Mr. Standring's Chemistry 30 Review Package 2013

Table of Contents:

Pages are in order, but many are missing. Just look for the page #'s at the bottom left or right corner.

Each section ends with a checklist of necessary knowledge, skills & connections

Solution packages are available for use in class

The Ten Study Habits of Successful Students

Successful students have good study habits. They apply these habits to all of their classes. Read about each study habit. Work to develop any study habit you do not have.

Successful students:

1. **Try not to do too much studying at one time.**
   If you try to do too much studying at one time, you will tire and your studying will not be very effective. Space the work you have to do over shorter periods of time. Taking short breaks will restore your mental energy.

2. **Plan specific times for studying.**
   Study time is any time you are doing something related to schoolwork. It can be completing assigned reading, working on a paper or project, or studying for a test. Schedule specific times throughout the week for your study time.

3. **Try to study at the same times each day.**
   Studying at the same times each day establishes a routine that becomes a regular part of your life, just like sleeping and eating. When a scheduled study time comes up during the day, you will be mentally prepared to begin studying.

4. **Set specific goals for their study times.**
   Goals will help you stay focused and monitor your progress. Simply sitting down to study has little value. You must be very clear about what you want to accomplish during your study times.

5. **Start studying when planned.**
   You may delay starting your studying because you don't like an assignment or think it is too hard. A delay in studying is called "procrastination." If you procrastinate for any reason, you will find it difficult to get everything done when you need to. You may rush to make up the time you wasted getting started, resulting in careless work and errors.

6. **Work on the assignment they find most difficult first.**
   Your most difficult assignment will require the most effort. Start with your most difficult assignment since this is when you have the most mental energy.

7. **Review their notes before beginning an assignment.**
   Reviewing your notes can help you make sure you are doing an assignment correctly. Also, your notes may include information that will help you complete an assignment.

8. **Tell their friends not to call (or text or instant message or Skype or whatever other new fandangled thing the kids use to communicate nowadays) them during their study times.**
   Two study problems can occur if your friends call you during your study times. First, your work is interrupted. It is not that easy to get back to what you were doing. Second, your friends may talk about things that will distract you from what you need to do. Here's a simple idea - turn off your cell phone during your study times.

9. **Call another student when they have difficulty with an assignment.**
   This is a case where "two heads may be better than one."

10. **Review their schoolwork over the weekend.**
    Yes, weekends should be fun time. But there is also time to do some review. This will help you be ready to go on Monday morning when another school week begins.

These ten study habits can help you throughout your education. Make sure they are your study habits.

Your notes here → ____

____

____

____

____

____

____

Student Worksheet LSM G.8

Graphic Organizers

Graphic organizers are a visual means of organizing information, note taking, showing relationships among ideas and concepts, and solidifying your understanding of a topic. Learning occurs as you make the organizers, which is why they are sometimes called “thinking tools.” There are many kinds of graphic organizers; select the style that best suits the intended purpose.

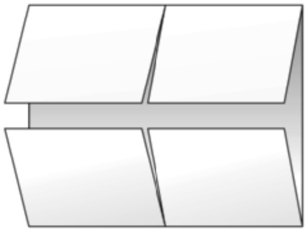
PMI Chart

A PMI chart is used to examine both sides of an issue. Positive aspects of a topic or issue are recorded in the **P** (plus) column. Negative aspects are recorded in the **M** (minus) column. Interesting or controversial questions are recorded in the **I** (interesting) column (**Table 1**).

Table 1 A PMI Chart

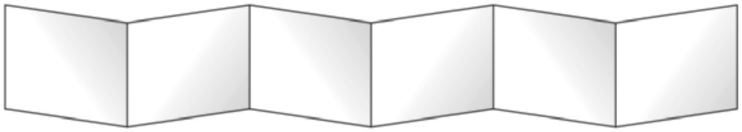
P	M	I

3-D Paper Organizers



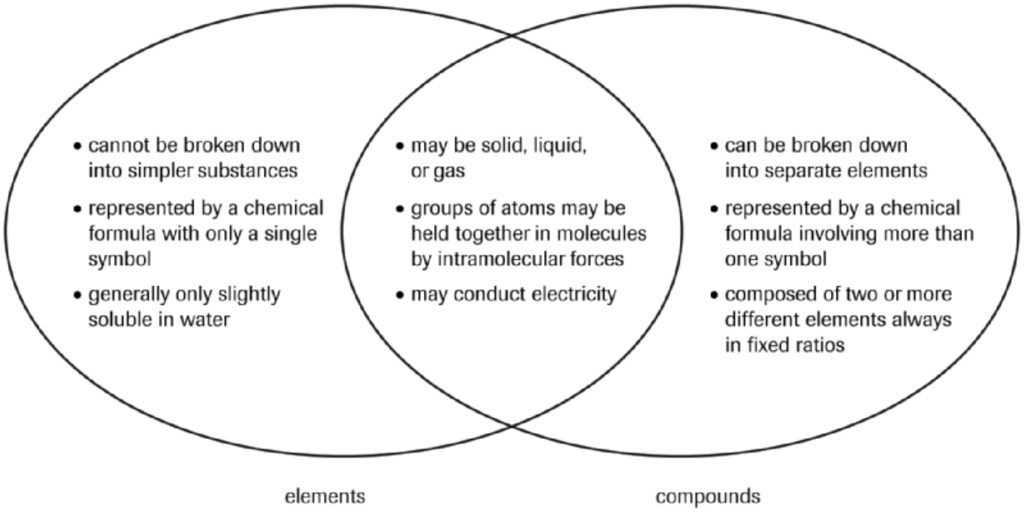
Graphic organizers that make use of cutting and folding paper are useful tools for some students. These manipulatives may be used to compare and contrast related concepts, or (in the form of flip-chart booklets) to list many items. They can also incorporate Venn diagrams. They provide a very strong organizational structure that supports student learning. All that is needed is brightly coloured sheets of strong paper, scissors, and possibly tape.

(These manipulatives have been popularized as Foldables ™ by Dinah Zike, whose books provide many creative applications.)



Venn Diagram

A Venn diagram is used to show similarities and differences in two or more concepts. Write all similarities between the concepts in the overlapping section of the circles and all unique features of each concept in the nonoverlapping parts of the appropriate circles.



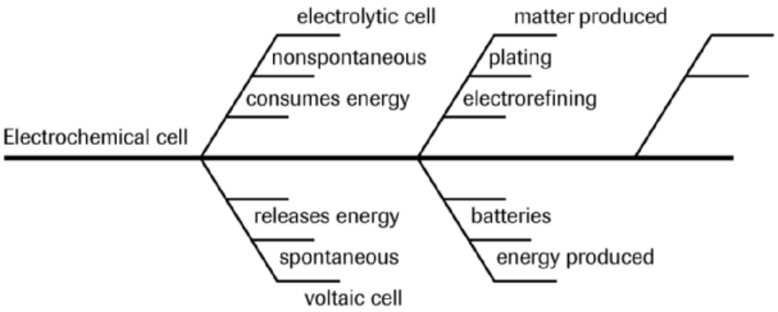
- cannot be broken down into simpler substances
- represented by a chemical formula with only a single symbol
- generally only slightly soluble in water

- may be solid, liquid, or gas
- groups of atoms may be held together in molecules by intramolecular forces
- may conduct electricity

- can be broken down into separate elements
- represented by a chemical formula involving more than one symbol
- composed of two or more different elements always in fixed ratios

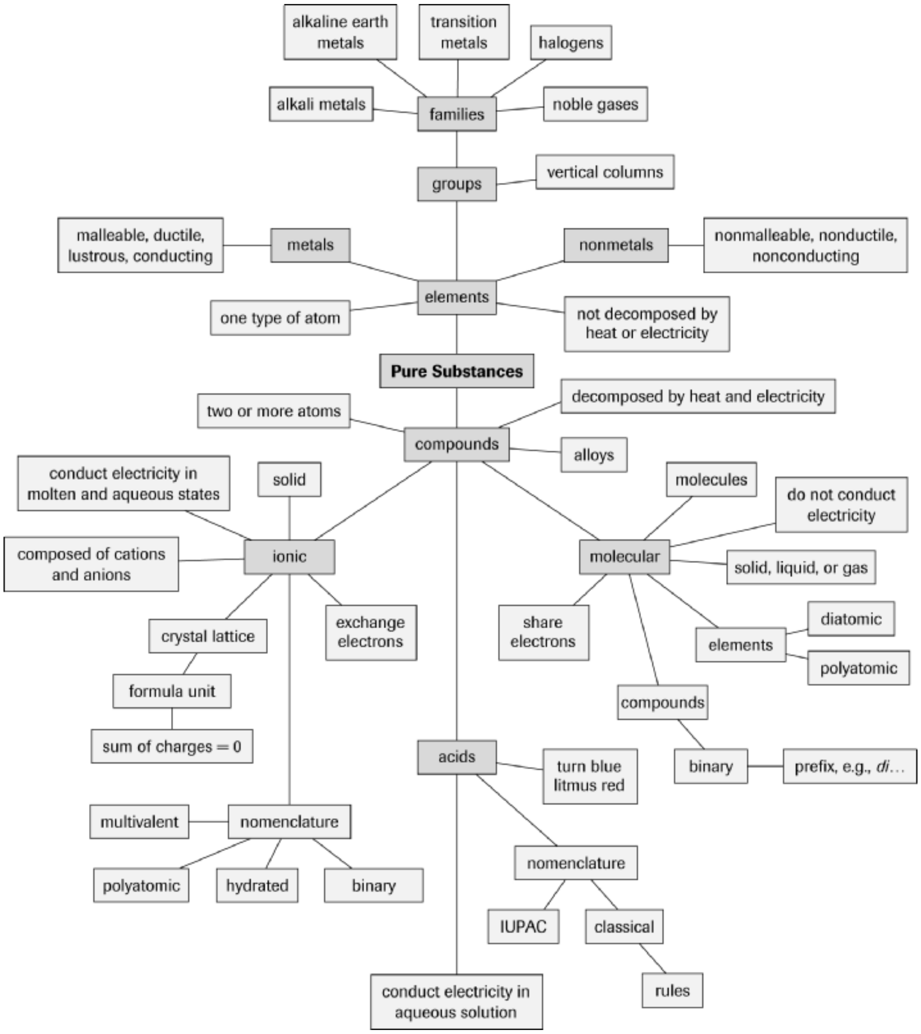
Fishbone Diagram

A fishbone diagram is used to identify separate causes and effects. In the head of the fish, identify the effect, topic, or result. At the end of each major bone, identify the major subtopics or categories. On the minor bones that attach to each major bone, add details about the subtopics or possible causes of each effect or result.



The Concept Map

Concept maps are used to show connections between ideas, using words or visuals. Put the central idea in the middle of a sheet of paper and organize the most closely related ideas around it. Draw lines between related ideas. Some concept maps include connecting words alongside the lines, to emphasize the relationship between terms.



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Student Worksheet LSM G.9

Directing Words

The following words are often used in a variety of questions that you may be asked to answer. Become familiar with these words and their specific meanings in context. This familiarity will also help to prepare you for Diploma Examination questions.

algebraically	using mathematical procedures that involve letters or symbols to represent numbers
analyze	to make a mathematical, chemical, or methodical examination of parts to determine the nature, proportion, function, interrelationship, etc. of the whole
compare	examine the character or qualities of two things by providing characteristics of both that point out their *similarities* and *differences*
conclude	state a logical end based on reasoning and/or evidence
contrast/distinguish	point out the *differences* between two things that have similar or comparable natures
criticize	point out the *demerits* of an item or issue
define	provide the essential qualities or meaning of a word or concept; make distinct and clear by marking out the limits
describe	give a written account or represent the characteristics of something by a figure, model, or picture
design/plan	construct a plan; i.e, a detailed sequence of actions for a specific purpose
determine	find a solution, to a specified degree of accuracy, to a problem by showing appropriate formulas, procedures, and calculations
enumerate	specify one by one or list in concise form and according to some order
evaluate	give the significance or worth of something by identifying the good and bad points or the advantages and disadvantages
explain	make clear what is not immediately obvious or entirely known; give the cause of or reason for; make known in detail
graphically	using a drawing that is produced electronically or by hand and that shows a relation between certain sets of numbers
how	show in what manner or way, with what meaning
hypothesize	form a tentative proposition intended as a possible explanation for an observed phenomenon; i.e., a possible cause for a specific effect. The proposition should be testable logically and/or empirically
identify	recognize and select as having the characteristics of something
illustrate	make clear by giving an example. The form of the example must be specified in the question; i.e., word description, sketch, or diagram
infer	form a generalization from sample data; arrive at a conclusion by reasoning from evidence
interpret	tell the meaning of something; present information in a new form that adds meaning to the original data

justify/show how	show reasons for or give facts that support a position
model	find a model (in mathematics, a model of a situation is a pattern that is supposed to represent or set a standard for a real situation) that does a good job of representing a situation
outline	give, in an organized fashion, the essential parts of something. The form of the outline must be specified in the question; i.e., list, flow chart, concept map
predict	tell in advance on the basis of empirical evidence and/or logic
prove	establish the truth or validity of a statement for the general case by giving factual evidence or logical argument
relate	show logical or causal connection between things
sketch	provide a drawing that represents the key features of an object or graph
solve	give a solution for a problem; i.e., explanation in words and/or numbers
summarize	give a brief account of the main points
trace	give a step-by-step description of the development
verify	establish, by substitution for a particular case or by geometric comparison, the truth of a statement
why	show the cause, reason, or purpose

This list is also available on the Alberta Education Web site, at: http://www.edc.gov.ab.ca/k_12/testing/diploma/bulletins/essential/directing_words.asp.

Solutions for Identifying and Defining Directing Words, Extra Exercise

1. (a) Evaluate: IV (Give the significance or worth of something by identifying the good and bad points or the advantages and disadvantages.)
   (b) Identify: V (Recognize and select as having the characteristics of something.)
   (c) Trace: I (Give a step-by-step description of the development.)
   (d) Infer: (Form a generalization from sample data; arrive at a conclusion by reasoning from evidence.)
   (e) Sketch: (Provide a drawing that represents the key features of an object or graph.)
   (f) Illustrate: (Make clear by giving an example. The form of the example must be specified in the question; i.e., word description, sketch, or diagram.)

2. (a) solve
   (b) explain
   (c) interpreted
   (d) relating

3. (a) Algebraically
   (b) Describe
   (c) How
   (d) Verify
   (e) Contrast/Distinguish
   (f) Determine

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Student Worksheet LSM 1.1A

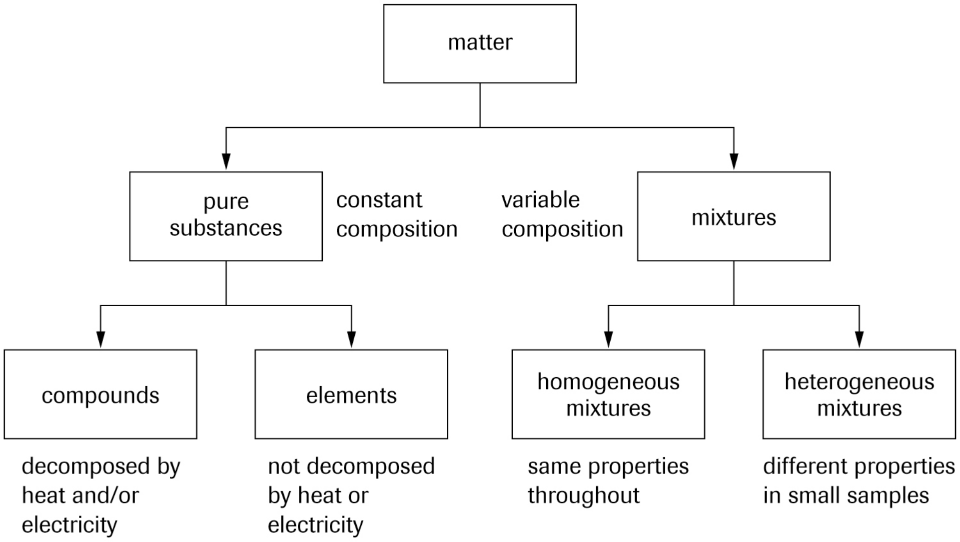
Classifying Knowledge, Extra Exercises

Complete the table below with the correct definitions.

Table 1 Definitions of Terms

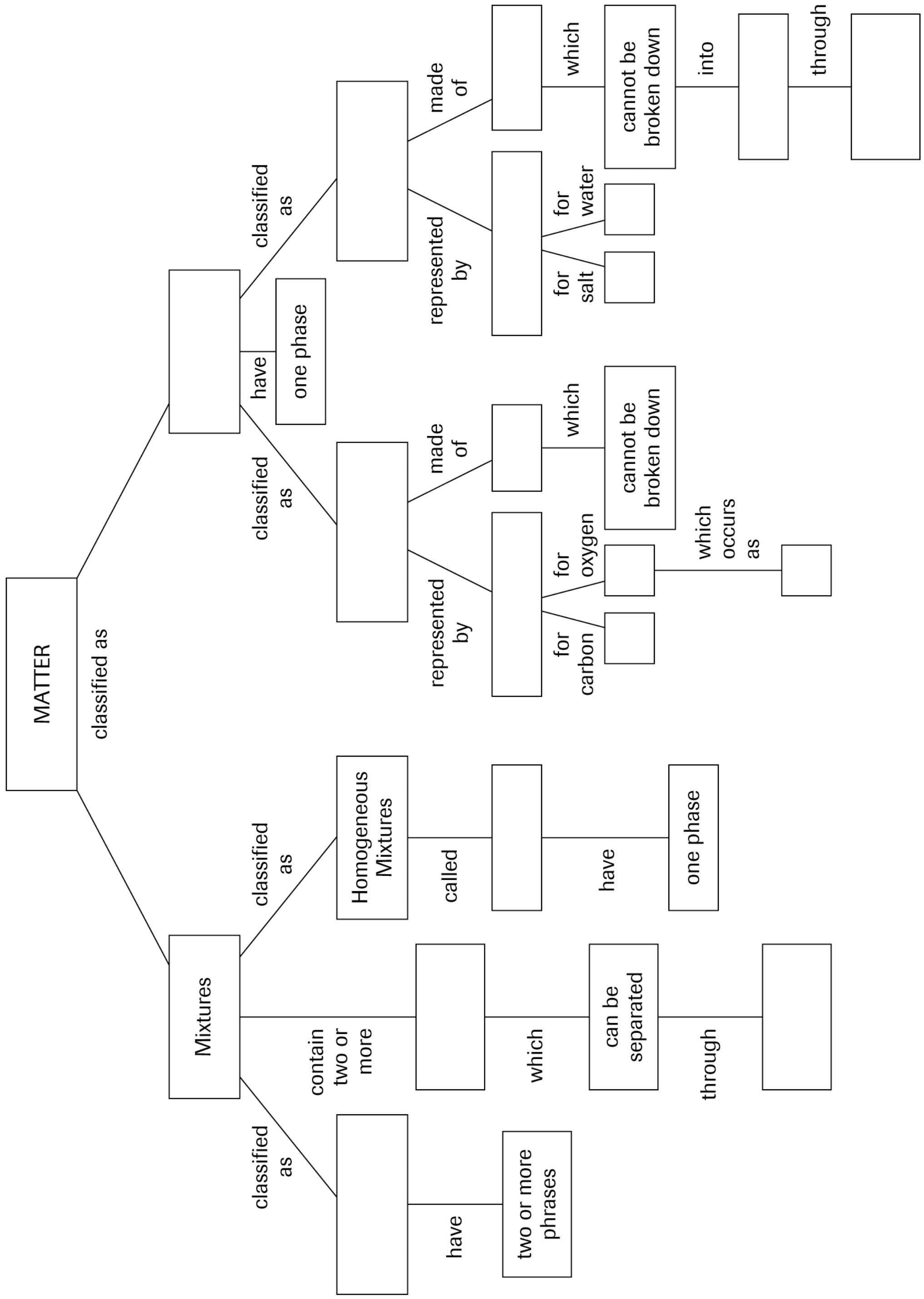
Term	Definition	Examples/Notes
observation		
qualitative observation		
quantitative observation		
interpretation		
empirical knowledge		
theoretical knowledge		
empirical hypothesis		
empirical definition		
generalization		
scientific law		

An Empirical Classification of Matter



Classification of Matter, Extra Exercises

Complete the following concept map.

