O}_2}$$

2. Consider the *unbalanced equation*: $\text{NaOH}(\text{aq}) + (\text{NH}_4)_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{NH}_3(\text{g})$. How many litres of ammonia, measured at STP, are formed by combining a solution containing 40.5 g of ammonium sulphate with a solution containing excess sodium hydroxide at STP?



$$40.5 \text{ g } (\text{NH}_4)_2\text{SO}_4 \times \frac{1 \text{ mol}}{132.1 \text{ g}} \times \frac{2 \text{ mol NH}_3}{1 \text{ mol } (\text{NH}_4)_2\text{SO}_4} \times \frac{22.4 \text{ L NH}_3}{1 \text{ mol NH}_3} = \boxed{13.7 \text{ L}}$$

3. What mass of HCl must be contained in a solution added to an excess of zinc arsenide to obtain 17.5 L of arsine (AsH_3) gas at STP?



$$17.5 \text{ L} \times \frac{1 \text{ mol AsH}_3}{22.4 \text{ L AsH}_3} \times \frac{6 \text{ mol HCl}}{2 \text{ mol AsH}_3} \times \frac{36.5 \text{ g HCl}}{1 \text{ mol HCl}} = \boxed{85.5 \text{ g}}$$

Quick Check (p. 214)

1. What is the concentration of a solution made by dissolving 117 g of sodium chloride in water to make a total volume of 1.5 L?

$$\frac{117 \text{ g NaCl}}{1.5 \text{ L}} \times \frac{1 \text{ mol}}{58.5 \text{ g}} = \boxed{1.3 \text{ mol/L}}$$

2. Use the molarity of the solution from question 1 to determine the mass of NaCl in 250 mL.

$$250 \text{ mL} \times \frac{10^{-3} \text{ L}}{1 \text{ mL}} \times \frac{1.3 \text{ mol}}{1 \text{ L}} \times \frac{58.5 \text{ g}}{1 \text{ mol}} = 19.5 \rightarrow \boxed{20. \text{ g}}$$

3. What volume of a solution of this same molarity would contain 39.0 g of sodium chloride?

$$39.0 \text{ g} \times \frac{1 \text{ mol}}{58.5 \text{ g}} \times \frac{1 \text{ L}}{1.3 \text{ mol}} = \boxed{0.50 \text{ L}}$$

Practice Problems — Calculations Involving Solution Volumes (p. 215)

Show all work for each of the following questions. Begin with a balanced chemical equation if one is not provided. Be sure to include the formulas as part of your units in each step. Give your final answer with the correct number of significant figures.

1. The exothermic reaction $2\text{K}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{KOH}(aq) + \text{H}_2(g)$ often releases enough heat to ignite the hydrogen gas it produces. What mass of potassium metal would be required to produce 250 mL of a 0.45 mol/L solution of potassium hydroxide?

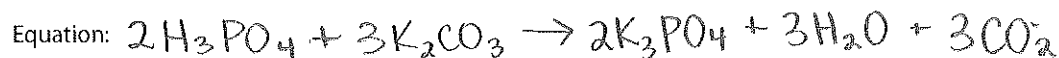
$$250\text{ mL} \times \frac{0.45\text{ mol KOH}}{1000\text{ mL}} \times \frac{2\text{ mol K}}{2\text{ mol KOH}} \times \frac{39.1\text{ g K}}{1\text{ mol K}} = \boxed{4.4\text{ g}}$$

2. What volume of 0.80 mol/L sodium iodide solution would completely react with 2.4×10^{24} molecules of chlorine gas?



$$2.4 \times 10^{24}\text{ molecules Cl}_2 \times \frac{1\text{ mol Cl}_2}{6.02 \times 10^{23}\text{ molec}} \times \frac{2\text{ mol NaI}}{1\text{ mol Cl}_2} \times \frac{1\text{ L}}{0.80\text{ mol NaI}} = 9.97\text{ L} \rightarrow \boxed{10.\text{ L}}$$

3. How many litres of carbon dioxide gas would be formed at STP if 1.5 L of 1.75 M phosphoric acid were reacted with excess potassium carbonate solution? Be cautious, as this question requires the use of both molar volume and molarity as conversion factors.



$$1.5\text{ L} \times \frac{1.75\text{ mol}}{1\text{ L}} \times \frac{3\text{ mol CO}_2}{2\text{ mol H}_3\text{PO}_4} \times \frac{22.4\text{ L CO}_2}{1\text{ mol CO}_2} = \boxed{88\text{ L}}$$

4.5 Activity: Stoichiometric Relationships of Hydrochloric Acid (p. 216)

Question

How can you measure the products of a reaction?

Background

Hydrochloric acid is the most commonly used acid in secondary school chemistry. It can be reacted with a variety of substances in several different types of chemical reactions. The products of such reactions may be separated and measured to study stoichiometric relationships in a variety of ways.

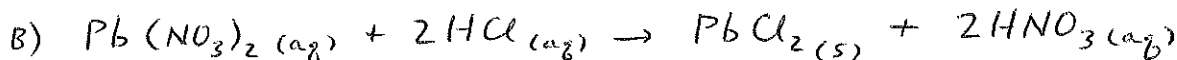
Procedure

- Use the theoretical information provided in the "Background" above to design two different experiments to isolate and measure a product of two different reactions, each consuming 750 mL of 1.0 M HCl. You must use a single and a double replacement reaction.
 - Describe each design in point-form.
 - Provide a labelled diagram of each design.
- Compare your designs with those of a classmate. Use your comparison to improve your designs.

Example Only :-

Results and Discussion

- Write a balanced chemical equation for each reaction you selected.



- Calculate the quantity of product you would expect to measure in each experiment.