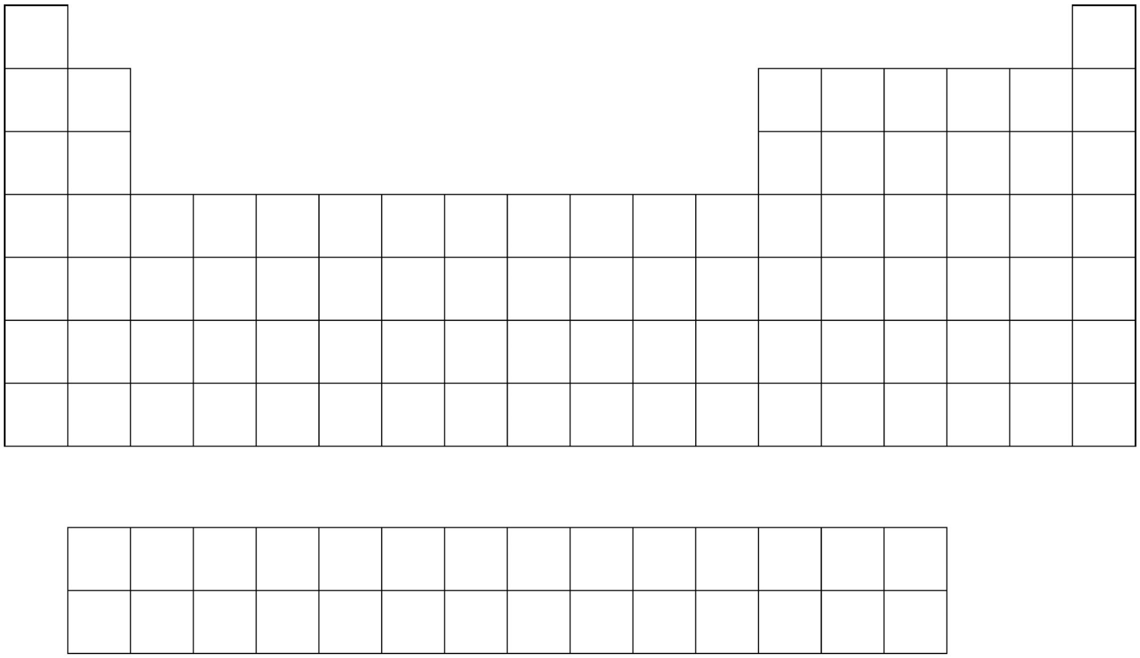


Student Worksheet LSM 1.3A

Make a Periodic Table, Extra Exercises

Follow the instructions to complete the blank periodic table below.

1. (a) Label the groups (at the top of each group) from 1 to 18 and in the Roman Numeral system (IA to VIIIA and IB to VIIIB).
   (b) Label the periods at both sides of the table.
2. (a) Lightly shade the metals with a light colour.
   (b) Lightly shade the nonmetals with another colour.
   (c) Leave the metalloids unshaded.
3. Label the transition metals, lanthanides, actinides, alkali metals, alkaline-earth metals, halogens, and noble gases.
4. Outline the “staircase” with a thick black line and label it.
5. (a) Circle the elements that are liquid at SATP.
   (b) Outline the elements with a thick black line that are gases at SATP.



Answer the following questions:

1. Define the following terms: period; group or family.
2. What does the “staircase” indicate?
3. How many elements are solid at SATP? Liquids at SATP? Gases at SATP?

Student Worksheet LSM 1.3C

Periodic Table, Extra Exercises

Use the periodic table on the inside front cover of your textbook to complete the following table

Table 1 Properties and Descriptions of Elements

Element name		Chemical symbol	Atomic number	Group number	Period number	Metal (m) or nonmetal (nm)	State at SATP	Family/Series names
1.	chlorine							
2.	magnesium							
3.			30					
4.		N						
5.				17	5			
6.			79					
7.					3			alkali metals
8.	thorium			——				
9.				12			liquid	
10.		Br						
11	argon							
12.				11	5			
13.			19					
14.	calcium							
15.				1			gas	
16.			58	——				

Student Worksheet LSM 1.4A

Bohr Diagrams and Energy Level Diagrams, Extra Exercises

Element name and symbol	Bohr diagram of atom	Energy level diagram of atom	# of valence electrons	Bohr diagram of ion	Energy level diagram of ion	Ionic charge (if an ion forms)
lithium						
beryllium						
F						
magnesium						
Al						

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Student Worksheet LSM 1.5A

Ionic Compounds, Extra Exercises

1. Write the formulas for the following compounds.
   (a) magnesium oxide ____________________
   (b) sodium fluoride ____________________
   (c) aluminium nitride ____________________
   (d) potassium sulfide ____________________
   (e) lithium iodide ____________________
   (f) calcium bromide ____________________
   (g) beryllium oxide ____________________
   (h) nickel chloride ____________________
   (i) magnesium nitride ____________________
   (j) aluminium sulfide ____________________
   (k) copper(I) bromide ____________________
   (l) tin(II) iodide ____________________
   (m) iron(III) chloride ____________________
   (n) calcium phosphide ____________________
   (o) lead(II) oxide ____________________
   (p) lead(IV) fluoride ____________________
   (q) tin(IV) bromide ____________________
   (r) copper(II) sulfide ____________________
   (s) iron(II) oxide ____________________
   (t) calcium nitride ____________________

2. Write the names for the following compounds.
   (a) Li_2O ____________________
   (b) $AlCl_3$ ____________________
   (c) MgS ____________________
   (d) CaO ____________________
   (e) KBr ____________________
   (f) BeF ____________________
   (g) Na_3N ____________________
   (h) Al_2O_3 ____________________
   (i) $CuCl_2$ ____________________
   (j) $FeBr_3$ ____________________
   (k) PbS ____________________
   (l) SnO_2 ____________________
   (m) Na_2S ____________________
   (n) Mg_3P_2 ____________________
   (o) NiO ____________________
   (p) CuI ____________________
   (q) $PbCl_4$ ____________________
   (r) FeP ____________________
   (s) CaF_2 ____________________
   (t) K_3P ____________________

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Molecular Compounds, Extra Exercises

1. Write the formulas for the following compounds.

   (a) carbon dioxide ____________________

   (b) silicon dioxide ____________________

   (c) water ____________________

   (d) carbon disulfide ____________________

   (e) sulfur trioxide ____________________

   (f) ammonia ____________________

   (g) carbon tetrachloride ____________________

   (h) hydrogen peroxide ____________________

   (i) methane ____________________

   (j) ozone ____________________

   (k) diphosphorus trioxide ____________________

   (l) nitrogen monoxide ____________________

   (m) chlorine dioxide ____________________

   (n) dinitrogen oxide ____________________

   (o) carbon monoxide ____________________

   (p) arsenic tribromide ____________________

   (q) phosphorus pentabromide ____________________

   (r) dinitrogen tetroxide ____________________

   (s) silicon carbide ____________________

   (t) sulfur dioxide ____________________

2. Write the names for the following compounds.

   (a) CF_4 ____________________

   (b) NH_3 ____________________

   (c) PBr_3 ____________________

   (d) O_3 ____________________

   (e) $F_2(g)$ ____________________

   (f) CS_2 ____________________

   (g) N_2O_4 ____________________

   (h) H_2O_2 ____________________

   (i) CO ____________________

   (j) SiC ____________________

   (k) P_2O_5 ____________________

   (l) CH_4 ____________________

   (m) SO_3 ____________________

   (n) H_2O ____________________

   (o) SiO_2 ____________________

   (p) PCl_5 ____________________

   (q) $I_2(g)$ ____________________

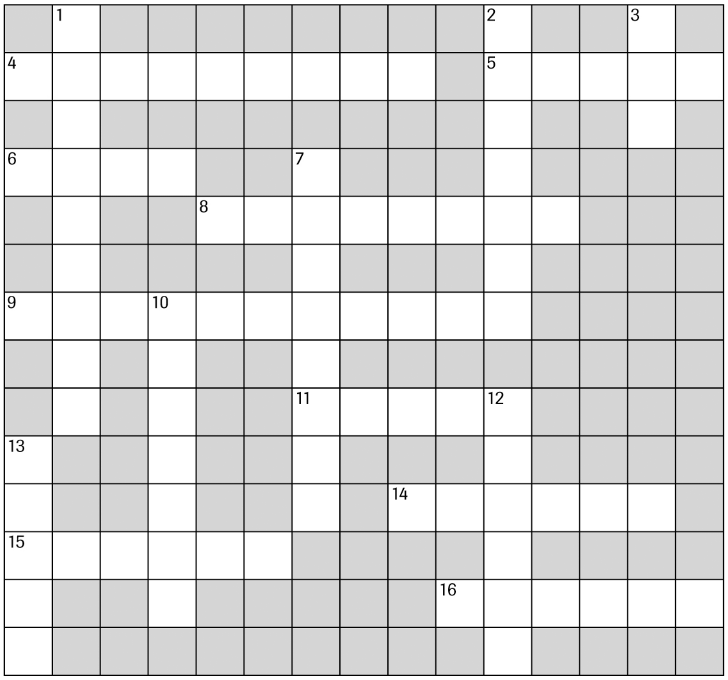
   (r) NO_2 ____________________

   (s) SF_4 ____________________

   (t) $H_2(g)$ ____________________

Student Worksheet LSM 1.6C

Crossword on Elements, Compounds, and the Periodic Table, Extra Exercises



Across

3. __________ is a vital ion in the human body.
5. An example of an element that has a stable electron arrangement is __________.
6. The overall charge on an ionic compound is __________.
8. The Group 17 elements are called __________.
9. A(n) __________ is a solution that conducts electricity.
11. Compounds that are made up of positive and negative ions are called __________.
14. __________ are shiny, malleable conductors of electricity.
15. Chemical families or __________ are elements located in the same vertical column.
16. __________ is an example of an alkali metal.

Down

1. __________ compounds are formed when nonmetals combine with other nonmetals.
2. Electrons in the outer shell of an atom are called __________ electrons.
3. A charged atom is called a(n) __________.
7. When a fluorine atom gains an electron, it forms a(n) __________ ion.
10. __________ is an example of an alkaline-earth metal.
12. A(n) __________ is a positively charged ion.
13. __________ is an example of a nonelectrolyte when it is dissolved in water.

Student Worksheet LSM 1.CS

Chapter 1 Checklist

Have you mastered the concepts, applications, and skills associated with the following items? Check them off when you are confident in your understanding.

Knowledge

classify matter as pure and mixtures as homogeneous and heterogeneous (1.2)	
interpret the periodic table of the elements (1.3)	
use atomic theory to explain the periodic table (1.4)	
classify elements and compounds, and know the properties of each class (1.3, 1.4)	
explain and predict chemical formulas for and name ionic and molecular compounds, acids, and bases (1.5, 1.6)	
identify the state of matter of substances (1.5, 1.6)	
write chemical equations when given reactants and products (1.5, 1.6)	
classify scientific knowledge as qualitative and quantitative, as observations and interpretations, and as empirical and theoretical (1.1)	

STS

describe the natures of science and technology (1.1)	
describe the application of some common chemicals (1.3, 1.5, 1.6)	

Skills

use a textbook, a periodic table, and other references efficiently and effectively (1.1–1.6)	
interpret and write laboratory reports (1.1, 1.2, 1.3, 1.4, 1.6)	
select and use diagnostic tests (1.2, 1.3, 1.4, 1.5, 1.6)	

Key Terms

1.1	science	
	technology	
	chemistry	
	observation	
	interpretation	
	empirical knowledge	
	theoretical knowledge	
	empirical hypothesis	
	empirical definition	
	generalization	
	scientific law	
	law of conservation of mass	
1.2	matter	
	pure substance	
	mixture	
	heterogeneous mixture	
	homogeneous mixture	
	element	
	entity	
	atom	
	compound	
	chemical formula	
1.3	periodic law	
	family	
	group	
	period	
	semi-metal	
	standard ambient temperature and pressure (SATP)	
	metal	
	nonmetal	
	alkali metal	
	alkaline-earth metal	
	halogen	
	noble gas	
	main group element	
	transition element	

LSM 1.CS (cont'd)

1.4	theoretical hypothesis	
	theoretical definition	
	theory	
	mass number	
	atomic number	
	ion	
	monatomic ion	
	cation	
	anion	
1.5	ionic compound	
	molecular compound	
	acid	
	base	
	neutral	
	aqueous solution	
	polyatomic ion	
	formula unit	
	empirical formula	
	hydrate	
1.6	molecule	
	molecular formula	
	diatomic molecule	

Solutions for Writing Chemical Formulas, Extra Exercises

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1. (a) Ca = 1 atom of calcium
   (b) N_2 = 2 atoms of nitrogen in the molecule
   (c) $Ba_3(PO_4)_2$ = 1 molecule of barium phosphate, which includes
   3 atoms of barium
   2 atoms of phosphorus
   8 atoms of oxygen
   Total number of atoms = 13
   (d) 2 H_2O = 2 molecules of water, which includes
   4 atoms of hydrogen
   2 atoms of oxygen
   Total number of atoms = 6
   (e) 3 $FeSO_4$ = 3 molecules of iron(II) sulfate, which includes
   3 atoms of iron
   3 atoms of sulfur
   12 atoms of oxygen
   Total number of atoms = 18
   (f) 4 $Cu(NO_3)_2$ = 4 molecules of copper(II) nitrate, which includes
   4 atoms of copper
   8 atoms of nitrogen
   24 atoms of oxygen
   Total number of atoms = 36
2. (a) K_2CrO_4

Type of atom	**# of atoms**
potassium	**2**
chromium	**1**
oxygen	**4**
Total	**7**

(b) 2 $(NH_4)_2Cr_2O_7$

Type of atom	**# of atoms**
nitrogen	**4**
hydrogen	**16**
chromium	**4**
oxygen	**14**
Total	**38**

Instructional Master LSM 2.3A

Summary: Balancing Chemical Equations

Step 1: Write the chemical formula for each reactant and product, including the state of matter for each one.

Step 2: Try balancing the atom or ion present in the greatest number. Find the lowest common multiple to obtain coefficients to balance this particular atom or ion.

Step 3: Repeat step 2 to balance each of the remaining atoms and ions.

Step 4: Check the final reaction equation to ensure that all atoms and ions are balanced.

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Student Worksheet LSM 2.4A

Molar Mass and Conversions, Extra Exercises

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1. Determine the molar mass of each of the following substances:
   (a) $MgI_2(s)$ (b) $Al(OH)_3(s)$

   (c) $(NH_4)_2CO_3(s)$ (d) $CoCl_2 \bullet 6H_2O(s)$

2. Convert each of the following masses into its chemical amount:
   (a) 8.40 g of $NaOH(s)$

   (b) 4.2 kg of $H_2O(l)$

3. Convert each of the following amounts into a mass in grams of the given substance:
   (a) 0.456 mol of $Al_2(SO_4)_3(s)$

   (b) 0.518 mmol of $CuSO_4 \bullet 5H_2O(s)$

4. Complete the following table.

Table 1 Molar Calculations

Substance	Molar mass (g/mol)	Mass (g)	Chemical amount (mol)
$CaCl_2(s)$		18.6	
$Al_2O_3(s)$			0.267
$Mg(OH)_2(s)$		35.00	
$Na_2CO_3 \bullet 10H_2O(s)$			0.150

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Instructional Master LSM 2.5A

Chemical Reactions

Table 1 Chemical Reactions

Reaction type	Generalization
formation	elements → compound
simple decomposition	compound → elements
complete combustion	substance + oxygen → most common oxides
single replacement	element + compound → element + compound (metal + compound → metal + compound nonmetal + compound → nonmetal + compound)
double replacement	compound + compound → compound + compound

Solubility of Ionic Compounds

SOLUBILITY OF IONIC COMPOUNDS AT SATP—GENERALIZATIONS*

Ion	**Cl^- Br^- I^-**	**S^{2-}**	**OH^-**	**SO_4^{2-}**	**CO_3^{2-} PO_4^{3-} SO_3^{2-}**	**CH_3COO^-**	**NO_3^- ClO_3^- ClO_4^-**	**Group 1 NH_4^+ H_3O^+ (H^+)**
Very soluble (aq) $\geq$ 0.1 mol/L	most	Group 1, NH_4^+, Group 2	Group 1, NH_4^+, Sr^{2+}, Ba^{2+}, Tl^+	most	Group 1, NH_4^+	most	all	all
Slightly soluble (s) $<$ 0.1 mol/L (at SATP)	Ag^+, Pb^{2+}, Tl^+, Hg_2^{2+}, Cu^+	most	most	Ag^+, Pb^{2+}, Ca^{2+}, Ba^{2+}, Sr^{2+}, Ra^{2+}	most	Ag^+, Hg_2^{2+}	none	none

*Although these are particularly reliable, all generalizations have exceptions. This textbook specifically identifies any reference to an ionic compound solubility that is an exception to these generalizations.

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Student Worksheet LSM 2.6B

Classifying Chemical Reactions, Extra Exercises

1. (a) Classify each of the following reactions as formation, simple decomposition, single replacement, or double replacement reactions.
   (b) Balance each equation and add symbols to indicate states of matter for all reactants and products.

   (i) $Cu + O_2 \rightarrow CuO$ ______________

   (ii) $Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe$ ______________

   (iii) $Ag + S \rightarrow Ag_2S$ ______________

   (iv) H_2O + electricity $\rightarrow H_2 + O_2$ ______________

   (v) $FeS + HCl \rightarrow FeCl_2 + H_2S$ ______________

   (vi) $NaCl \rightarrow Na + Cl_2$ ______________

   (vii) $NaOH + HCl \rightarrow H_2O + NaCl$ ______________

   (viii) $Zn + HCl \rightarrow ZnCl_2 + H_2$ ______________

2. Write balanced chemical equations for the following:
   (a) The decomposition reaction of hydrogen sulfide.

   (b) The single displacement reaction of copper metal and silver nitrate.

   (c) The synthesis reaction of sodium and fluorine.

   (d) The double displacement reaction of aluminium sulfate and calcium hydroxide.

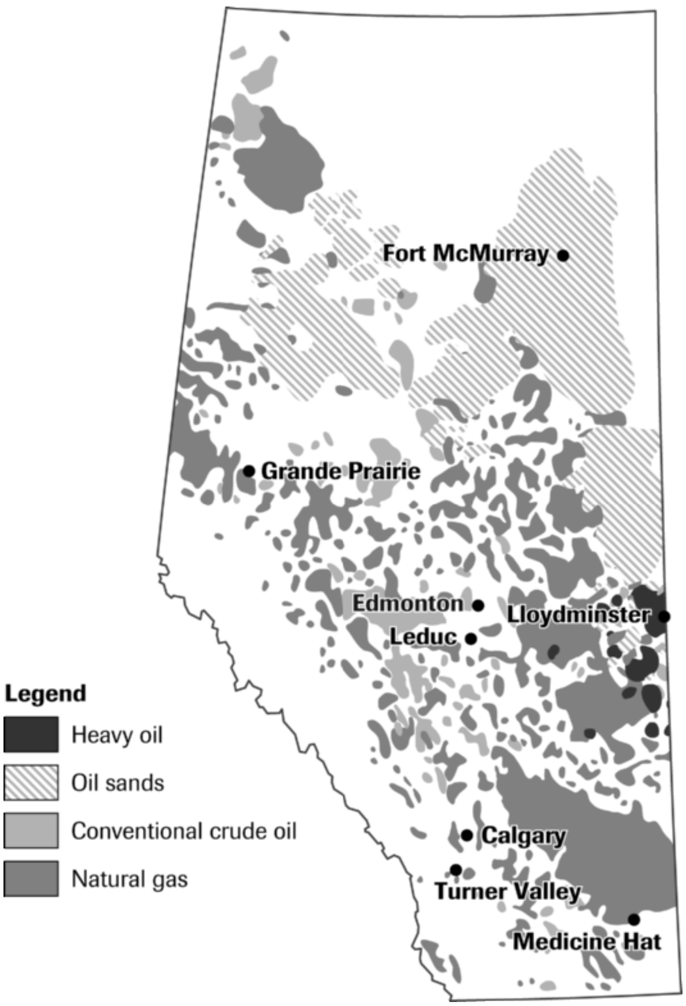
Student Worksheet LSM 2.6D

Predicting Chemical Reactions, Extra Exercises

For each of the following questions, classify the reaction type (formation, simple decomposition, combustion, single replacement, double replacement, or other), and predict the balanced chemical equation. Provide a word equation as well.

1. $Al(s) + O_2(g) \rightarrow$
2. $Ag_2O(s) \rightarrow$
3. $Br_2(l) + KI(aq) \rightarrow$
4. A strip of zinc metal is placed into a copper(II) nitrate solution.
5. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow$
6. Sulfuric acid is neutralized by aqueous sodium hydroxide.
7. $\rightarrow CuS(s) + NaCH_3COO(aq)$
8. $CuS(s) + O_2(g) \rightarrow$
9. Propane burns in air.
10. $Na_2CO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$

Fossil Fuel Industries in Alberta



Examples of Hydrocarbons, Alkanes, and Related Alkyl Groups

Table 1 Examples of Hydrocarbons

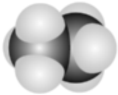
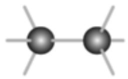
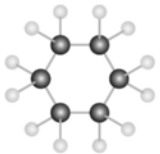
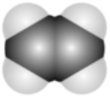
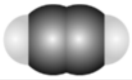
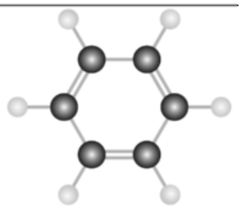
Hydrocarbon group	Example	Formula	Spacefill diagram	Bond and angles diagram
Aliphatic				
alkane	ethane	CH_3CH_3		
	cyclohexane	C_6H_{12}		
alkene	ethene	CH_2CH_2		
alkyne	ethyne	CHCH		
Aromatic				
	benzene	C_6H_6		

Table 3 The Alkane Family of Organic Compounds

IUPAC Name	Formula
methane	$CH_4(g)$
ethane	$C_2H_6(g)$
propane	$C_3H_8(g)$
butane	$C_4H_{10}(g)$
pentane	$C_5H_{12}(l)$
hexane	$C_6H_{14}(l)$
heptane	$C_7H_{16}(l)$
octane	$C_8H_{18}(l)$
nonane	$C_9H_{20}(l)$
decane	$C_{10}H_{22}(l)$
–ane	C_nH_{2n+2}

Table 4 Examples of Alkyl Branches

Branch	Name
$-CH_3$	methyl
$-C_2H_5$ $(-CH_2CH_3)$	ethyl
$-C_3H_7$ $(-CH_2CH_2CH_3)$	propyl

Naming and Drawing Branched Alkanes

SUMMARY Naming Branched Alkanes

Step 1: Identify the longest continuous chain of carbon atoms—the parent chain—in the structural formula. Number the carbon atoms, starting from the end closest to the branch(es), so that the numbers are the lowest possible.

Step 2: Identify any branches and their location number on the parent chain.

Step 3: Write the complete IUPAC name, following this format:
(number of location) – (branch name)(parent chain).

SUMMARY Drawing Branched Alkane Structural Formulas

Step 1: Draw a straight chain containing the number of carbon atoms represented by the name of the parent chain, and number the carbon atoms from left to right.

Step 2: Attach all branches to their numbered locations on the parent chain.

Step 3: Add enough hydrogen atoms to show that each carbon has four single bonds.

Learning Tip

Be careful when choosing the longest continuous chain (parent chain) because the structure may twist and turn. If your name ends up with an alkyl branch on the number 1 position, then you have made a mistake in identifying the parent chain. For example, 1,1-dimethylpropane is incorrect and should be 2-methylbutane.

Learning Tip

When naming hydrocarbons, not only must the branches be put in alphabetical order but also the set of numbers must be the lowest set. The lowest set of numbers is defined by IUPAC as being the set that starts with the lowest number; e.g., 2, 4 and 5 are lower than 3, 3 and 4. Start numbering the longest continuous chain from the end of the molecule with the closest branch.

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Student Worksheet LSM 9.2C

Naming Aliphatic and Aromatic Hydrocarbons, Extra Exercises

---

1. For each of the following IUPAC names, draw a structural diagram.

(a) methylpentane

(b) octane

(c) 2,3,4-trimethylpentane

(d) ethene

(e) propyne

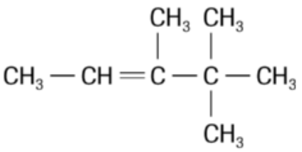
(f) methylpropene

(g) cyclohexane

2. For each of the following structural diagrams, write the IUPAC name.

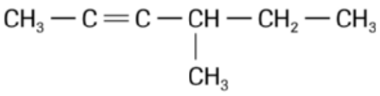
(a) $CH_3-CH_2-CH_2-CH=CH_2$

(b)



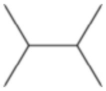
$CH_3-CH=C(CH_3)-C(CH_3)_2-CH_3$

(c)



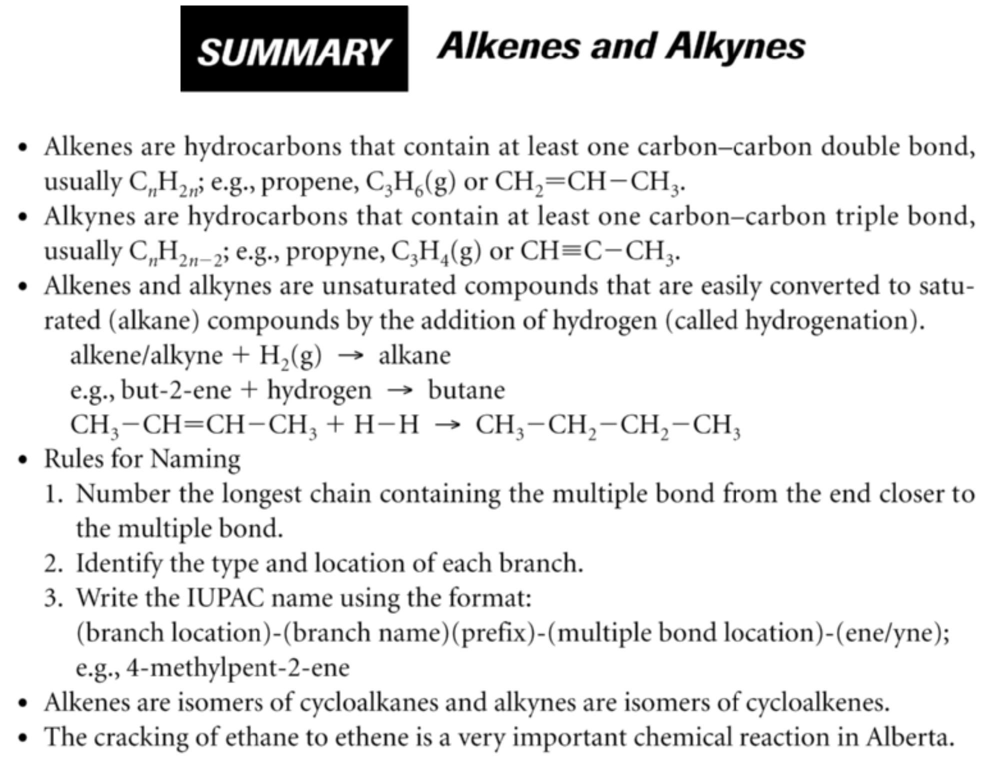
$CH_3-C=C-CH(CH_3)-CH_2-CH_3$

(d)



Instructional Master LSM 9.3B

Alkenes and Alkynes



SUMMARY Alkenes and Alkynes

- Alkenes are hydrocarbons that contain at least one carbon–carbon double bond, usually C_nH_{2n}; e.g., propene, $C_3H_6(g)$ or $CH_2{=}CH{-}CH_3$.
- Alkynes are hydrocarbons that contain at least one carbon–carbon triple bond, usually C_nH_{2n-2}; e.g., propyne, $C_3H_4(g)$ or $CH{\equiv}C{-}CH_3$.
- Alkenes and alkynes are unsaturated compounds that are easily converted to saturated (alkane) compounds by the addition of hydrogen (called hydrogenation).
  alkene/alkyne + $H_2(g)$ → alkane
  e.g., but-2-ene + hydrogen → butane
  $CH_3{-}CH{=}CH{-}CH_3 + H{-}H \rightarrow CH_3{-}CH_2{-}CH_2{-}CH_3$
- Rules for Naming
  1. Number the longest chain containing the multiple bond from the end closer to the multiple bond.
  2. Identify the type and location of each branch.
  3. Write the IUPAC name using the format:
     (branch location)-(branch name)(prefix)-(multiple bond location)-(ene/yne); e.g., 4-methylpent-2-ene
- Alkenes are isomers of cycloalkanes and alkynes are isomers of cycloalkenes.
- The cracking of ethane to ethene is a very important chemical reaction in Alberta.

Naming Aromatic Hydrocarbons

SUMMARY Naming Aromatic Hydrocarbons

1. If an alkyl branch is attached to a benzene ring, the compound is named as an alkylbenzene. Alternatively, the benzene ring may be considered as a branch of a large molecule; in this case, the benzene ring is called a phenyl branch.
2. If more than one alkyl branch is attached to a benzene ring, the branches are numbered using the lowest numbers possible, starting with one of the branches. Given a choice between two sets of lowest numbers, choose the set that is in both numerical and alphabetical order. See Sample Problem 9.3.

Learning Tip

You may encounter a naming system—known as the classical system—for benzene rings with two substituted groups. The 1,2-, 1,3-, and 1,4-arrangements are denoted by the prefixes ortho- (*o*), meta- (*m*), and para- (*p*), respectively. These names, such as *o*-dimethylbenzene, are still used in industry, and you may encounter them in other references.

Fractional Distillation

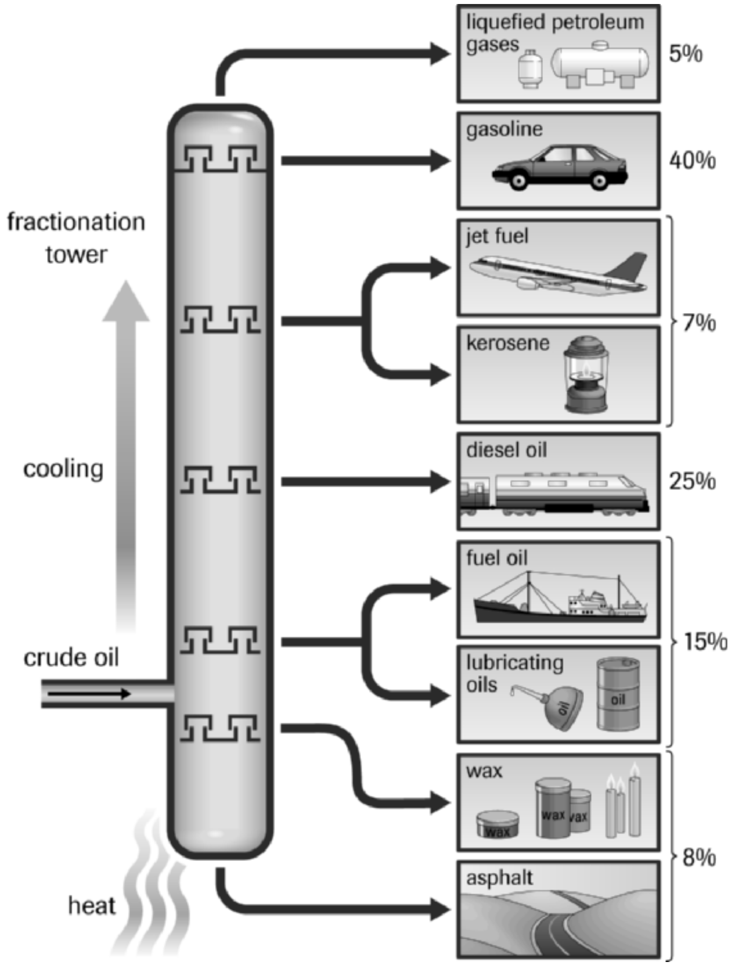


Figure 2
A fractional distillation (fractionation) tower contains trays positioned at various levels. Heated crude oil enters near the bottom of the tower. The bottom of the tower is kept hot, and the temperature gradually decreases toward the top of the tower. The concentration of components with lower boiling points increases from the bottom to the top of the tower.

The percentage distributions shown vary with the type of crude oil and with seasonal demands.

Crude Oil Refining

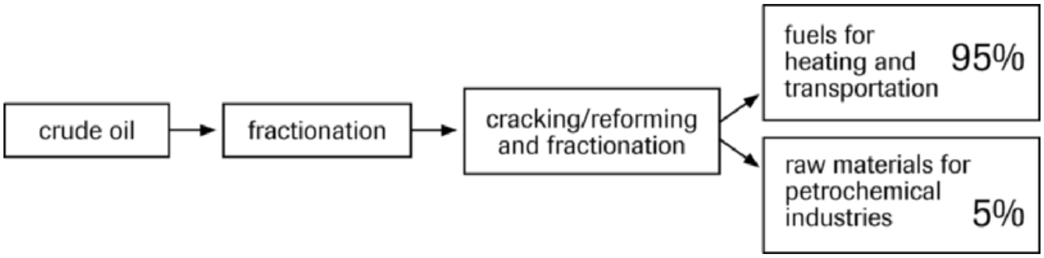
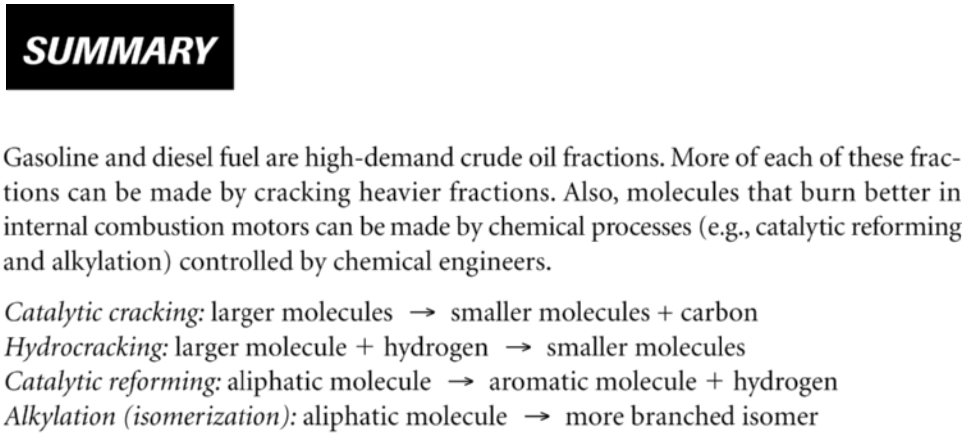


Figure 4
Only 5% of the original mass of crude oil is used as starting chemicals (called feedstock) in the manufacture of solvents, greases, plastics, synthetic fibres, and pharmaceuticals.