A) $750\text{ mL} \times \frac{1.0\text{ mol HCl}}{1000\text{ mL}} \times \frac{1\text{ mol H}_2}{2\text{ mol HCl}} \times \frac{22.4\text{ L H}_2}{1\text{ mol H}_2} = 8.4\text{ L}$

B) $750\text{ mL} \times \frac{1.0\text{ mol HCl}}{1000\text{ mL}} \times \frac{1\text{ mol PbCl}_2}{2\text{ mol HCl}} \times \frac{278.2\text{ g PbCl}_2}{1\text{ mol PbCl}_2} = 104.3 \rightarrow 1.0 \times 10^2\text{ g}$

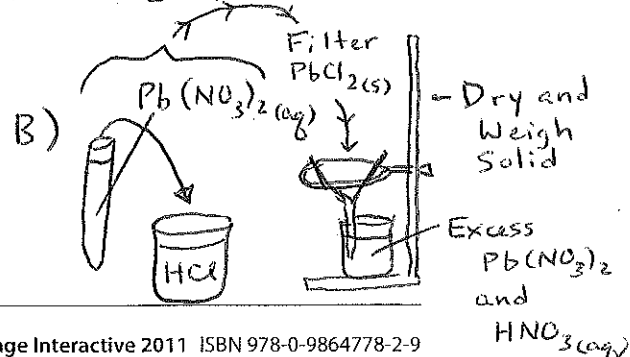
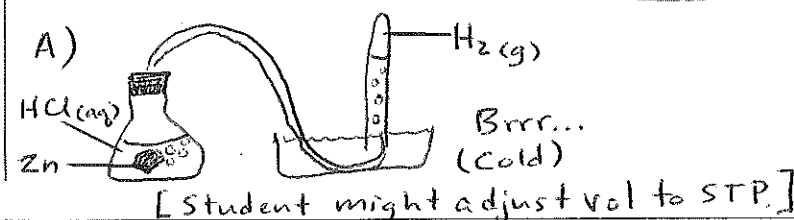
- Exchange your answers to questions 1 and 2 with the same classmate with whom you compared experimental designs. Beginning with the classmate's answers to question 2, calculate the volume of 1.0 mol/L HCl required to form the quantity of product he or she determined for each experiment. Your classmate should perform the same calculations with your question 2 answers.

A) $8.4\text{ L H}_2 \times \frac{1\text{ mol H}_2}{22.4\text{ L}} \times \frac{2\text{ mol HCl}}{1\text{ mol H}_2} \times \frac{1000\text{ mL}}{1.0\text{ mol HCl}} = 750\text{ mL}$

B) $104.3\text{ g PbCl}_2 \times \frac{1\text{ mol PbCl}_2}{278.2\text{ g}} \times \frac{2\text{ mol HCl}}{1\text{ mol PbCl}_2} \times \frac{1000\text{ mL}}{1.0\text{ mol HCl}} = 750\text{ mL}$

- What answer should all students get for question 3?

(see #3) $\rightarrow 750\text{ mL}$



4.5 Review Questions (p. 217)

$$1. a) 1.26 \text{ mol M} \times \frac{2 \text{ mol CuO}}{1 \text{ mol M}} = 2.52 \text{ mol CuO} //$$

$$b) 1.5 \text{ Kg} \times \frac{10^3 \text{ g}}{1 \text{ Kg}} \times \frac{1 \text{ mol M}}{221.0 \text{ g}} \times \frac{2 \text{ mol CuO}}{1 \text{ mol M}} \times \frac{79.5 \text{ g}}{1 \text{ mol}} = 1100 \text{ g} //$$

$$c) 706 \text{ g} \times \frac{1 \text{ mol}}{79.5 \text{ g CuO}} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol CuO}} \times \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} = 99.5 \text{ L} //$$

$$2. a) 3160 \text{ g} \times \frac{1 \text{ mol CH}_3\text{NO}_2}{61.0 \text{ g}} \times \frac{2 \text{ mol N}_2}{4 \text{ mol CH}_3\text{NO}_2} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 580. \text{ L} //$$

$$b) 955 \text{ g} \times \frac{1 \text{ mol N}_2}{28.0 \text{ g}} \times \frac{4 \text{ mol CH}_3\text{NO}_2}{2 \text{ mol N}_2} \times \frac{61.0 \text{ g}}{1 \text{ mol}} = 4160 \text{ g} //$$

$$c) 3.5 \times 10^{25} \text{ molec N}_2 \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ molec}} \times \frac{6 \text{ mol H}_2\text{O}}{2 \text{ mol N}_2} \times \frac{18.0 \text{ g}}{1 \text{ mol}} = 3100 \text{ g} //$$

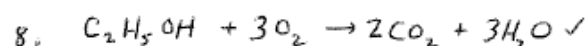
$$3. 10.0 \text{ mL} \times \frac{0.45 \text{ mol HCl}}{1000 \text{ mL}} \times \frac{1 \text{ mol Zn}}{2 \text{ mol HCl}} \times \frac{65.4 \text{ g}}{1 \text{ mol}} = 0.15 \text{ g}$$

$$4. 12.2 \text{ g Na} \times \frac{1 \text{ mol Na}}{23.0 \text{ g Na}} \times \frac{124.7 \text{ KJ}}{4 \text{ mol Na}} = 16.5 \text{ KJ} //$$

$$5. 3.225 \text{ g} \times \frac{1 \text{ mol H}_2\text{C}_2\text{O}_4}{90.0 \text{ g}} \times \frac{2 \text{ mol KMnO}_4}{5 \text{ mol H}_2\text{C}_2\text{O}_4} \times \frac{1000 \text{ mL}}{0.250 \text{ mol}} = 57.3 \text{ mL}$$

$$6. 2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3 \quad 4.56 \text{ Kg} \times \frac{1 \text{ mol}}{0.1335 \text{ Kg}} \times \frac{3 \text{ mol Cl}_2}{2 \text{ mol AlCl}_3} \times \frac{71.0 \text{ g}}{1 \text{ mol}} = 3640 \text{ g} //$$

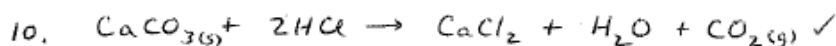
$$7. \text{H}_2\text{SO}_4 + 2\text{KOH} \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O} \quad 0.034 \text{ mol KOH} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol KOH}} = 0.017 \text{ mol} //$$



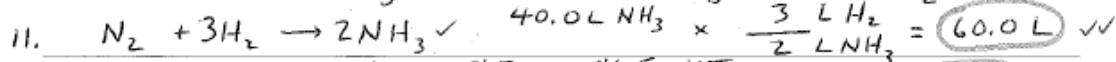
$$35.00 \text{ g C}_2\text{H}_5\text{OH} \times \frac{1 \text{ mol}}{46.0 \text{ g}} \times \frac{3 \text{ mol H}_2\text{O}}{1 \text{ mol C}_2\text{H}_5\text{OH}} \times \frac{18.0 \text{ g}}{1 \text{ mol}} = 41.1 \text{ g} //$$



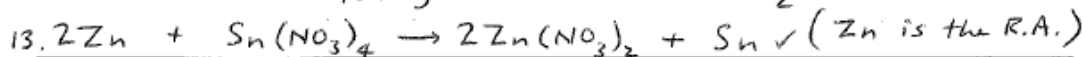
$$21.7 \text{ L} \times \frac{1 \text{ mol H}_2\text{S}}{22.4 \text{ L}} \times \frac{1 \text{ mol FeS}}{1 \text{ mol H}_2\text{S}} \times \frac{87.9 \text{ g}}{1 \text{ mol}} = 85.2 \text{ g} //$$



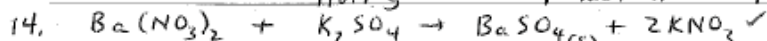
$$15.0 \text{ g CaCO}_3 \times \frac{1 \text{ mol CaCO}_3}{100.1 \text{ g}} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CaCO}_3} \times \frac{44.0 \text{ g CO}_2}{1 \text{ mol CO}_2} = 6.59 \text{ g} //$$



$$12. 5.00 \text{ g PbI}_2 \times \frac{1 \text{ mol PbI}_2}{461.0 \text{ g}} \times \frac{46.5 \text{ KJ}}{1 \text{ mol PbI}_2} = 0.504 \text{ KJ} //$$



$$27.5 \text{ g Sn} \times \frac{1 \text{ mol Sn}}{118.7 \text{ g Sn}} \times \frac{2 \text{ mol Zn}}{1 \text{ mol Sn}} \times \frac{65.3 \text{ g Zn}}{1 \text{ mol Zn}} = 30.3 \text{ g} //$$



$$6.5 \text{ mol Ba}(\text{NO}_3)_2 \times \frac{1 \text{ mol BaSO}_4}{1 \text{ mol Ba}(\text{NO}_3)_2} \times \frac{233.4 \text{ g}}{1 \text{ mol}} = 1500 \text{ g} //$$

15. See Q 10.
for EQN

$$12.2 \text{ L CO}_2 \times \frac{1 \text{ mol CO}_2}{22.4 \text{ L CO}_2} \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol CO}_2} \times \frac{100.1 \text{ g}}{1 \text{ mol}} = 54.5 \text{ g} //$$

3 16) $\text{N}_2\text{O}_4 \rightarrow 2\text{NO}_2 \quad \Delta H = 56 \text{ kJ/mol N}_2\text{O}_4 \quad \checkmark$
 $1.25 \text{ g NO}_2 \times \frac{1 \text{ mol NO}_2}{46.0 \text{ g NO}_2} \times \frac{1 \text{ mol N}_2\text{O}_4}{2 \text{ mol NO}_2} \times \frac{56 \text{ kJ}}{1 \text{ mol N}_2\text{O}_4} = 0.76 \text{ kJ} \quad \checkmark$

3 17) $2\text{HBr} + \text{K}_2\text{CO}_3 \rightarrow 2\text{KBr} + \text{H}_2\text{O} + \text{CO}_2 \quad \checkmark$
 $450 \text{ mL} \times \frac{0.500 \text{ mol HBr}}{1000 \text{ mL}} \times \frac{1 \text{ mol K}_2\text{CO}_3}{2 \text{ mol HBr}} \times \frac{138.2 \text{ g}}{1 \text{ mol}} = 16 \text{ g} \quad \checkmark$

3 18) $2\text{CH}_3\text{COOH} + \text{Pb}(\text{NO}_3)_2 \rightarrow \text{Pb}(\text{CH}_3\text{COO})_2(\text{s}) + 2\text{HNO}_3 \quad \checkmark$
 $8.64 \text{ g} \times \frac{1 \text{ mol Pb}(\text{CH}_3\text{COO})_2}{325.2 \text{ g}} \times \frac{2 \text{ mol CH}_3\text{COOH}}{1 \text{ mol Pb}(\text{CH}_3\text{COO})_2} \times \frac{1 \text{ L}}{2.5 \text{ mol}} = 0.021 \text{ L} \quad (21 \text{ mL}) \quad \checkmark$