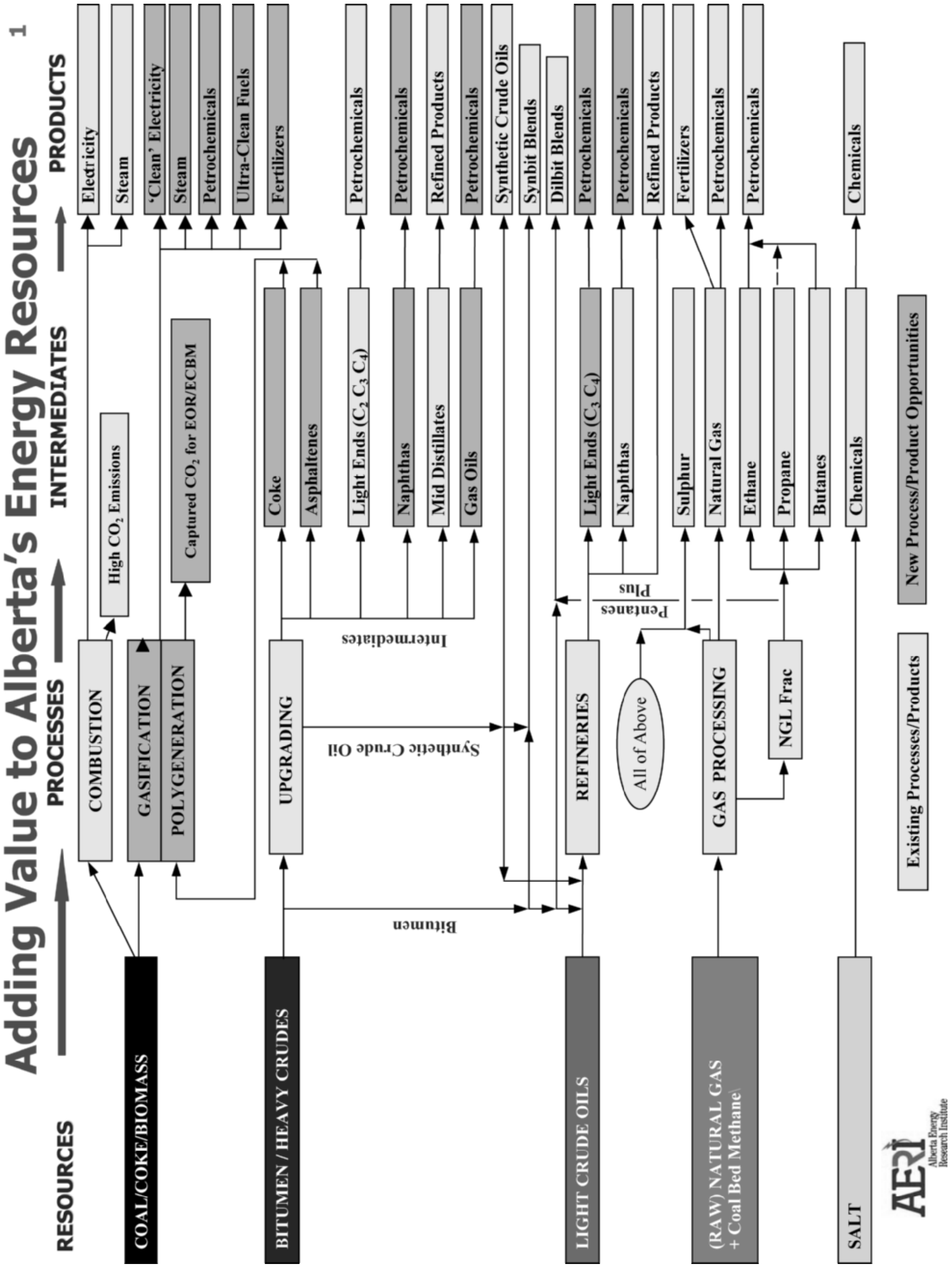


SUMMARY

Gasoline and diesel fuel are high-demand crude oil fractions. More of each of these fractions can be made by cracking heavier fractions. Also, molecules that burn better in internal combustion motors can be made by chemical processes (e.g., catalytic reforming and alkylation) controlled by chemical engineers.

Catalytic cracking: larger molecules → smaller molecules + carbon
Hydrocracking: larger molecule + hydrogen → smaller molecules
Catalytic reforming: aliphatic molecule → aromatic molecule + hydrogen
Alkylation (isomerization): aliphatic molecule → more branched isomer

Adding Value to Alberta’s Energy Resources



Instructional Master LSM 9.6B

Complete and Incomplete Combustion

SUMMARY ***Complete and Incomplete Combustion***

Complete Combustion

hydrocarbon + (excess) $O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Incomplete Combustion

hydrocarbon + (insufficient) $O_2(g) \rightarrow x\text{C}(s) + y\text{CO}(g) + z\text{CO}_2(g) + H_2O(g)$

(The ratio of x:y:z largely depends on the proportion of oxygen available.)

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Student Worksheet LSM 9.CS

Chapter Checklist

Have you mastered the concepts, applications, and skills associated with the following items? Check them off when you are confident in your understanding.

Knowledge

define organic compounds, recognizing inorganic exceptions (9.1)	
identify and describe organic compounds in everyday life, as well as their origins and applications (all sections)	
name and draw structures for saturated and unsaturated aliphatic and aromatic hydrocarbons (9.2 to 9.6)	
classify organic compounds from their structural formulas (9.2 to 9.5)	
define and use the concept of structural isomerism and relate to properties of isomers (9.2 to 9.4)	
compare boiling points and solubility of organic compounds (9.2 to 9.4)	
describe fractional distillation and solvent extraction (9.1, 9.2, 9.3, 9.5)	
define, provide examples of, predict products, and write and interpret balanced equations for combustion reactions (9.1, 9.2, 9.5, 9.6)	
describe major reactions for producing energy and economically important compounds from fossil fuels (all sections)	

STS

illustrate how science and technology are developed to meet societal needs and expand human capabilities (all sections)	
describe interactions of science, technology, and society (all sections)	
illustrate how science and technology have both intended and unintended consequences (9.1, 9.2, 9.5, 9.6)	

Skills

initiating and planning	describe procedures for safe handling, storing, and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information (9.5)	
performing and recording	separate a mixture of organic compounds based on boiling point differences (9.5)	
	build molecular models depicting the structures of selected organic and inorganic compounds (9.3)	
analyzing and interpreting	follow IUPAC guidelines for naming and formulas and by compiling evidence to compare the properties of structural isomers (all sections)	
	compile and organize data to compare the properties of structural isomers (9.3)	
	investigate sources of greenhouse gases, that is, methane, carbon dioxide, water, and dinitrogen oxide (nitrous oxide) and the issue of climate change (all sections)	
communication and teamwork	work cooperatively in addressing problems and apply skills and conventions of science in communicating information and ideas and in assessing results by preparing reports on topics related to organic chemistry (all sections)	

Key Terms

9.1	organic chemistry	
	hydrocarbon	
	refining	
9.2	alkane	
	homologous series	
	saturated hydrocarbon	
	structural isomer	
	alkyl branch	
	cycloalkane	
9.3	unsaturated compound	
	hydrogenation	
	alkene	
	alkyne	
	cycloalkene	
	cracking	
	aliphatic hydrocarbon	
9.4	aromatic	
	phenyl group	
9.5	fractionation	
	fraction	
	solvent extraction	
	hydrocracking	
	catalytic reforming	

Key STS

appreciate that Alberta is a source of a variety of fossil fuels *(CS: Fossil Fuel Industries in Alberta; WA: CS—Coal in Alberta; Section 9.2; EI: Coalbed Methane; DYK: Ethane Extraction Plants; Section 9.3 Questions 14, 15; CS: The Athabasca Oil Sands; Chapter 9 Review Questions 39, 41)*	
understand the importance of hydrocarbons as fuels and to produce useful products *(Exploration: Burning Fossil Fuels; Section 9.1; WA: CS—Coal in Alberta; Section 9.2 Questions 1, 4, 9, 10, 11; Section 9.4 Question 4; Section 9.5; CC: The Petroleum Industry; CS: Octane Number; Section 9.5 Questions 9, 10, 13; Section 9.6)*	
recognize the physical and chemical processes involved in refining natural gas, crude oil, and oil sands *(Section 9.1; Section 9.2; EI: Coalbed Methane; Section 9.3; Section 9.5; Investigation 9.3; WA: CA—Karl Chuang; CS: Octane Number; Investigation 9.4; Investigation 9.5; CS: The Athabasca Oil Sands)*	
compare hydrocracking and catalytic reforming *(Section 9.5)*	
understand the role of science and technology in improving the quality of gasoline by improving the octane rating, reducing pollutants, and adding oxygenators *(Section 9.2 Table 2; Section 9.5; CS: Octane Number; Section 9.5 Questions 9, 13; Section 9.6; DYK: Ethanol in Gasoline; Section 9.6 Question 2)*	
consider the appropriate use and environmental effects of fossil fuel extraction *(Exploration: Burning Fossil Fuels; CS: Fossil Fuel Industries in Alberta; WA: CS—Coal in Alberta; EI: Coalbed Methane; CS: The Athabasca Oil Sands; Section 9.5 Question 10; Chapter 9 Review questions 38, 39, 41, 44)*	
explore various careers associated with the petroleum industry *(CC: Petroleum Engineer; CC: The Petroleum Industry)*	

Key Skills

distinguish types of organic compounds from given structural formulas *(all sections)*	
separate hydrocarbons by distillation *(Investigation 9.3)*	
compare properties of organic compounds *(Investigations 9.1; Section 9.2 Question 10; Investigation 9.2; Lab Exercise 9.A; Lab Exercise 9.B; Investigation 9.3)*	
use IUPAC organic nomenclature *(all sections)*	

Petrochemicals and Their Derivatives

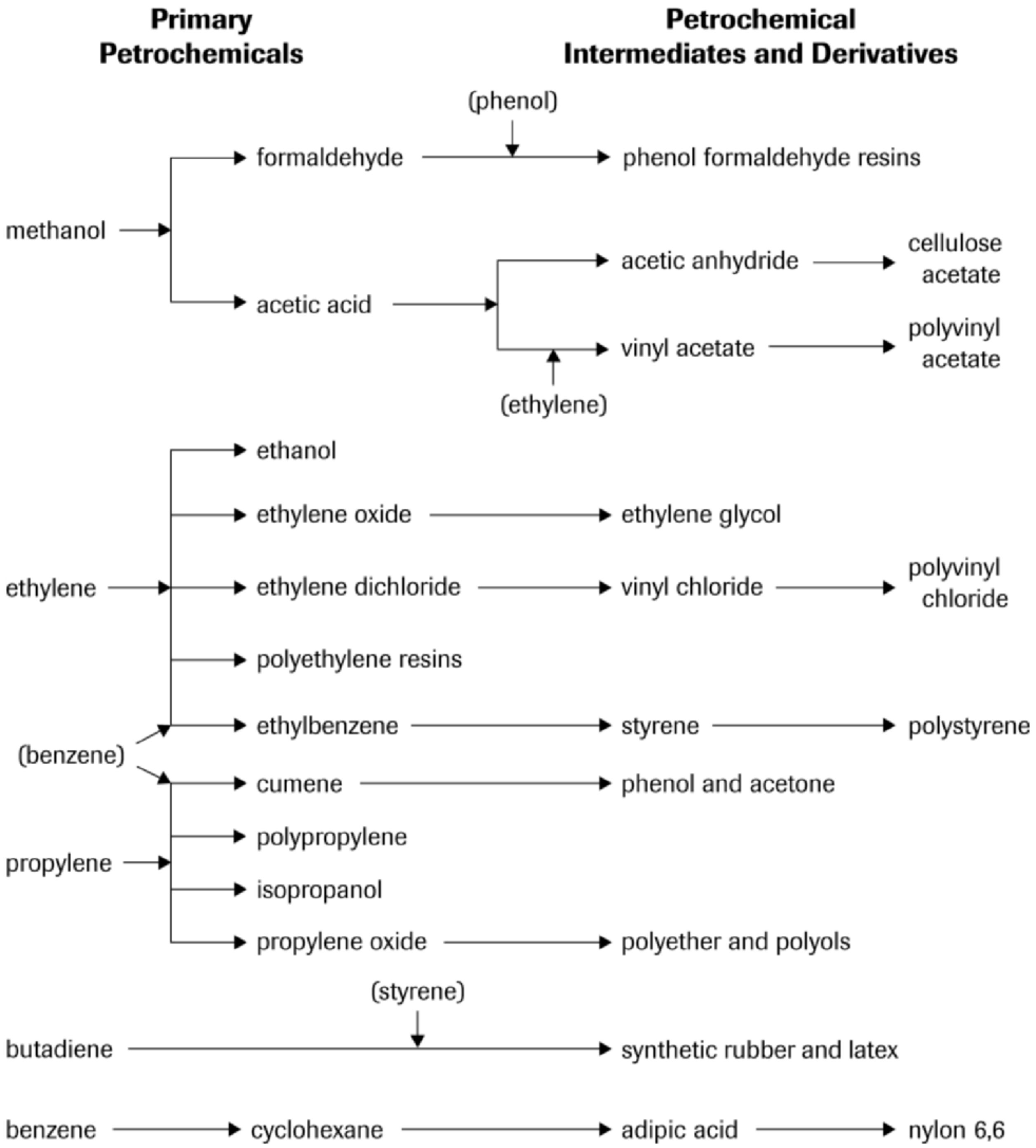


Figure 2
This figure, provided by industry, communicates primary petrochemicals, intermediates, and derivatives. You do not have to memorize this figure, or the common industrial names (except acetic acid, ethylene, and propylene).

The Ethylene Industry

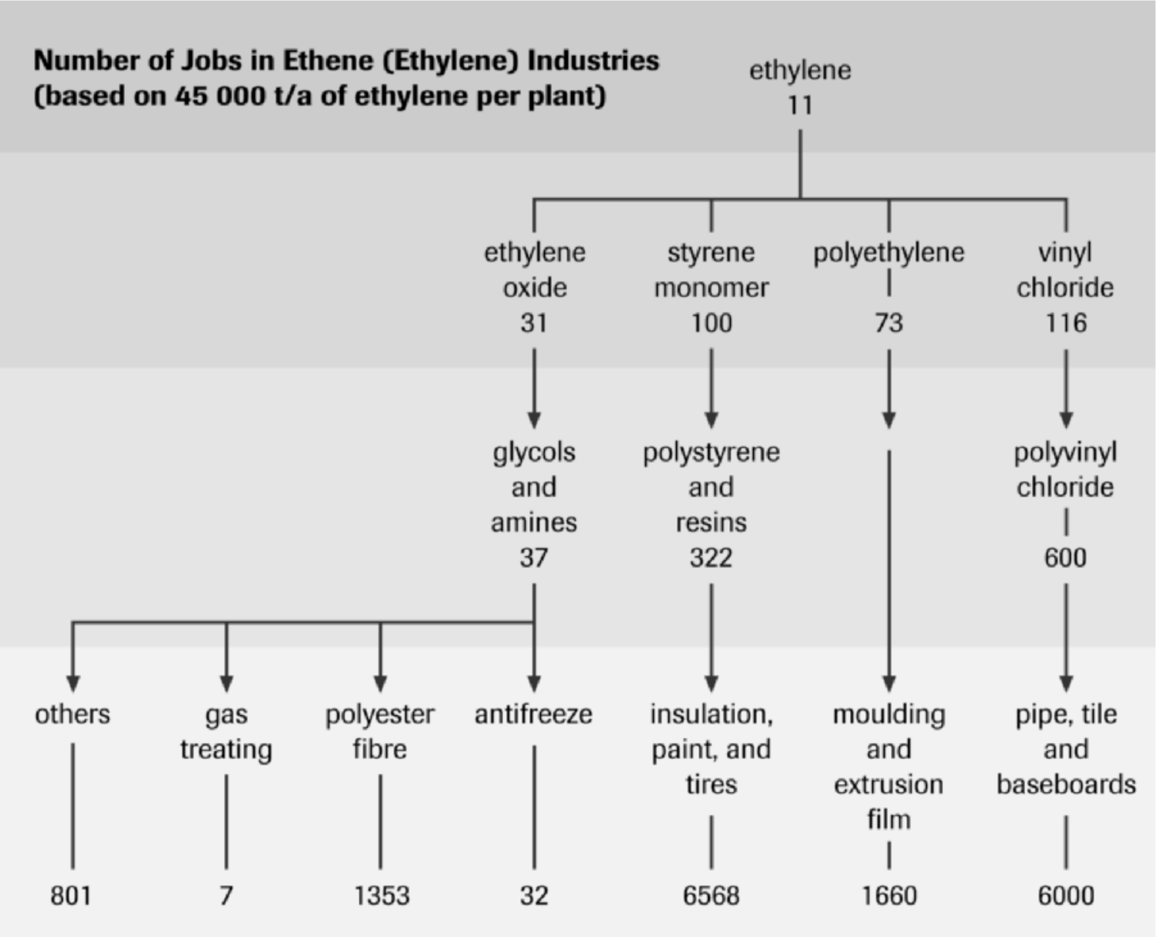


Figure 3
Number of jobs in ethene industries, based on 45 000 t/a of ethene per plant. This is called a multiplier effect. For every 11 jobs in the ethene plant, there are thousands of jobs created elsewhere.

Student Worksheet LSM 10.2A

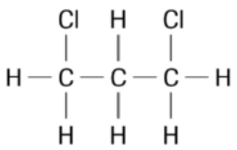
Identification and Naming of Organic Halides, Extra Exercises

---

1. In the following questions, the IUPAC names of a variety of organic compounds are provided. Draw a structural diagram for each compound.
    (a) 2-bromopentane
    (b) 1,4-dichlorobenzene
    (c) 1,1-dichloro-2,2-difluoroethane
    (d) 1,3,5-trichlorohexane
    (e) 1-fluorobut-2-ene
    (f) 1-chloro-3-fluorocyclopentane

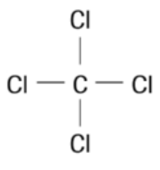
2. Give the IUPAC name for the following organic compound represented by the given structural formula.

(a)



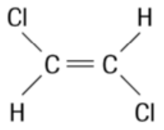
$$\begin{array}{ccccccc} & & Cl & & H & & Cl & \\ & & | & & | & & | & \\ H & - & C & - & C & - & C & - & H \\ & & | & & | & & | & \\ & & H & & H & & H & \end{array}$$

(b)

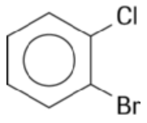


$$\begin{array}{ccccc} & & Cl & & \\ & & | & & \\ Cl & - & C & - & Cl \\ & & | & & \\ & & Cl & & \end{array}$$

(c)



(d)



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Student Worksheet LSM 10.2C

Reactions of Hydrocarbons, Extra Exercises

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1. Write an equation for each of the following reactions. Use molecular or structural formulas and classify the reaction as combustion, addition, substitution, hydrogenation, alogenation, hydration, or hydrohalogenation.

   (a) Octane burns with oxygen gas.

   (b) 2-methylpent-1-ene reacts with hydrogen.

   (c) Ethyne and excess oxygen react.

   (d) 3-methylbut-1-yne reacts with excess hydrogen.

   (e) Chlorine reacts with ethane.

   (f) Water and ethane are placed in concentrated sulfuric acid with a catalyst.

   (g) Bromine combines with benzene.

   (h) Propene and HBr combine.

   (i) Cyclohexene reacts with hydrogen gas.

   (j) Acetylene reacts with excess hydrogen.

Organic Halides

SUMMARY Organic Halides

Properties

Organic halides

- may be polar or nonpolar molecules or may have a relatively nonpolar (hydrocarbon) end and a polar (halide) end
- have higher boiling points than similar hydrocarbons
- have very low solubility in water but higher solubility (especially for small molecules) than similar hydrocarbons
- are typically good solvents for organic materials such as fats, oils, waxes, gums, resins, and/or rubber

Preparation

- Addition reactions with halogens or hydrogen halides
  alkenes and alkynes → organic halides
  e.g., $CH_2{=}CH_2 + Cl_2 \rightarrow CH_2Cl{-}CH_2Cl$
- Substitution reactions with halogens
  alkanes and aromatics $\xrightarrow{\text{light}}$ organic halides
  e.g., $CH_3{-}CH_2Cl + Cl_2 \xrightarrow{\text{light}} CH_2Cl{-}CH_2Cl + H{-}Cl$

Student Worksheet LSM 10.2F

Organic Reactions, Extra Exercises

For each of the following questions, state the organic reaction type, draw condensed structural diagrams for all reactants and products, and name all products. (You do not need to balance the equations.)

1. Propane reacts with fluorine

2. Chloroethane reacts with hydroxide ions

3. Ethanol, CH_3CH_2OH, burns in an automobile engine.

4. Chloromethane is produced by reacting methane and chlorine.

5. Bromine and ethene react to form an organic halide.

6. Hydrogen chloride and ethene react to produce an organic halide.

Student Worksheet LSM 10.2H

Applications of Organic Reactions, Extra Exercises

1. Research examples of the use of organic chemistry to address health, safety, or environmental problems, and write a report or present one such case study. Examples of topics include leaded and unleaded gasoline, solvents in dry cleaning, aerosol propellants, and pesticides and fertilizers.

   www.science.nelson.com GO

2. Why was mustard gas such an effective weapon, both during the First World War and more recently? Research its properties and effects, and defences developed against it.

   www.science.nelson.com GO

3. Long-term replacements for CFCs and HCFCs include hydrofluorocarbons (HFCs), which have no chlorine atoms. HFCs are now being used in refrigerators, automobile air conditioners, aerosol cans, and as foaming agents.
   (a) Research the latest information on HFCs and on alternative chemicals or technologies that are being used in refrigerators and air conditioners. Present your findings in a brief report.
   (b) Name, and draw structural formulas for two chlorine-free organic halides that could potentially be used in place of CFCs and HCFCs.

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Instructional Master LSM 10.3A

Alcohols and Elimination Reactions

SUMMARY Alcohols and Elimination Reactions

Alcohols

Functional group:

–OH, hydroxyl group

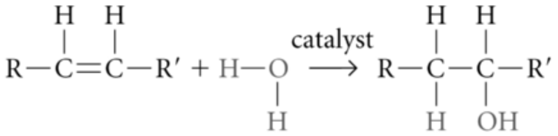
Naming alcohols:

- Drop the "e" from the alkane name and add "ol"; e.g., ethane becomes ethanol.
- If necessary, add a number (or numbers) to communicate where the hydroxyl group(s) is (are) located; e.g., propan-1-ol and propan-2-ol.
- If the alcohol has two or three hydroxyl groups, it is a diol or a triol, respectively; e.g., ethane-1,2-diol and propane-1,2,3-triol. For diols and triols, do not drop the "e" from the alkane name.

Preparation:

Addition reactions with water

- alkenes + water → alcohols

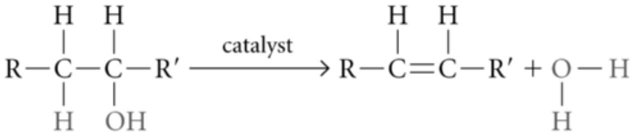


$$\mathrm{R-\overset{H}{C}=\overset{H}{C}-R' + H-\underset{H}{O} \xrightarrow{\text{catalyst}} R-\underset{H}{\overset{H}{C}}-\underset{OH}{\overset{H}{C}}-R'}$$

(R and R′ can be the same or different alkyl groups. Isomers often result.)

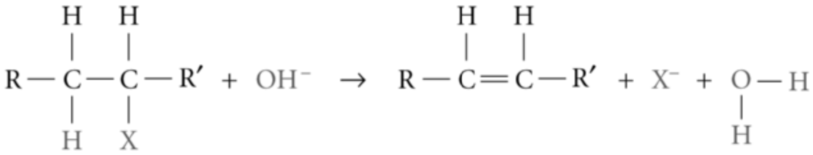
Elimination Reactions

- alcohols → alkenes + water



$$\mathrm{R-\underset{H}{\overset{H}{C}}-\underset{OH}{\overset{H}{C}}-R' \xrightarrow{\text{catalyst}} R-\overset{H}{C}=\overset{H}{C}-R' + \underset{H}{O}-H}$$

- organic halides + OH⁻ → alkenes + halide ion + water



$$\mathrm{R-\underset{H}{\overset{H}{C}}-\underset{X}{\overset{H}{C}}-R' + OH^- \rightarrow R-\overset{H}{C}=\overset{H}{C}-R' + X^- + \underset{H}{O}-H}$$

Student Worksheet LSM 10.3C

Alcohols and Elimination Reactions, Extra Exercises

---

1. Draw a structural diagram for each compound named.

(a) butan-1-ol

(b) 2-methylpropan-2-ol

(c) 1,2-dimethylbutan-2-ol

(d) pentan-2-ol

(e) cyclohexane-1,4-diol

2. For each of the following questions, draw condensed structural diagrams for all reactants and products, and name all products. (You do not need to balance the equations.)
   (a) Butan-1-ol reacts in the presence of concentrated sulfuric acid.

   (b) A 5-carbon alkene is formed when a secondary alcohol undergoes an elimination reaction.

   (c) In the presence of a strong base, 2-chlorobutane forms releases water and another small entity as it is transformed to an unsaturated compound.

3. Considering the technological difficulties and the economic cost of producing ethene, which of the following solutions would you choose? Justify your selection:
   - burn and/or export the natural gas without removing the ethane
   - remove the ethane and export it separately in a special pipeline
   - crack the ethane to ethene and export it by pipeline
   - use the ethene in Alberta to produce tertiary products, such as plastics

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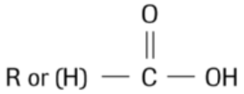
Instructional Master LSM 10.4A

Condensation Reactions of Carboxylic Acids to Form Esters

Learning Tip

The general formula for a carboxylic acid is

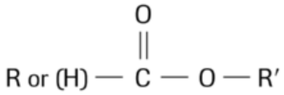


$$\text{R or (H)}-\overset{\overset{\displaystyle O}{\|}}{C}-OH$$

The functional group for carboxylic acids is the carboxyl group, –COOH.

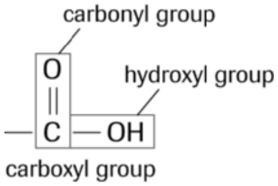
Learning Tip

The general formula for an ester is



$$\text{R or (H)}-\overset{\overset{\displaystyle O}{\|}}{C}-O-R'$$

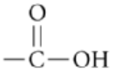
where R′ is the branch from the alcohol and R (or H) is from the acid.



SUMMARY Carboxylic Acids, Esters, and Esterification

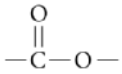
Functional groups

- carboxylic acid: –COOH carboxyl group



$$-\overset{\overset{\displaystyle O}{\|}}{C}-OH$$

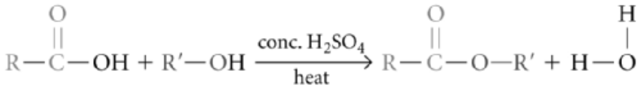
- ester: –COO– alkylated carboxyl group



$$-\overset{\overset{\displaystyle O}{\|}}{C}-O-$$

Esterification reaction

carboxylic acid + alcohol → ester + water

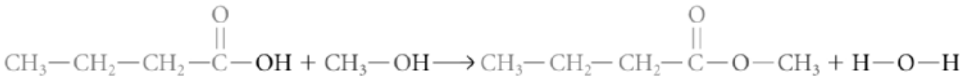


$$R-\overset{\overset{\displaystyle O}{\|}}{C}-OH + R'-OH \xrightarrow[\text{heat}]{\text{conc. } H_2SO_4} R-\overset{\overset{\displaystyle O}{\|}}{C}-O-R' + H-\overset{\overset{\displaystyle H}{|}}{O}$$

carboxylic acid alcohol ester water

(For esterification, R can be replaced by an H, but R' cannot be replaced by an H.)

e.g.,

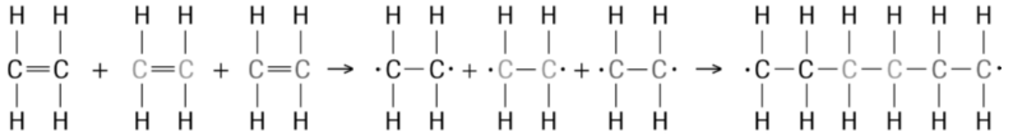


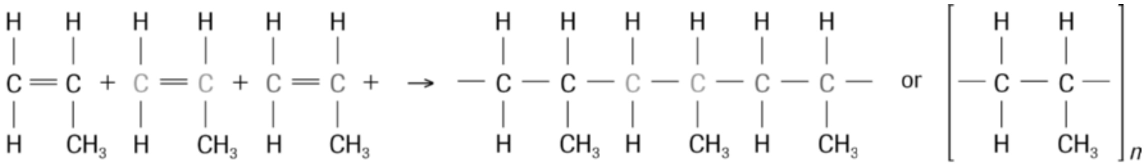
$$CH_3-CH_2-CH_2-\overset{\overset{\displaystyle O}{\|}}{C}-OH + CH_3-OH \longrightarrow CH_3-CH_2-CH_2-\overset{\overset{\displaystyle O}{\|}}{C}-O-CH_3 + H-O-H$$

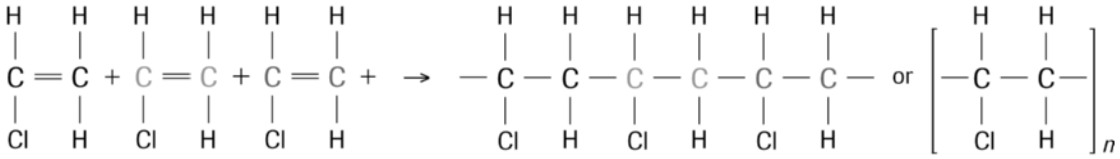
butanoic acid + methanol ⟶ methyl butanoate + water

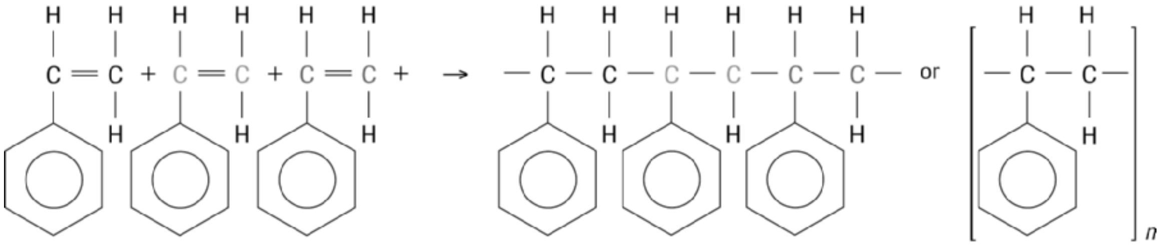
Instructional Master LSM 10.5A

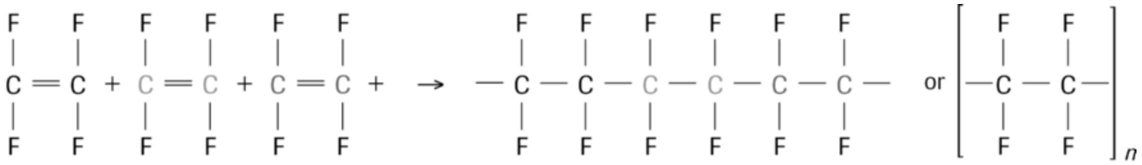
Polymer Addition Reactions





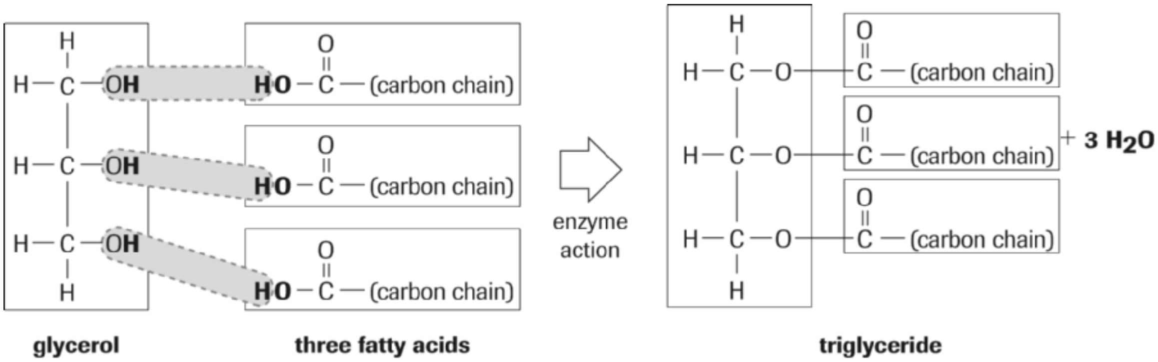






Instructional Master LSM 10.5B

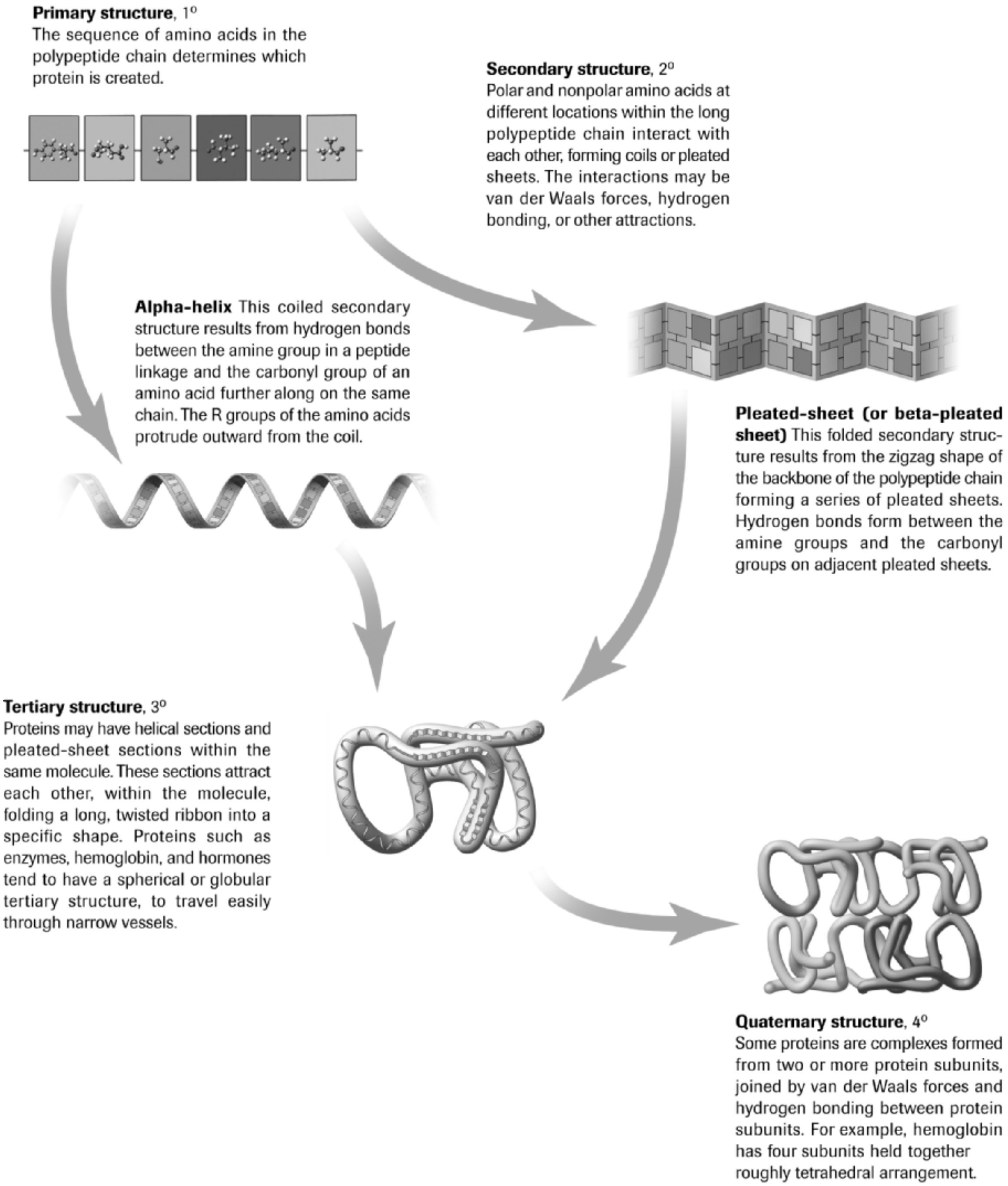
Fats and Oils



Triglycerides are formed by the union of glycerol and three fatty acids. Note the removal of water in the synthesis. The terms monoglyceride and diglyceride are used to describe the joining of glycerol with one or two fatty acids, respectively.