3 19) $3\text{NaOH} + \text{H}_3\text{PO}_4 \rightarrow \text{Na}_3\text{PO}_4 + 3\text{H}_2\text{O} \quad \checkmark$
 $25.0 \text{ mL} \times \frac{0.356 \text{ mol H}_3\text{PO}_4}{1000 \text{ mL}} \times \frac{3 \text{ mol NaOH}}{1 \text{ mol H}_3\text{PO}_4} \times \frac{1000 \text{ mL}}{0.610 \text{ mol}} = 43.8 \text{ mL} \quad \checkmark$

3 20) $\text{Zn} + 2\text{HI} \rightarrow \text{ZnI}_2 + \text{H}_2 \quad \checkmark$
 $150 \text{ mL} \times \frac{0.185 \text{ mol HI}}{1000 \text{ mL}} \times \frac{1 \text{ mol H}_2}{2 \text{ mol HI}} \times \frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2} = 0.31 \text{ L} \quad \checkmark$

4.6 Stoichiometry in the Real World — Excess/Limiting Amounts, Percentage Yield, and Impurities

Warm Up (p. 220)

Suppose you are working in an ice cream shop. A group of customers come in and they would like to order ice cream sundaes. The shop's computer page reads as follows for ice cream sundaes:

Sundae Ingredients in Stock	Number of Items	Sundae Recipe
Scoops vanilla ice cream	six scoops	one scoop vanilla ice cream
Chocolate syrup	200 mL	50 ml chocolate syrup
Super sweet red cherries	475 cherries	one cherry on top
		(Do NOT change the recipe)

- What is the maximum number of sundaes you can make for customers? **4**
- What item *limits* the number of sundaes you can make? **Chocolate Syrup**
- How much of each of the excess items will still be left as sundae stock?
2 scoops of ice cream
471 cherries



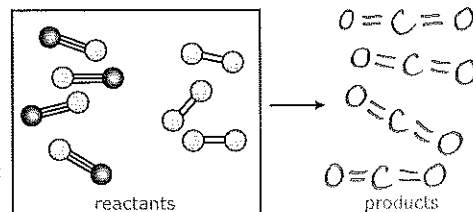
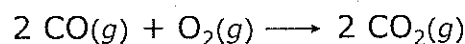
Quick Check (p. 221)

Consider the following situation:

- Which of the reactants is in excess? **O₂**
- Sketch any remaining reactants and products to show what will be present once the reaction is complete.

O=O
excess O₂

- How many CO₂ particles are formed? **4**
- How many excess reactant particles remain? **1**



Practice Problems — Limiting and Excess Stoichiometry (p. 223)

- 25.0 g each of zinc metal and HCl dissolved in aqueous solution are reacted together. What volume of hydrogen gas, measured at STP, is formed and what mass of excess reactant is left over?



$$25.0\text{g Zn} \times \frac{1\text{mol}}{65.4\text{g}} \times \frac{1\text{mol H}_2}{1\text{mol Zn}} \times \frac{22.4\text{L H}_2}{1\text{mol H}_2} = 8.56\text{L (Zn is in excess)}$$

$$25.0\text{g HCl} \times \frac{1\text{mol}}{36.5\text{g}} \times \frac{1\text{mol H}_2}{2\text{mol HCl}} \times \frac{22.4\text{L}}{1\text{mol}} = \boxed{7.67\text{L}}$$

$$25.0\text{g HCl} \times \frac{1\text{mol HCl}}{36.5\text{g HCl}} \times \frac{1\text{mol Zn}}{2\text{mol HCl}} \times \frac{65.4\text{g Zn}}{1\text{mol}} = 22.4\text{g Zn}_{\text{use}}$$

$$\therefore 25.0 - 22.4 = \boxed{2.6\text{g excess Zn}}$$

2. An acid spill containing 12.0 g of pure sulphuric acid is neutralized by 80.0 g of sodium bicarbonate (baking soda). What volume of water is formed? What mass of excess reactant is left over? (Reminder: density of water = 1.00 g/mL) $2\text{NaHCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O} + 2\text{CO}_2$

$$\text{excess} \rightarrow 80.0\text{g} \times \frac{1\text{ mol}}{84.0\text{g}} \times \frac{2\text{ mol H}_2\text{O}}{2\text{ mol NaHCO}_3} \times \frac{18.0\text{g H}_2\text{O}}{1\text{ mol H}_2\text{O}} \times \frac{1\text{ mL H}_2\text{O}}{1.00\text{g H}_2\text{O}} = 17.1\text{ mL}$$

$$\text{limiting} \rightarrow 12.0\text{g} \times \frac{1\text{ mol}}{98.1\text{g}} \times \frac{2\text{ mol H}_2\text{O}}{1\text{ mol H}_2\text{SO}_4} \times \frac{18.0\text{g H}_2\text{O}}{1\text{ mol H}_2\text{O}} \times \frac{1\text{ mL}}{1.00\text{g H}_2\text{O}} = 4.40\text{ mL}$$

$$12.0\text{g H}_2\text{SO}_4 \times \frac{1\text{ mol}}{98.1\text{g}} \times \frac{2\text{ mol NaHCO}_3}{1\text{ mol H}_2\text{SO}_4} \times \frac{84.0\text{g}}{1\text{ mol}} = 20.6\text{g NaHCO}_3 \text{ used}$$