Instructional Master LSM 10.5D

Four Levels of Proteins



Primary structure, 1°
The sequence of amino acids in the polypeptide chain determines which protein is created.

Secondary structure, 2°
Polar and nonpolar amino acids at different locations within the long polypeptide chain interact with each other, forming coils or pleated sheets. The interactions may be van der Waals forces, hydrogen bonding, or other attractions.

Alpha-helix This coiled secondary structure results from hydrogen bonds between the amine group in a peptide linkage and the carbonyl group of an amino acid further along on the same chain. The R groups of the amino acids protrude outward from the coil.

Pleated-sheet (or beta-pleated sheet) This folded secondary structure results from the zigzag shape of the backbone of the polypeptide chain forming a series of pleated sheets. Hydrogen bonds form between the amine groups and the carbonyl groups on adjacent pleated sheets.

Tertiary structure, 3°
Proteins may have helical sections and pleated-sheet sections within the same molecule. These sections attract each other, within the molecule, folding a long, twisted ribbon into a specific shape. Proteins such as enzymes, hemoglobin, and hormones tend to have a spherical or globular tertiary structure, to travel easily through narrow vessels.

Quaternary structure, 4°
Some proteins are complexes formed from two or more protein subunits, joined by van der Waals forces and hydrogen bonding between protein subunits. For example, hemoglobin has four subunits held together roughly tetrahedral arrangement.

Proteins are very complicated compounds, with up to four levels of organization giving each protein a unique physical shape with unique physical characteristics.

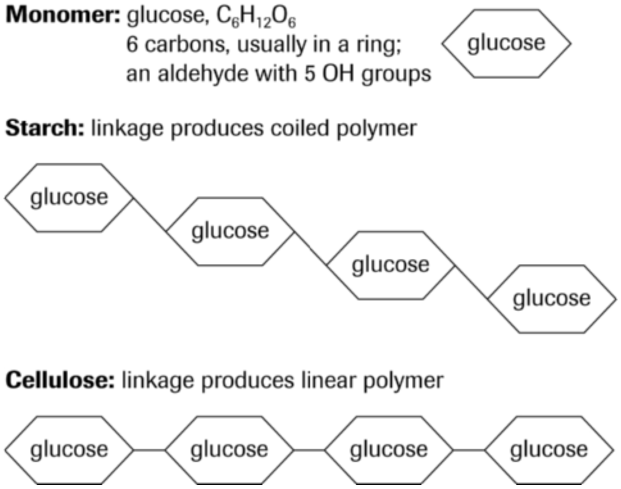
Instructional Master LSM 10.5E

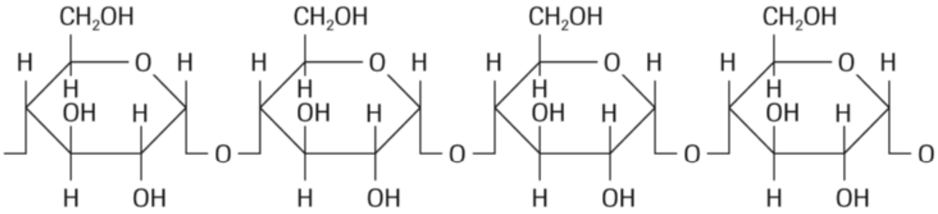
Carbohydrates—Starch and Cellulose

Monomer: glucose, $C_6H_{12}O_6$
6 carbons, usually in a ring;
an aldehyde with 5 OH groups

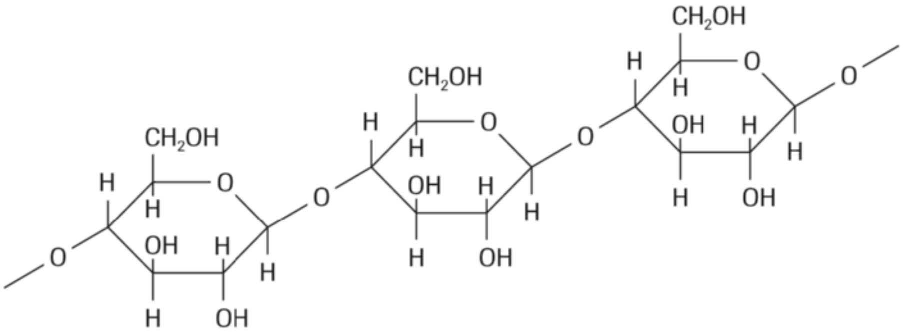
Starch: linkage produces coiled polymer

Cellulose: linkage produces linear polymer





(a) starch



(b) cellulose

Figure 11
The difference in linkages between glucose monomers gives very different three-dimensional structures.
(a) In starch, the polymer takes on a tightly coiled helical structure.
(b) In cellulose, the linked monomers can rotate, allowing formation of straight fibres.

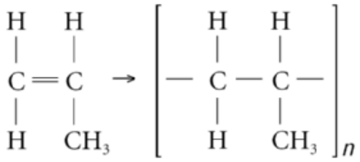
Polymers—Plastics, Fibres, and Food

SUMMARY Polymers—Plastics, Fibres, and Food

Polymers are a common part our everyday world and include natural polymers (lipids (triesters), proteins, and carbohydrates) and synthetic polymers (e.g., nylon, cellulose acetate, and polyesters). For the structural formula equations of these polymers, see previous pages.

Addition Polymers

Synthetic Addition Polymer (e.g., polypropene)



Condensation Polymers (Structural Analogs)

Polyesters

- Natural "Polyester" (e.g., butter):
  glycerol + fatty acid → fat or oil + water
  e.g., $C_3H_5(OH)_3 + 3\ C_{17}H_{35}COOH \rightarrow C_3H_5(OOCC_{17}H_{35})_3 + 3\ H_2O$
- Synthetic Polyester (e.g., Dacron)
  dicarboxylic acid + polyalcohol → polyester + water
  e.g., $HOOC - C_6H_4 - COOH + HO - CH_2 - CH_2 - OH \rightarrow$
  $[-OOC - C_6H_4 - COO - CH_2 - CH_2-]_n + n\,H_2O$

Polyamides

- Natural (proteins; polypeptides):
  amino acid + amino acid + … → protein + water
  e.g., $H_2N - CH_2 - COOH + H_2N - CHCH_3 - COOH \rightarrow$
  $[-NH - CH_2 - CONH - CHCH_3 - CO-]_n + n\,H_2O$
- Synthetic (e.g., nylon 6,6):
  dicarboxylic acid + di-amine → nylon + water
  e.g., $HOOC - (CH_2)_4 - COOH + H_2N - (CH_2)_6 - NH_2 \rightarrow$
  $[- OC - (CH_2)_4 - CONH - (CH_2)_6 - NH-]_n + n\,H_2O$

Polysaccharides

- Natural (e.g., starch):
  glucose + glucose → starch + water
  glucose + glucose → cellulose + water
  (The glucose molecules that polymerize to produce starch or cellulose have slightly different stereochemical formulas.)
  $C_6H_{12}O_6 + C_6H_{12}O_6 \rightarrow [-C_6H_{10}O_5-]_n + n\,H_2O$
- Synthetic (e.g., cellulose triacetate):
  cellulose + acetic acid + acetic anhydride → cellulose acetate + ...
  $[-C_6H_{10}O_5-]_n + CH_3COOH + (CH_3CO)_2O \rightarrow [-C_6H_7O_5(OCH_3)_3-]_n + ...$

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Student Worksheet LSM 10.CS

Chapter Checklist

Have you mastered the concepts, applications, and skills associated with the following items? Check them off when you are confident in your understanding.

Knowledge

identify and describe significant organic compounds in daily life, demonstrating generalized knowledge of their origins and applications (all sections)	
name and draw structural, condensed structural and line diagrams and formulas for organic halides, alcohols, carboxylic acids, and esters (all sections)	
identify types of compounds from their functional groups, given the structural formula and name the functional groups (10.2, 10.3, 10.4)	
compare, both within a homologous series and between compounds with different functional groups, the boiling points and solubility of examples of alcohols and carboxylic acids (10.3, 10.4)	
define, illustrate, and provide examples of simple addition, substitution, elimination, and esterification (condensation) reactions (10.2, 10.3, 10.4, 10.5)	
predict products and write and interpret balanced equations for the above reactions (10.2, 10.3, 10.4, 10.5)	
define, illustrate, and give examples of monomers, polymers, and polymerization in living and non-living systems (10.5)	
relate the reactions described above to major reactions for producing economically important compounds from fossil fuels (all sections)	

STS

illustrate how science and technology are developed to meet societal needs and expand human capabilities (all sections)	
describe interactions of science, technology, and society (all sections)	
illustrate how science and technology have both intended and unintended consequences (10.5)	

Skills

initiating and planning	predict the ester formed from an alcohol and an organic acid (10.4)	
	describe procedures for safe handling, storing and disposal of materials used in the laboratory, with reference to labelling information (10.2, 10.3, 10.4, 10.5)	
	design an experiment to compare the properties of organic to inorganic compounds (10.4)	
performing and recording	build molecular models depicting the structures of selected organic and inorganic compounds (10.4)	
	perform an experiment to investigate the reactions of organic compounds (10.2, 10.3, 10.4, 10.5)	
analyzing and interpreting	follow appropriate IUPAC guidelines in writing the names and formulas of organic compounds (all sections)	
	compile and organize data to compare the properties of structural isomers (10.3)	

Skills (cont'd)

	interpret results of a test to distinguish between a saturated and an unsaturated aliphatic using aqueous bromine solution (10.2)	
	use appropriate chemical symbols and nomenclature in writing organic chemical reactions (all sections)	
	use models to illustrate polymerization (10.5)	
communi-cation and teamwork	work cooperatively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results (10.3, 10.4, 10.5)	

Key Terms

10.2	hydrocarbon derivative	
	organic halide	
	functional group	
	addition reaction	
	substitution reaction	
10.3	alcohol	
	hydroxyl group	
	elimination reaction	
10.4	carboxylic acid	
	carboxyl group	
	condensation reaction	
	ester	
	esterification reaction	
	ester functional group	
10.5	polymer	
	monomer	
	polymerization	
	addition polymerization	
	condensation polymerization	
	polyester	
	polyamide	
	polypeptide	

Key STS

name common organic compounds in familiar products (Section 10.3; Section 10.4; DYK: Water-Soluble Vitamins; Section 10.5; WA: WQ—Teflon: Healthy or Hazardous?; BC: Correlation versus Cause and Effect; EI: Natural or Artificial Polymers?; DYK: Cotton—A Natural Polymer)	
recognize petrochemical industry products and by-products *(all sections)*	
recognize the similarity between natural and synthetic polymers (BC: Natural Esters; Section 10.4 Questions 7, 9, 10, 13; Section 10.5; BC: Correlation versus Cause and Effect; EI: Natural or Artificial Polymers?; DYK: Cotton—A Natural Polymer; WA: CS—Cellulose Acetate)	
understand the reasons for the modification of gasoline (Section 10.3; DYK: Ethanol in Alberta; DYK: Fill Up with Methanol; WA: WQ— Cellulosic Ethanol; Section 10.3 Question 12; also see Chapter 9)	
weigh the positive and negative effects of various reactions involving organic compounds (Exploration: Burning Fossil Fuels; Section 10.1 Questions 1, 2; DYK: Ozone Depletion; Section 10.2 Questions 5, 8; Section 10.3; DYK: Ethanol in Alberta; DYK: Alcohol Poisoning; WA: CS—Cellulose Acetate; Section 10.3 Questions 12, 13; Section 10.4 Question 13; WA: WQ—Teflon: Healthy or Hazardous?; WA: CS—Recycling Plasics; DYK: Biodegradable Polymers; DYK: Aboriginal Technologies; EI: Natural or Artificial Polymers?; WA: CS—Cellulose Acetate; Section 10.5 Questions 4, 5, 6, 9, 13	
explore various careers associated with the petrochemical industry *(CC: Chemical Engineer; CC: Chemical Employee)*	

Key Skills

prepare and react organic compounds *(Investigations 10.1, 10.2, 10.3, 10.5)*	

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Student Worksheet LSM 11.0A

Thermal Energy Calculations, Extra Exercises

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1. Calculate the quantity of heat required to warm 1.25 L of water from 22.0 °C to 98.0 °C in an electric kettle.

2. What mass of aluminum in a car engine will absorb 1.00 MJ of heat when the temperature rises from 22 °C to 102 °C after the car is started?

3. In a laboratory experiment, 2.00 kJ of heat flowed to a 100 g sample of a liquid solvent, causing a temperature increase from 15.40 °C to 21.37 °C. Calculate the specific heat capacity of the liquid.

4. A human body loses about 360 kJ of heat every hour. Assuming that an average human body is equivalent to about 60 kg of water, what temperature decrease would this heat transfer cause? (Of course, this heat is replaced by body metabolism.)

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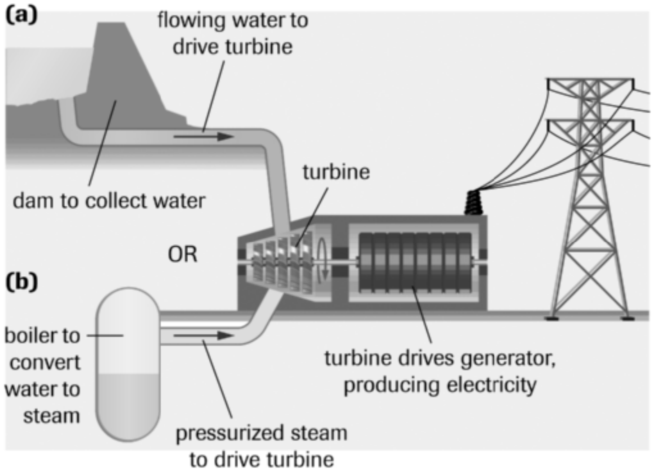
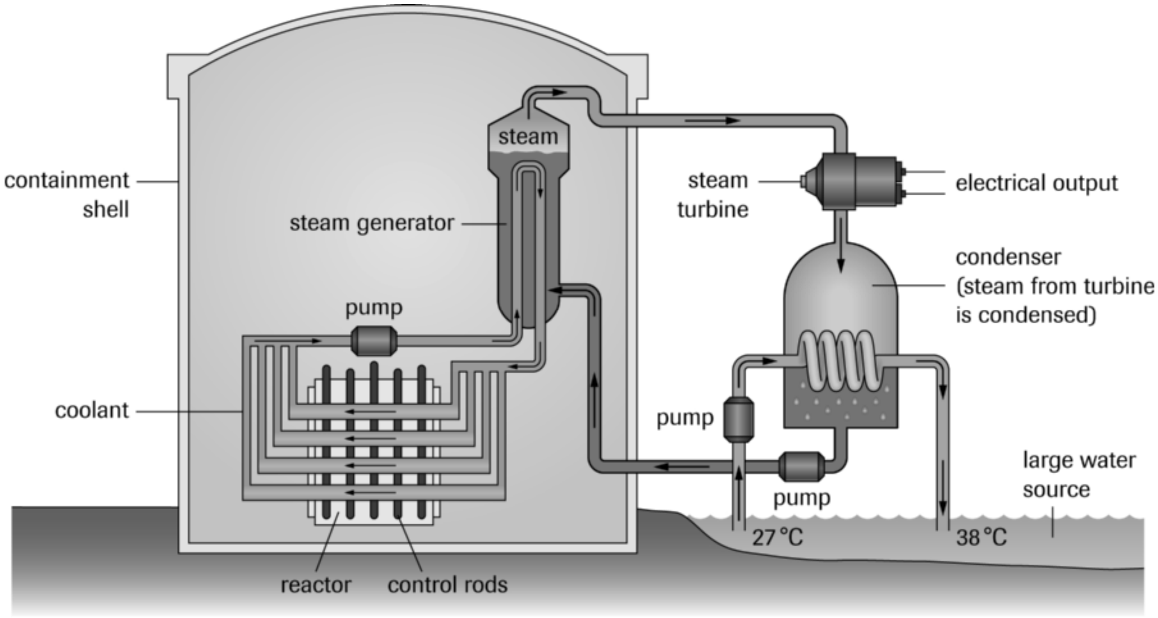
Instructional Master LSM 11.1A

Energy Production



There are many energy sources that can be used to spin a turbine. If the turbine is connected to a generator, the energy of motion is converted to electrical energy.

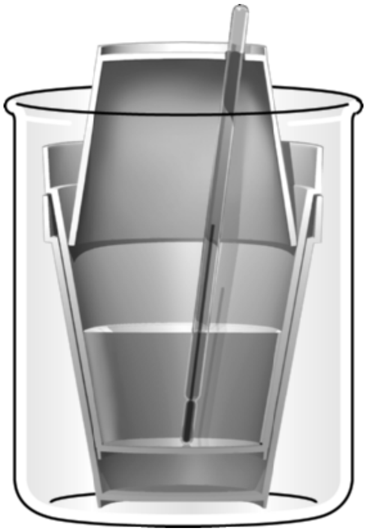


In nuclear reactors, the energy released in nuclear fission is absorbed by a primary coolant, which then transfers the energy to a secondary coolant. The energy from the secondary coolant is used in a turbine or engine to generate electrical energy.

Instructional Master LSM 11.2A

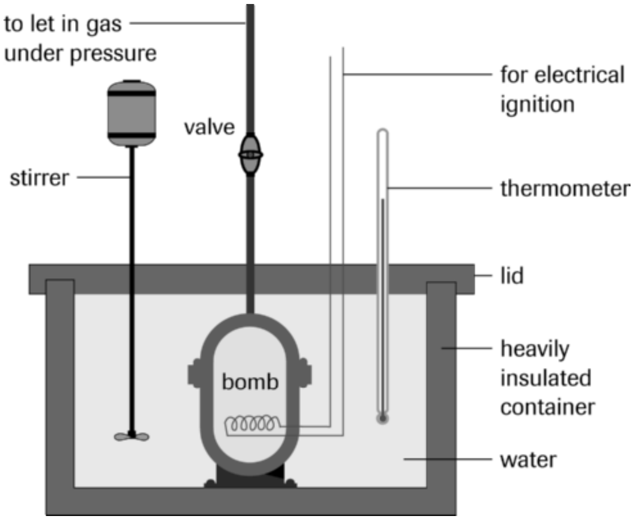
Calorimeters

Simple Calorimeter



A simple laboratory calorimeter consists of an insulated container made of three nested polystyrene cups, a measured quantity of water, and a thermometer. The chemical system is placed in or dissolved in the water of the calorimeter. Energy transfers between the chemical system and the surrounding water are monitored by measuring changes in the water temperature.