3. When ammonia is passed over hot calcium, calcium hydride and nitrogen gas are formed. What mass of calcium hydride results when 20.0 L of ammonia measured at STP and 150 g of calcium are reacted? What quantity of excess reactant remains? $2\text{NH}_3 + 3\text{Ca} \rightarrow 3\text{CaH}_2 + \text{N}_2$

$$20.0\text{ L NH}_3 \times \frac{1\text{ mol NH}_3}{22.4\text{ L}} \times \frac{3\text{ mol CaH}_2}{2\text{ mol NH}_3} \times \frac{42.1\text{ g}}{1\text{ mol}} = 56.4\text{ g}$$

$$\text{excess} \rightarrow 150\text{g Ca} \times \frac{1\text{ mol}}{40.1\text{g}} \times \frac{3\text{ mol CaH}_2}{3\text{ mol Ca}} \times \frac{42.1\text{ g}}{1\text{ mol}} = 157\text{ g}$$

$$20.0\text{ L NH}_3 \times \frac{1\text{ mol}}{22.4\text{ L}} \times \frac{3\text{ mol Ca}}{2\text{ mol NH}_3} \times \frac{40.1\text{ g}}{1\text{ mol}} = 53.7\text{ g Ca used}$$

$$150 - 53.7 = 96.3 \rightarrow 1.0 \times 10^2\text{ g left Ca}$$

Quick Check (p. 224)

1. Determine the percentage purity of the following:

(a) A 4.5 g piece of calcium contains only 3.8 g of pure calcium metal.

$$\frac{3.8\text{ g}}{4.5\text{ g}} \times 100\% = 84\%$$

(b) Pure phosphorus makes up 17.5 g of a sample of white phosphorus weighing 20.0 g.

$$\frac{17.5\text{ g}}{20.0\text{ g}} \times 100\% = 87.5\%$$

2. What mass of pure potassium is contained in a 0.90 g piece of 90.0% pure potassium?

$$\frac{90.0}{100} (0.90\text{ g}) = 0.81\text{ g}$$

Practice Problems — Percentage Purity (p. 226)

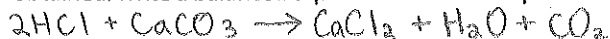
1. The thermite reaction is extremely exothermic, producing temperatures in excess of 2500°C. These high temperatures were used to weld iron tracks together during the early days of the railway. Thermite is a mixture of powdered aluminum and iron(III) oxide (rust) reacting as follows: $2\text{Al(s)} + \text{Fe}_2\text{O}_3\text{(s)} \rightarrow \text{Al}_2\text{O}_3\text{(s)} + 2\text{Fe(s)}$. If 2.44 g of 95% pure aluminum is reacted, how many grams of aluminum oxide can be produced?

$$0.95(2.44\text{g Al}) \times \frac{1\text{ mol}}{27.0\text{g}} \times \frac{1\text{ mol Al}_2\text{O}_3}{2\text{ mol Al}} \times \frac{102.0\text{ g}}{1\text{ mol}} = 4.4\text{ g}$$

2. How many grams of 73.0% pure iron(III) oxide are required to form 24.5 g of pure white hot iron metal in the thermite reaction? (Hint: Begin by determining the mass of pure oxide required.)

$$24.5\text{g Fe} \times \frac{1\text{ mol Fe}}{55.8\text{g}} \times \frac{1\text{ mol Fe}_2\text{O}_3}{2\text{ mol Fe}} \times \frac{159.6\text{g Fe}_2\text{O}_3 \text{ pure}}{1\text{ mol Fe}_2\text{O}_3} \times \frac{100\text{ impure}}{73.0\text{ pure Fe}_2\text{O}_3} = 48.0\text{ g}$$

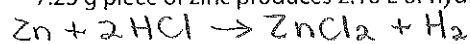
3. When hydrochloric acid is added to 5.73 g of contaminated calcium carbonate, 2.49 g of carbon dioxide is obtained. Write a balanced equation and find the percentage purity of the calcium carbonate.



$$2.49\text{g CO}_2 \times \frac{1\text{ mol CO}_2}{44.0\text{g}} \times \frac{1\text{ mol CaCO}_3}{1\text{ mol CO}_2} \times \frac{100.1\text{g}}{1\text{ mol}} = 5.66\text{g pure} \therefore \frac{5.66\text{g}}{5.73\text{g}} \times 100\% = 98.9\%$$

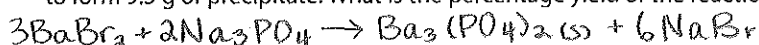
Practice Problems — Percentage Yield (p. 227)

1. A chunk of zinc metal reacts with an excess of hydrochloric acid solution. What is the percentage yield if a 7.23 g piece of zinc produces 2.16 L of hydrogen gas at STP? Begin with a balanced equation.



$$7.23 \text{ g Zn} \times \frac{1 \text{ mol}}{65.4 \text{ g}} \times \frac{1 \text{ mol H}_2}{1 \text{ mol Zn}} \times \frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2} = 2.48 \text{ L} \quad \therefore \frac{2.16 \text{ L}}{2.48 \text{ L}} \times 100\% = 87.2\%$$

2. A solution containing 15.2 g of barium bromide is reacted with a solution containing excess sodium phosphate to form 9.5 g of precipitate. What is the percentage yield of the reaction? Begin with a balanced equation.



$$15.2 \text{ g BaBr}_2 \times \frac{1 \text{ mol}}{297.1 \text{ g}} \times \frac{1 \text{ mol Ba}_3(\text{PO}_4)_2}{3 \text{ mol BaBr}_2} \times \frac{601.9 \text{ g}}{1 \text{ mol}} = 10.3 \text{ g} \quad \therefore \frac{9.5 \text{ g}}{10.3 \text{ g}} \times 100\% = 93\%$$

3. Copper(II) oxide reacts with hydrogen gas to form water and copper metal. From this reaction, 3.6 g of copper metal was obtained with a yield of 32.5%. What mass of copper(II) oxide was reacted with the excess hydrogen gas? Begin with a balanced equation. $\text{CuO} + \text{H}_2 \rightarrow \text{H}_2\text{O} + \text{Cu}$ (bal)

$$\text{yield of } 100\% \quad \frac{3.6 \text{ g}}{0.325} = 11 \text{ g} \quad 11 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.5 \text{ g}} \times \frac{1 \text{ mol CuO}}{1 \text{ mol Cu}} \times \frac{79.5 \text{ g CuO}}{1 \text{ mol CuO}} = 14 \text{ g}$$

4.6 Activity: Charting and Graphing Stoichiometry (p. 228)

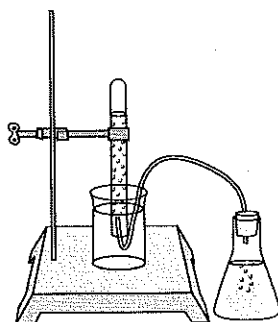
Question

What are some different ways to represent stoichiometric data?