Bomb Calorimeter



The reactants are placed inside the calorimeter's bomb, which is surrounded by the calorimeter water. Once the calorimeter is sealed and the initial temperature measured, the combustion reaction is initiated by an electric heater or spark. Stirring is essential in order to obtain a uniform final temperature for the water.

Student Worksheet LSM 11.2B

Bomb Calorimetry, Extra Exercise (Extension)

Bomb calorimeters are used in research to measure enthalpy changes of combustion of fuels, foodstuffs, crops, and explosives. It is a useful technology, but calorimeters that are larger and more sophisticated than polystyrene cups usually have a noticeable heat transfer to or from the calorimeter materials. The total energy change of the calorimeter is the sum of the energy changes of all of the components.

$$\underset{\text{calorimeter}}{Q} = \underset{\text{water}}{mc\Delta t} + \underset{\text{containers}}{m_1c_1\Delta t_1} + \underset{\text{stirrer}}{m_2c_2\Delta t_2} + \underset{\text{thermometer}}{m_3c_3\Delta t_3}$$

Because the temperature change is identical for all components and the same components are used over and over again, this total energy calculation can be simplified. The different constants in the equation can be replaced by a single constant, C, the heat capacity of the particular calorimeter. The **heat capacity** of a calorimeter is the total energy absorbed or released per degree Celsius for the calorimeter and its contents. Its units are J/°C or kJ/°C.

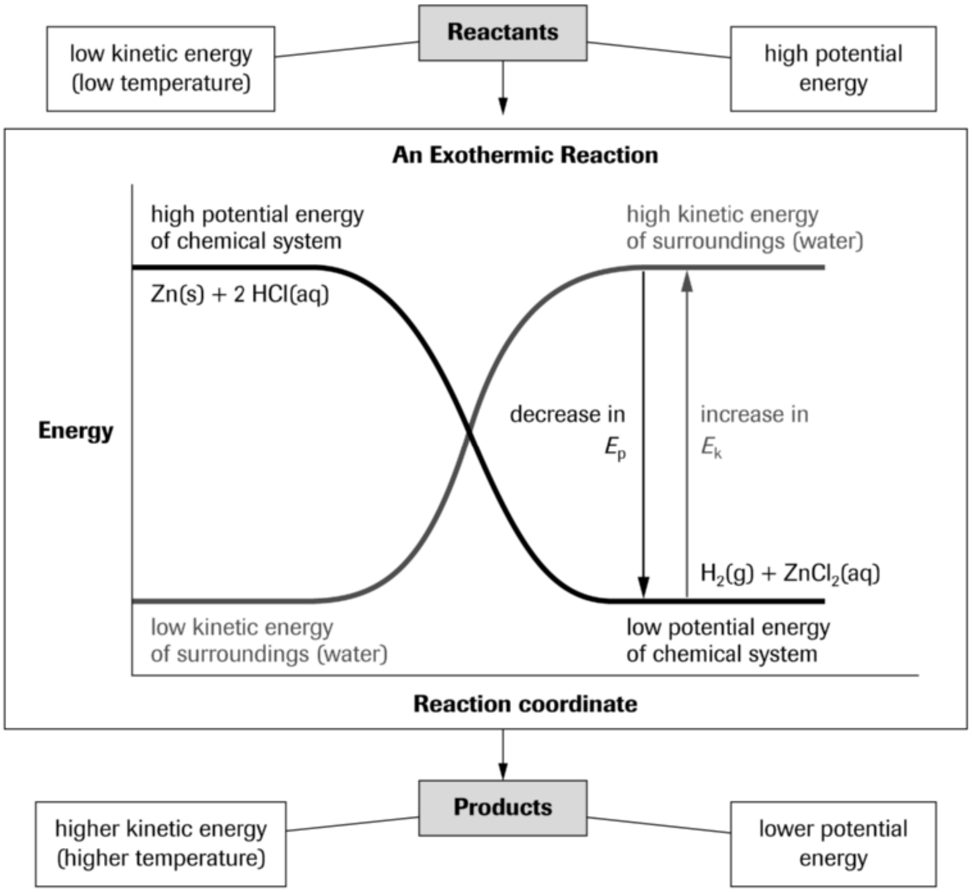
$$Q = C\Delta t$$

1. An oxygen bomb calorimeter has a heat capacity of 6.49 kJ/°C. The complete combustion of 1.12 g of ethyne, $C_2H_2(g)$, produces a temperature change from 18.60 °C to 27.15 °C. Calculate the molar enthalpy of combustion, Δ_cH_m, for ethyne.

2. Canadian inventors have developed zeolite, a natural aluminium silicate mineral, as a storage medium for solar heat. Zeolite releases heat when hydrated with water. In a test, zeolite is used to heat water in a tank that has a heat capacity of 157 kJ/°C. What is the enthalpy change of hydration (ΔH_h) for zeolite if the temperature of the water increases from 27 °C to 73 °C?

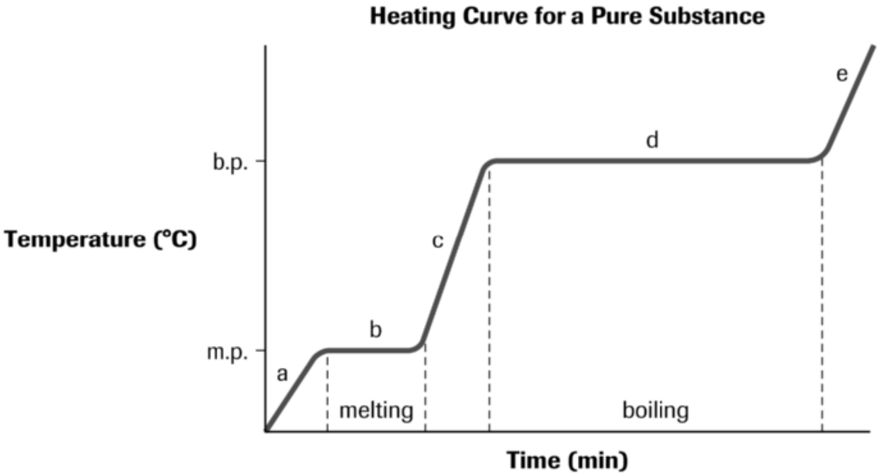
3. Besides the molar enthalpy of combustion as determined in a bomb calorimeter, what other properties or factors are involved in evaluating automobile fuels such as propane, ethanol, and hydrogen?

Energy Transfer Concept Diagram

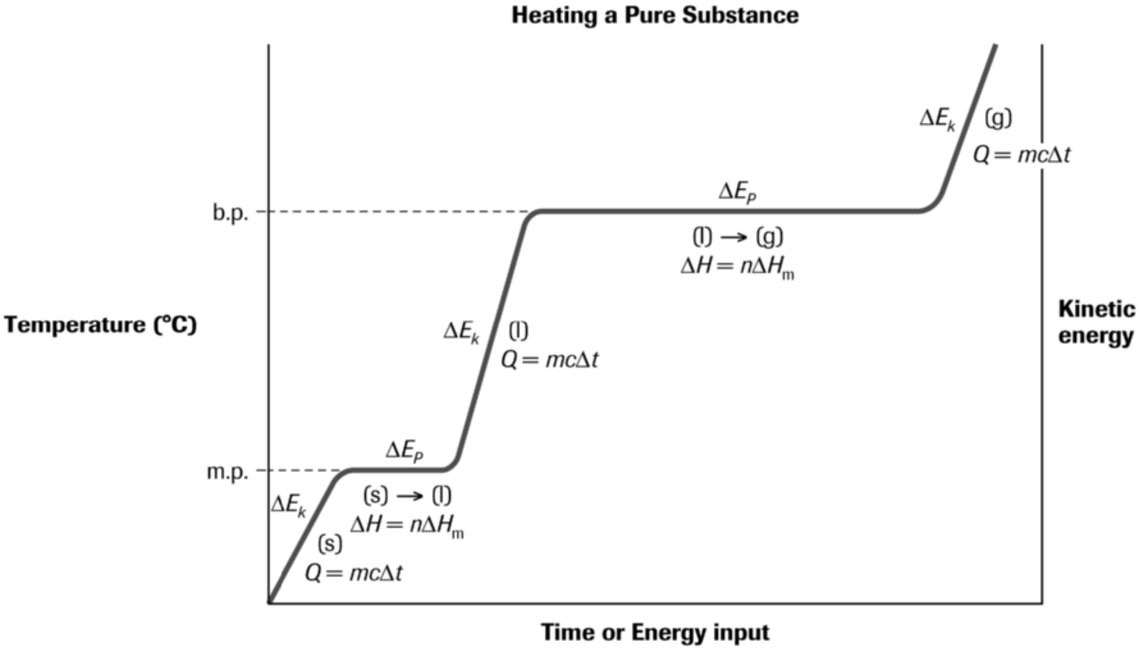


During an exothermic chemical reaction, some of the potential energy is converted to kinetic energy, which then gets transferred to the surroundings.

Heating Curves (Extension)



As a pure substance is heated, the following changes can occur:
a an increase in the temperature of the solid (Q);
b a phase change from solid to liquid (ΔH);
c an increase in the temperature of the liquid (Q);
d a phase change from liquid to gas (ΔH); and
e an increase in the temperature of the gas (Q).



Student Worksheet LSM 11.2I

Enthalpy Changes and Calorimetry, Extra Exercises

1. An experiment produces evidence that the evaporation of 4.00 g of liquid butane, $C_4H_{10}(l)$, requires a gain in enthalpy of 1.67 kJ. Find the molar enthalpy of vaporization for butane from this evidence.

2. A 40.0 kg industrial calorimeter has an overall specific heat capacity of 1.00 J/(g•°C). Complete combustion of 1.00 g of hydrogen in this calorimeter causes a temperature increase of 3.54 °C. Calculate the molar enthalpy of combustion for hydrogen from this evidence.

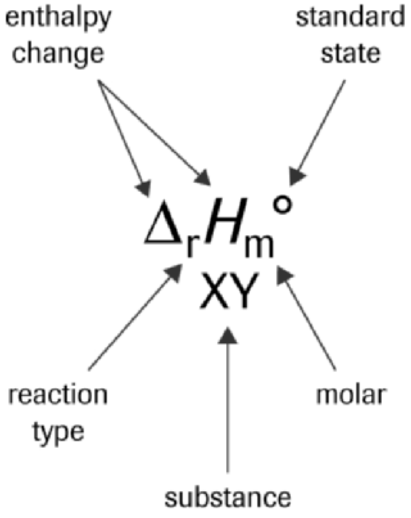
3. Combustion of 3.50 g of ethanol, $C_2H_5OH(l)$, in a calorimeter with an overall specific heat capacity of 1.52 J/(g•°C) causes a temperature increase from 19.88 °C to 26.18 °C. The mass of the calorimeter, including water, is 10.0 kg. Find the molar enthalpy of combustion for ethanol from this evidence.

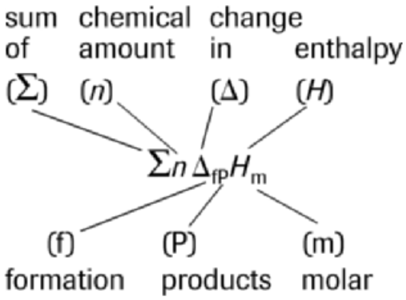
Thermochemistry Symbols and Equations

$$\overbrace{\Delta_r H}^{\text{enthalpy change of reaction}} = \overbrace{n}^{\text{chemical amount}}\ \overbrace{\Delta_r H_m}^{\text{molar enthalpy of reaction}}$$

for example, $\Delta_c H = n\ \Delta_c H_m$ (c for combustion)

$$\underbrace{n}_{\text{chemical amount of fuel}}\ \underbrace{\Delta_c H_m}_{\text{molar enthalpy of combustion of fuel}} = \underbrace{m}_{\text{mass of water}}\underbrace{c}_{\text{specific heat capacity of water}}\underbrace{\Delta t}_{\text{temperature change of water}}$$

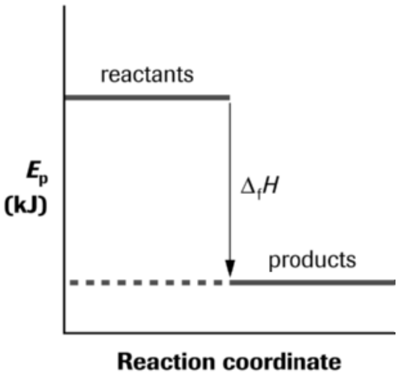




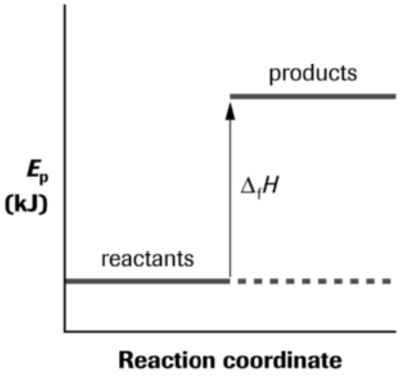
Instructional Master LSM 11.3B

Exothermic and Endothermic Changes

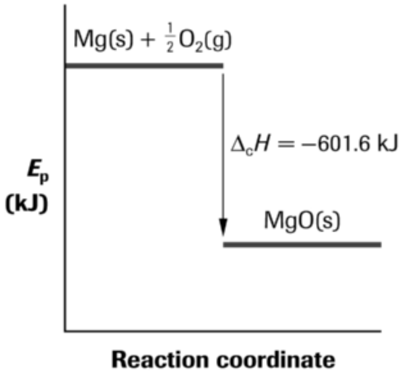
Exothermic Reaction



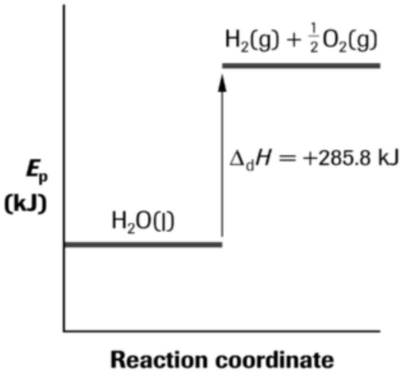
Endothermic Reaction



(a) **Combustion of Magnesium**



(b) **Decomposition of Water**



Communicating Energy Changes

SUMMARY Four Ways of Communicating Energy Changes

	Exothermic Changes	Endothermic Changes
1. Molar Enthalpy	$\Delta_r H_m < 0$	$\Delta_r H_m > 0$
2. Enthalpy Change	reactants → products; $\Delta_r H < 0$	reactants → products; $\Delta_r H > 0$
3. Term in a Balanced Equation	reactants → products + energy	reactants + energy → products
4. Chemical Potential Energy Diagram	E_p (reactants) > E_p (products)	E_p (reactants) < E_p (products)

Student Worksheet LSM 11.3D

Representing Enthalpy Changes, Extra Exercises

1. Iron(II) sulfide ore is roasted according to the following chemical equation.
   $4\ FeS(s) + 7\ O_2(g) \rightarrow 2\ Fe_2O_3(s) + 4\ SO_2(g)$ $\Delta_cH = -2456$ kJ

   (a) Rewrite this chemical equation including the energy as a term in the balanced equation.

   (b) What is the molar enthalpy for iron(II) sulfide in this reaction?

   (c) What is the molar enthalpy for iron(III) oxide in this reaction?

2. Boron reacts with hydrogen to form diboron hexahydride (diborane) gas. The molar enthalpy of reaction for boron is +15.7 kJ/mol. Write the balanced chemical equation using whole number coefficients and
   (a) including the energy change as a Δ_rH.

   (b) including the energy change as a term in the balanced equation.

3. The molar enthalpy of combustion for octane, $C_8H_{18(l)}$, is reported to be –1.3 MJ/mol. Write the balanced chemical equation using whole number coefficients and
   (a) including the energy change as a Δ_rH.

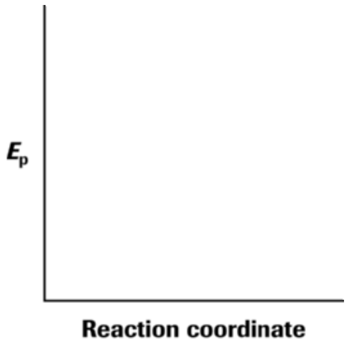
   (b) including the energy change as a term in the balanced equation.

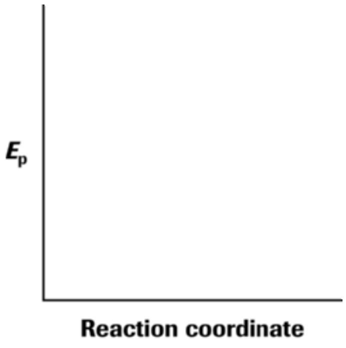
4. Draw potential energy diagrams to communicate the following chemical reactions. Assume SATP conditions.