Background

A student performs a single replacement reaction between pure magnesium metal and hydrochloric acid under two different conditions. The quantities involved are measured and recorded in tabular and graphic forms.

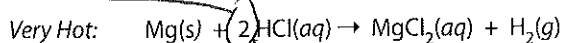
The reactions are performed in 250.0 mL of HCl(aq) . The magnesium is weighed before and after the reactions are completed. The reactions are allowed to proceed for the same period of time in both trials. In both cases, once the reaction is completed, the volume of gas is measured under STP conditions.



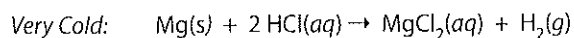
Measured Items	Very Hot	Very Cold
Initial mass of Mg(s)	1.00 g	1.00 g
Final mass of Mg(s)	0.00 g	0.75 g
Volume of $\text{H}_2(\text{g})$	? mL	? mL
Volume of HCl(aq)	250.0 mL	250 mL ← (250.0 mL)
Initial Concentration of HCl(aq)	1.00 mol/L	1.00 mol/L

Procedure

- Complete the following tables for the reaction under very hot and very cold conditions. Include signs in the "change in number of moles row" to indicate an increase or decrease in the number of moles.



Reaction species	Mg(s)	HCl(aq)	MgCl ₂ (aq)	H ₂ (g)
Initial number of moles	0.0412	0.250	0	0
Change in number of moles	-0.0412	-0.0823	+0.0412	+0.0412
Final number of moles	0	0.168	0.0412	0.0412



Reaction species	Mg(s)	HCl(aq)	MgCl ₂ (aq)	H ₂ (g)
Initial number of moles	0.0412	0.250	0	0
Change in number of moles	-0.0103	-0.0206	+0.0103	+0.0103
Final number of moles	0.0309	0.229	0.0103	0.0103

$$n_{\text{Mg initial}} = 1.00 \text{ g} \times \frac{1 \text{ mol}}{24.3 \text{ g}} = 0.0412 \text{ mol}$$

$$n_{\text{HCl initial}} = 0.250 \text{ L} \times \frac{1.00 \text{ mol}}{\text{L}} = 0.250 \text{ mol}$$

$$n_{\text{Mg final (cold)}} = 0.75 \text{ g} \times \frac{1 \text{ mol}}{24.3 \text{ g}} = 0.0309 \text{ mol}$$

↑
sf

Results and Discussion

- What is the limiting reagent in the reaction under both conditions?

Magnesium

- What conditions are required to complete the reaction?

Very Hot Temperature (in this time period).

- What volume of hydrogen gas (measured at STP) is formed under very hot conditions? Under very cold conditions?

HOT $0.0412 \text{ mol} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 0.923 \text{ L}$

COLD $0.0103 \text{ mol} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 0.231 \text{ L}$

- What is the percentage yield under very cold conditions?

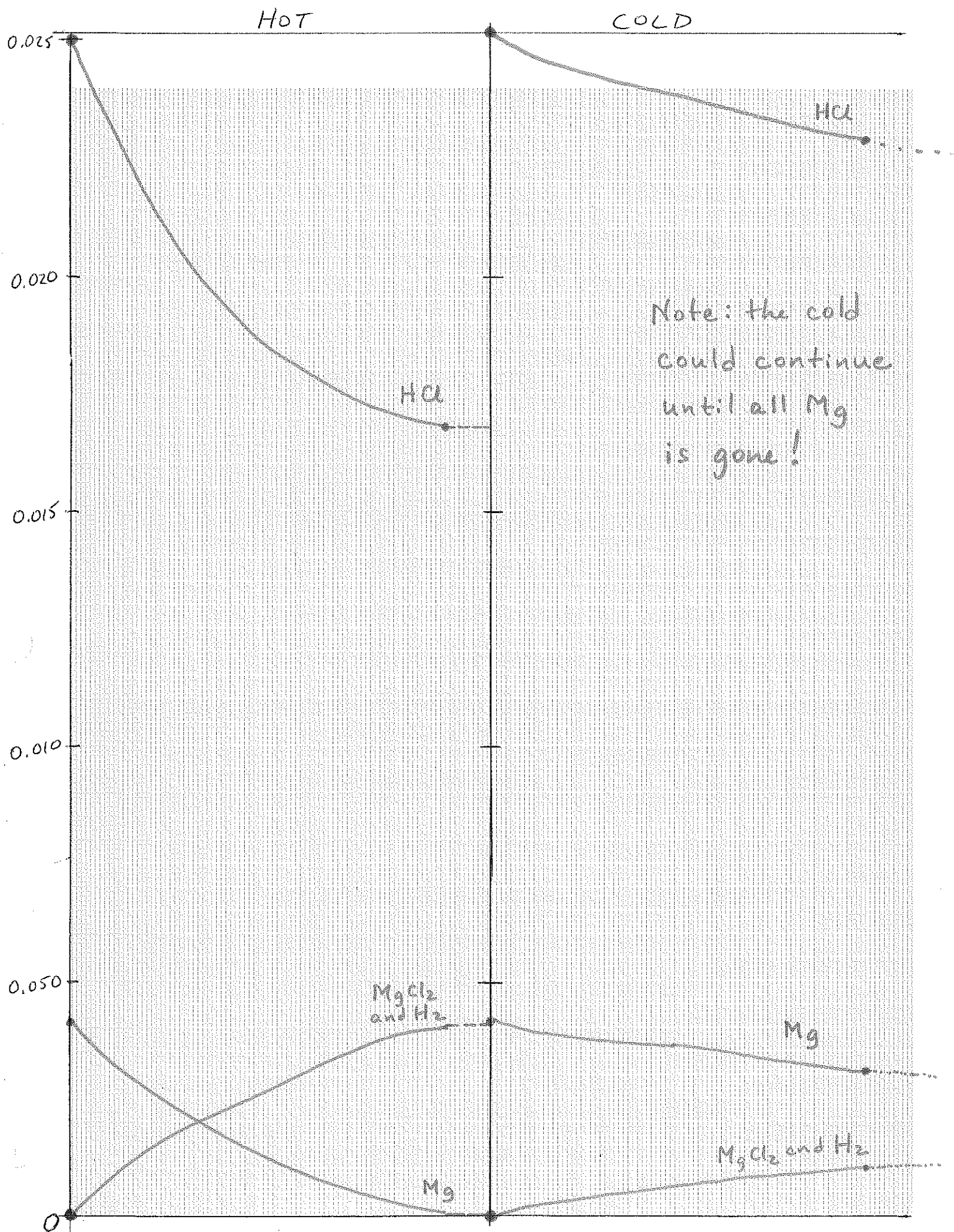
$$\frac{0.231 \text{ L}}{0.923 \text{ L}} \times 100\% = 25.0\% \quad (\text{Notice a quarter of the Mg only was used.})$$

- What is the final concentration of magnesium chloride solution under very hot conditions? Under very cold conditions?

HOT $\frac{0.0412 \text{ mol}}{0.2500 \text{ L}} = 0.165 \text{ mol/L}$

COLD $\frac{0.0103 \text{ mol}}{0.2500 \text{ L}} = 0.0412 \text{ mol/L} \quad (\frac{1}{4} [\text{MgCl}_2]_{\text{HOT}})$

- Using the graph "paper" below, sketch a graph to show the change in the number of moles of each species



4.6 Review Questions (p. 230)

1. NO
- Reactants may be impure
 - Rxn may not go 100% to completion.
 - One reactant may be in excess.

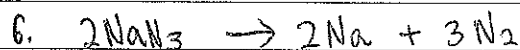
2. What are the reactants that are not completely consumed called?

Excess

3. $(\text{obtained} / \text{expected}) \times 100\%$

4. Rarely. need to apply % purity. (makes product mass smaller)

5. $\frac{0.132 \text{ g}}{100.132 \text{ g}} \times 100\% = 0.132\%$



$0.85 (\text{NaN}_3 \text{ } 120 \text{ g}) = 102 \text{ g pure NaN}_3$

$102 \text{ g NaN}_3 \times \frac{1 \text{ mol NaN}_3}{65.0 \text{ g}} \times \frac{3 \text{ mol N}_2}{2 \text{ mol NaN}_3} \times \frac{22.4 \text{ L N}_2}{1 \text{ mol N}_2} = 53 \text{ L}$

7. Limiting

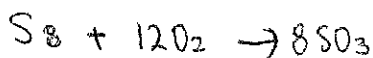
$4.22 \text{ g} \times \frac{1 \text{ mol}}{169.9 \text{ g}} \times \frac{3 \text{ mol AgCl}}{3 \text{ mol AgNO}_3} \times \frac{143.4 \text{ g}}{1 \text{ mol}} = 3.56 \text{ g}$

Excess

$7.73 \text{ g AlCl}_3 \times \frac{1 \text{ mol}}{133.5 \text{ g}} \times \frac{3 \text{ mol AgCl}}{1 \text{ mol AlCl}_3} \times \frac{143.4 \text{ g}}{1 \text{ mol}} = 24.9 \text{ g}$

8. 100% yield would produce

$\frac{8.00 \text{ mol}}{0.915} = 8.74 \text{ mol, so } 8.74 \text{ mol GeF}_3\text{H} \times \frac{1 \text{ mol GeH}_4}{4 \text{ mol GeF}_3\text{H}} = 2.19 \text{ mol}$



9. Excess

$$\hookrightarrow 5.00\text{g S}_8 \times \frac{1\text{mol S}_8}{256.8\text{g}} \times \frac{8\text{mol SO}_3}{1\text{mol S}_8} \times \frac{80.1\text{g}}{1\text{mol}} = \underline{12.5\text{g}}$$

Limits

$$\hookrightarrow 5.00\text{g O}_2 \times \frac{1\text{mol}}{32.0\text{g}} \times \frac{8\text{mol SO}_3}{12\text{mol O}_2} \times \frac{80.1\text{g}}{1\text{mol}} = \underline{8.34\text{g}}$$

10. Limiting

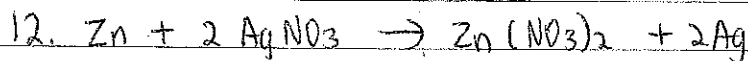
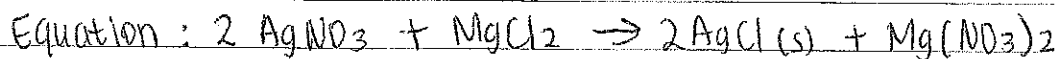
$$\hookrightarrow 40.0\text{g O}_2 \times \frac{1\text{mol O}_2}{32.0\text{g}} \times \frac{8\text{mol SO}_3}{12\text{mol O}_2} \times \frac{80.1\text{g}}{1\text{mol}} = \underline{66.8\text{g}}$$

Excess

$$\hookrightarrow 48.0\text{g S}_8 \times \frac{1\text{mol S}_8}{256.8\text{g}} \times \frac{8\text{mol SO}_3}{1\text{mol S}_8} \times \frac{80.1\text{g}}{1\text{mol}} = \underline{120.\text{g}}$$

$$\% \text{ yield} = \frac{63.2\text{g}}{66.75\text{g}} \times 100\% = \underline{94.7\%}$$

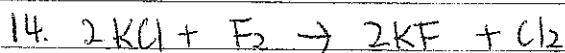
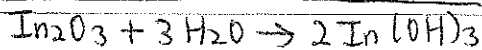
$$\hookrightarrow 8.95\text{g impure AgCl} \times \frac{75.0\text{g pure}}{100\text{g impure}} \times \frac{1\text{mol AgCl}}{143.4\text{g}} \times \frac{2\text{mol AgNO}_3}{2\text{mol AgCl}} \times \frac{1\text{L}}{0.105\text{mol}} = \underline{0.446\text{L}}$$



$$\left(145.0\text{mL} \times \frac{0.095\text{mol AgNO}_3}{1000\text{mL}} \times \frac{2\text{mol Ag}}{2\text{mol AgNO}_3} \times \frac{107.9\text{g}}{1\text{mol}} \right) \times 0.97 = \underline{1.4\text{g}}$$

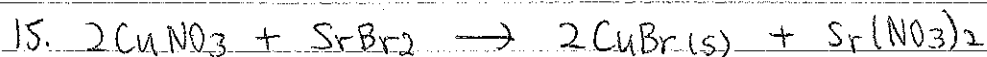
$$13. \text{Th yield} = 8.92\text{g In}_2\text{O}_3 \times \frac{1\text{mol}}{277.6\text{g}} \times \frac{2\text{mol In(OH)}_3}{1\text{mol In}_2\text{O}_3} \times \frac{165.8\text{g}}{1\text{mol}} = \underline{10.7\text{g}}$$

$$\therefore \frac{10.1\text{g}}{10.7\text{g}} \times 100\% = 94.8\%$$



(at STP)

$$39.8\text{g} \times \frac{84.0\text{g Pure}}{100\text{g impure}} \times \frac{1\text{mol KCl}}{74.6\text{g KCl}} \times \frac{1\text{mol Cl}_2}{2\text{mol KCl}} \times \frac{22.4\text{L}}{1\text{mol}} = \underline{5.02\text{L}}$$



Excess

$$46.7\text{g CuNO}_3 \times \frac{1\text{mol}}{125.5\text{g}} \times \frac{2\text{mol CuBr}}{2\text{mol CuNO}_3} \times \frac{143.4\text{g}}{1\text{mol}} = 53.4\text{g}$$

Limiting

$$30.8\text{g SrBr}_2 \times \frac{1\text{mol}}{247.4\text{g}} \times \frac{2\text{mol CuBr}}{1\text{mol SrBr}_2} \times \frac{143.4\text{g}}{1\text{mol}} = 35.7\text{g ppt form}$$

CuNO₃ in excess

used up

$$30.8\text{g SrBr}_2 \times \frac{1\text{mol}}{247.4\text{g}} \times \frac{2\text{mol CuNO}_3}{1\text{mol SrBr}_2} \times \frac{125.5\text{g}}{1\text{mol}} = 31.2\text{g CuNO}_3$$

$$\therefore 46.7 - 31.2 = 15.5\text{g left over.}$$

$$16. 20.0\text{g Ca} \times \frac{1\text{mol}}{40.1\text{g}} \times \frac{1\text{mol H}_2}{1\text{mol Ca}} \times \frac{22.4\text{L}}{1\text{mol}} = 11.2\text{L}$$

$$18.0\text{mL} \times \frac{1.00\text{g}}{1\text{mL}} \times \frac{1\text{mol}}{18.0\text{g}} \times \frac{1\text{mol H}_2}{2\text{mol H}_2\text{O}} \times \frac{22.4\text{L}}{1\text{mol}} = 11.2\text{L}$$

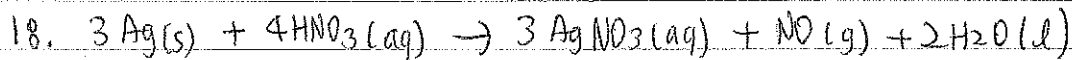
$$\text{Stoichiometric Amounts!! (ALL CONSUMED)} \therefore \frac{10.0\text{L}}{11.2\text{L}} \times 100\% = 89.5\% \text{ yield}$$



$$\text{pure reacted} = 4.41\text{g AgCl} \times \frac{1\text{mol}}{143.4\text{g}} \times \frac{1\text{mol CaCl}_2}{2\text{mol AgCl}} \times \frac{1000\text{mL}}{0.103\text{mol}} = 149\text{mL}$$

Purity was

$$\therefore \frac{149\text{mL}}{250\text{mL}} \times 100\% = 59.7 \rightarrow 60\%$$



$$0.3295\text{g AgCl} \times \frac{1\text{mol}}{143.4\text{g}} \times \frac{1\text{mol AgNO}_3}{1\text{mol AgCl}} \times \frac{3\text{mol Ag}}{3\text{mol AgNO}_3} \times \frac{107.9\text{g Ag Pure}}{1\text{mol Ag}} = 0.2479\text{g}$$

$$\text{Purity} = \frac{0.2479\text{g}}{0.7294\text{g}} \times 100\% = 33.99\%$$