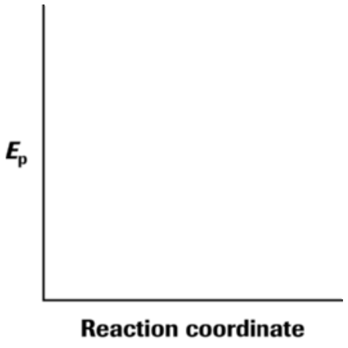
   (a) the formation of chromium (III) oxide

   (b) the simple decomposition of silver iodide

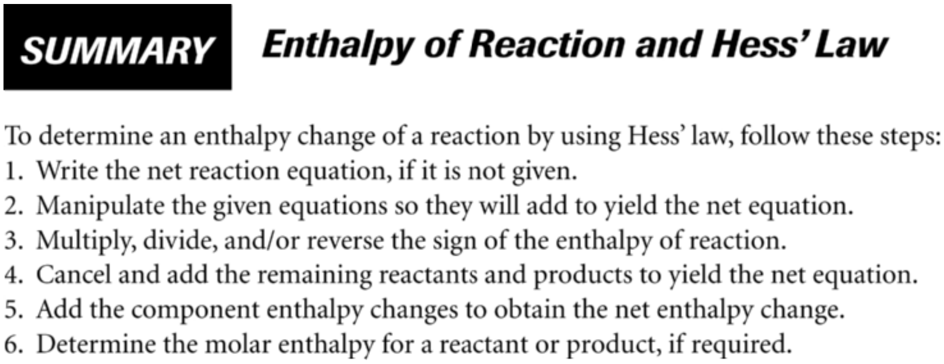
   (c) the formation of carbon disulfide





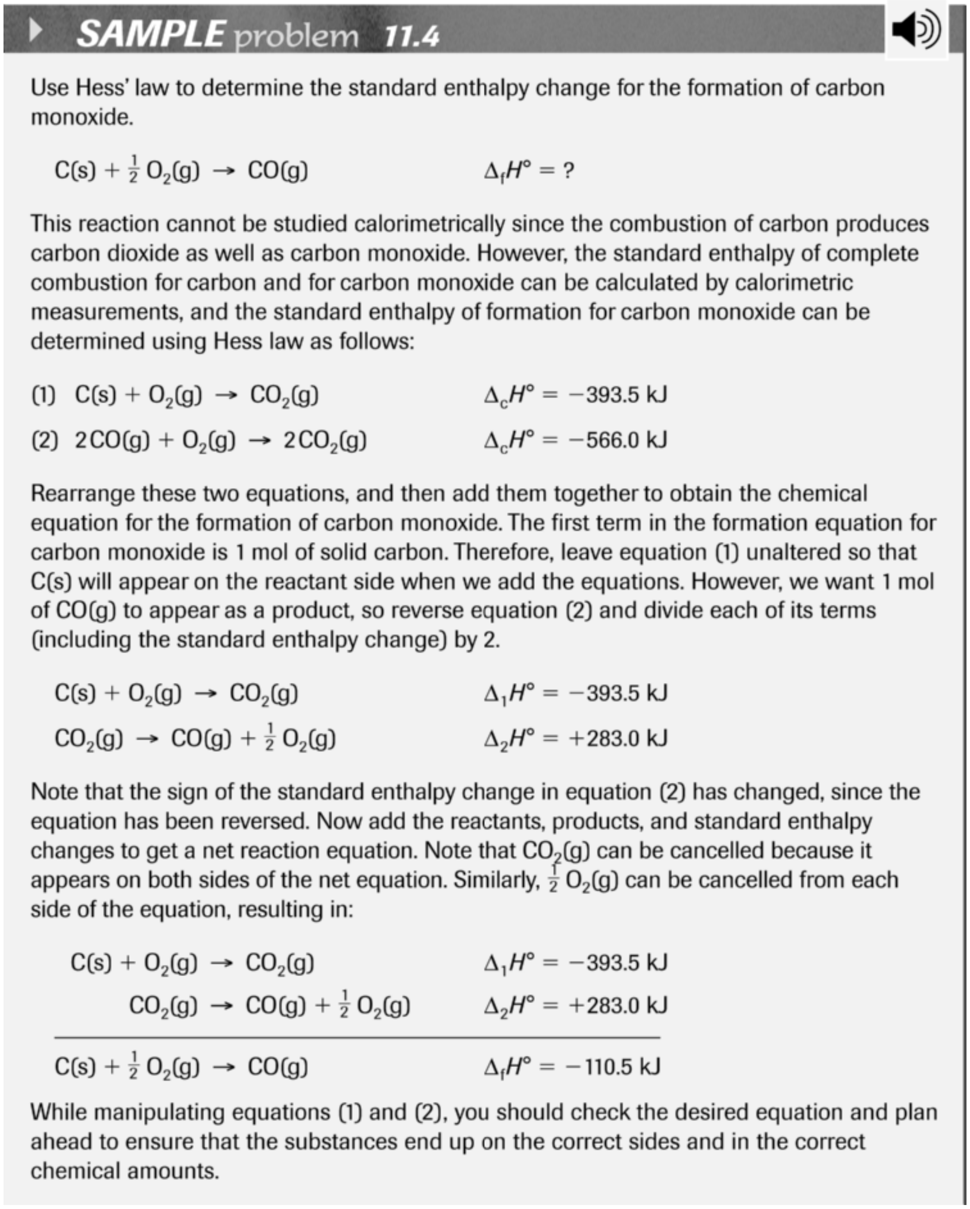


Hess' Law



SUMMARY Enthalpy of Reaction and Hess' Law

To determine an enthalpy change of a reaction by using Hess' law, follow these steps:
1. Write the net reaction equation, if it is not given.
2. Manipulate the given equations so they will add to yield the net equation.
3. Multiply, divide, and/or reverse the sign of the enthalpy of reaction.
4. Cancel and add the remaining reactants and products to yield the net equation.
5. Add the component enthalpy changes to obtain the net enthalpy change.
6. Determine the molar enthalpy for a reactant or product, if required.



SAMPLE problem 11.4

Use Hess' law to determine the standard enthalpy change for the formation of carbon monoxide.

$$C(s) + \frac{1}{2} O_2(g) \rightarrow CO(g) \qquad \Delta_f H^\circ = ?$$

This reaction cannot be studied calorimetrically since the combustion of carbon produces carbon dioxide as well as carbon monoxide. However, the standard enthalpy of complete combustion for carbon and for carbon monoxide can be calculated by calorimetric measurements, and the standard enthalpy of formation for carbon monoxide can be determined using Hess law as follows:

$$(1) \quad C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta_c H^\circ = -393.5\ \text{kJ}$$

$$(2) \quad 2CO(g) + O_2(g) \rightarrow 2CO_2(g) \qquad \Delta_c H^\circ = -566.0\ \text{kJ}$$

Rearrange these two equations, and then add them together to obtain the chemical equation for the formation of carbon monoxide. The first term in the formation equation for carbon monoxide is 1 mol of solid carbon. Therefore, leave equation (1) unaltered so that C(s) will appear on the reactant side when we add the equations. However, we want 1 mol of CO(g) to appear as a product, so reverse equation (2) and divide each of its terms (including the standard enthalpy change) by 2.

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta_1 H^\circ = -393.5\ \text{kJ}$$

$$CO_2(g) \rightarrow CO(g) + \frac{1}{2} O_2(g) \qquad \Delta_2 H^\circ = +283.0\ \text{kJ}$$

Note that the sign of the standard enthalpy change in equation (2) has changed, since the equation has been reversed. Now add the reactants, products, and standard enthalpy changes to get a net reaction equation. Note that $CO_2(g)$ can be cancelled because it appears on both sides of the net equation. Similarly, $\frac{1}{2} O_2(g)$ can be cancelled from each side of the equation, resulting in:

$$\begin{array}{ll} C(s) + O_2(g) \rightarrow CO_2(g) & \Delta_1 H^\circ = -393.5\ \text{kJ} \\ \quad CO_2(g) \rightarrow CO(g) + \frac{1}{2} O_2(g) & \Delta_2 H^\circ = +283.0\ \text{kJ} \\ \hline C(s) + \frac{1}{2} O_2(g) \rightarrow CO(g) & \Delta_f H^\circ = -110.5\ \text{kJ} \end{array}$$

While manipulating equations (1) and (2), you should check the desired equation and plan ahead to ensure that the substances end up on the correct sides and in the correct chemical amounts.

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Student Worksheet LSM 11.4E

Predicting ΔH Using Hess' Law, Extra Exercises

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1. The enthalpy changes for the formation of two tungsten bromides are shown below.
   $W(s) + 2\ Br_2(l) \rightarrow WBr_4(s)$ $\Delta_1 H° = –146.7$ kJ
   $W(s) + 3\ Br_2(l) \rightarrow WBr_6(s)$ $\Delta_2 H° = –184.4$ kJ

   Calculate the standard enthalpy change for the following reaction:
   $Br_2(l) + WBr_4(s) \rightarrow WBr_6(s)$

2. Given: (1) $N_2O_4(g) \rightarrow 2\ NO_2(g)$ $\Delta_1 H° = +58$ kJ
   (2) $NO(g) + \frac{1}{2}\ O_2(g) \rightarrow NO_2(g)$ $\Delta_2 H° = –56$kJ

   Calculate the standard enthalpy change for the following reaction:
   $2\ NO(g) + O_2(g) \rightarrow N_2O_4(g)$

3. Use the following reactions and enthalpy changes to predict the standard enthalpy change for:
   $2\ NO_2(g) + 2\ H_2O(g) \rightarrow 3\ O_2(g) + N_2H_4(g)$
   (1) $\frac{1}{2}\ N_2(g) + O_2(g) \rightarrow NO_2(g)$ $\Delta_1 H° = +33.2$ kJ
   (2) $H_2(g) + \frac{1}{2}\ O_2(g) \rightarrow H_2O(g)$ $\Delta_2 H° = –241.8$ kJ
   (3) $N_2(g) + 2\ H_2(g) \rightarrow N_2H_4(g)$ $\Delta_3 H° = +47.6$ kJ

4. Use the following formation reaction evidence to calculate the standard enthalpy change for the complete combustion of cycloheptane.
   (1) $C(s) + O_2(g) \rightarrow CO\ (g)$ $\Delta_1 H° = –393.5$ kJ
   (2) $H_2(g) + \frac{1}{2}\ O_2(g) \rightarrow H_2O(g)$ $\Delta_2 H° = –241.8$ kJ
   (3) $7\ C(s) + 7\ H_2\ (g) \rightarrow C_7H_{14}(l)$ $\Delta_3 H° = +115.0$ kJ

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Student Worksheet LSM 11.4G

Hess' Law, Extra Exercises

Instructions

1. Cut the known equations into strips. You will have a known equation A on one side of the strip, and its reverse known equation B on the other side.
2. Group the known equations with their proper target equations below.
3. In each group, flip the known equations until the products and the reactants match those of the target equation.
4. Attach (using tape or glue) the known equations on this sheet so the appropriate known equations are above the target equations.
5. Number the known equations in each group.
6. Multiply the known equations in each group until the stoichiometry of the known equations matches that of the target equation.
7. Add the equations and enthalpies to solve for the target equation.

(a) What is the enthalpy change for the formation of two moles of nitrogen monoxide from its elements?

$N_2(g) + O_2(g) \rightarrow 2\ NO(g) \qquad \Delta H° = ?$

(b) What is the enthalpy change for the formation of one mole of butane, $C_4H_{10}(g)$, from its elements?

$4\ C(s) + 5\ H_2(g) \rightarrow C_4H_{10}(g) \qquad \Delta H° = ?$

(c) Determine the enthalpy change involved in the formation of two moles of liquid propanol.

$6\ C(s) + 8\ H_2(g) + O_2(g) \rightarrow 2\ C_3H_7OH(l) \qquad \Delta H° = ?$

Known Equations A

*[**Teacher note:** Photocopy Known Equations A on one side of a sheet and Known Equations B on the reverse side of the same sheet.]*

$\mathbf{C(s) + O_2(g) \rightarrow CO_2(g)}$	$\mathbf{\Delta \mathit{H}° = -393.5\ kJ}$
$\mathbf{C(s) + O_2(g) \rightarrow CO_2(g)}$	$\mathbf{\Delta \mathit{H}° = -393.5\ kJ}$
$\mathbf{C_3H_7OH(l) + 9/2\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)}$	$\mathbf{\Delta \mathit{H}° = -2008\ kJ}$
$\mathbf{C_4H_{10}(g) + 13/2\ O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)}$	$\mathbf{\Delta \mathit{H}° = -2657\ kJ}$
$\mathbf{\frac{1}{2}\ N_2(g) + O_2(g) \rightarrow NO_2(g)}$	$\mathbf{\Delta \mathit{H}° = +34\ kJ}$
$\mathbf{H_2O(l) + H_2(g) \rightarrow \frac{1}{2}\ O_2(g)}$	$\mathbf{\Delta \mathit{H}° = +286\ kJ}$
$\mathbf{2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)}$	$\mathbf{\Delta \mathit{H}° = -483.6\ kJ}$
$\mathbf{NO_2(g) \rightarrow NO(g) + \frac{1}{2}\ O_2(g)}$	$\mathbf{\Delta \mathit{H}° = +56\ kJ}$

Known Equations B

*[**Teacher note:** Photocopy Known Equations A on one side of a sheet and Known Equations B on the reverse side of the same sheet.]*

$CO_2(g) \rightarrow C(s) + O_2(g)$ **$\Delta H° = +393.5$ kJ**

$CO_2(g) \rightarrow C(s) + O_2(g)$ **$\Delta H° = +393.5$ kJ**

$3\ CO_2(g) + 4\ H_2O(l) \rightarrow C_3H_7OH(l) + \frac{9}{2}\ O_2(g)$ **$\Delta H° = +2008$ kJ**

$4\ CO_2(g) + 5\ H_2O(g) \rightarrow C_4H_{10}(g) + \frac{13}{2}\ O_2(g)$ **$\Delta H° = +2657$ kJ**

$NO_2(g) \rightarrow \frac{1}{2}\ N_2(g) + O_2(g)$ **$\Delta H° = -34$ kJ**

$\frac{1}{2}\ O_2(g) \rightarrow H_2O(l) + H_2(g)$ **$\Delta H° = -286$ kJ**

$2\ H_2O(g) \rightarrow 2\ H_2(g) + O_2(g)$ **$\Delta H° = +483.6$ kJ**

$\frac{1}{2}\ O_2(g) + NO(g) \rightarrow NO_2(g)$ **$\Delta H° = -56$ kJ**

Instructional Master LSM 11.5A

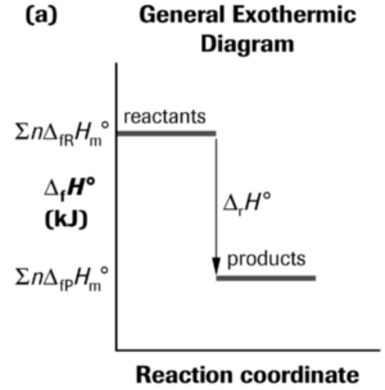
Enthalpies of Formation—Symbols of Use

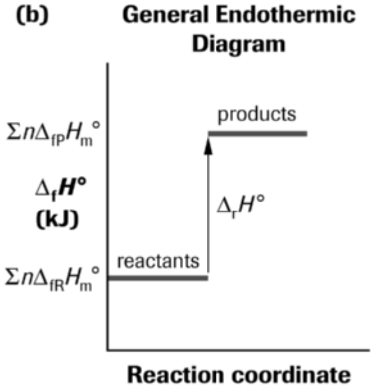
SUMMARY

Using Enthalpies of Formation to Predict Standard Enthalpy Changes, $\Delta_r H°$

According to Hess' law, the net enthalpy change for a chemical reaction is equal to the sum of the chemical amounts times the molar enthalpies of formation of the products minus the sum of the chemical amounts times the molar enthalpies of formation of the reactants.

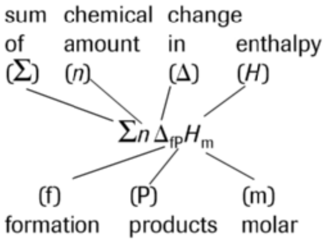
$$\Delta_r H° = \Sigma n \Delta_{f\,P} H_m° - \Sigma n \Delta_{f\,R} H_m°$$





Learning Tip

The symbols for this method of calculating the enthalpy of reaction are meant to help but, initially, can be confusing.



This set of symbols translates into: "the sum of the chemical amounts times the molar enthalpies of formation of the products." This set of symbols changes from P for the products to R for the reactants. Learn the concept (that is, the change in enthalpy is the energy of the products minus the energy of the reactants), and the equation then makes sense. See the Summary above.

Instructional Master **LSM 11.5C**

Summary of Energy Calculations

Table 1 Heat Transfer, Q (heating/cooling)

Equation	Empirical constant	Symbol	Typical units
$Q = mc\Delta t$	specific heat capacity	c	$\frac{\text{J}}{\text{g} \bullet {}^\circ\text{C}}$

Table 2 Enthalpy Change, $\Delta_r H_{mol} = n\Delta_r H$

Quantity	Alternative names	Symbol	Typical units
enthalpy change	heat of reaction change in heat content	$\Delta_r H$	kJ
molar enthalpy	molar heat heat content per mole	$\Delta_r H_m$* (substance)	$\frac{\text{kJ}}{\text{mol}}$

* In some resources the subscript m is omitted from the symbol. In this case, it is usually possible to infer from the context whether the symbol refers to molar enthalpy or enthalpy change.

Table 3 Classification of Reactions and $\Delta_r H$

Reaction type	Definition	Symbol	ΔH Names
formation	elements → compound	$\Delta_f H$	enthalpy of formation or heat of formation
combustion	substance + $O_2(g)$ → common oxides	$\Delta_c H$	enthalpy of combustion or heat of combustion
all other types*		$\Delta_r H$	enthalpy of reaction or heat of reaction

(superscript $^\circ$ means SATP(standard) conditions)

*If the reaction type has a specific name, then that name may be used to label enthalpy or ΔH; e.g., simple decomposition, $\Delta_{sd}H$ or neutralization $\Delta_n H$. Molar quantities are designated as, for example, $\Delta_f H_m$ and have units of kilojoules per mole (kJ/mol).

Problem-Solving Summary

1. Calorimetry Method

$$\underset{\text{(system)}}{\Delta_r H} = \underset{\text{(water equivalent)}}{Q}$$

$$\underset{\text{(substance)}}{n\Delta_r H_m} = mc\Delta t$$

2. Hess's Law $\Delta_r H^\circ = \Sigma\Delta_r H^\circ$

3. Formation Method $\Delta_r H^\circ = \Sigma n\Delta_{fp}H_m{}^\circ - \Sigma n\Delta_{fR}H_m{}^\circ$

Student Worksheet LSM 11.5D

Predicting ΔH Using Standard Enthalpies of Formation, Extra Exercises

1. Laboratory quantities of ethene can be prepared by an elimination reaction of ethanol using an acid catalyst. Calculate the enthalpy change for the conversion of ethanol into ethene and water using standard enthalpies of formation.

2. Calculate the molar enthalpy of combustion for acetic acid using standard enthalpies of formation.

3. An initial step in the production of iron in a blast furnace involves the conversion of iron(III) oxide and carbon monoxide into iron(II, III) oxide, $Fe_3O_4(s)$, and carbon dioxide. Calculate the enthalpy of this reaction using standard enthalpies of formation.

4. The fertilizer urea is produced along with liquid water by the reaction of ammonia and carbon dioxide. Calculate the enthalpy of this reaction using standard enthalpies of formation.

Student Worksheet LSM 11.5F

Multi-Step Energy Calculations, Extra Exercises

1. For the following combustion, what mass of carbon dioxide is produced when 1500 kJ of energy is released?
   $2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 4\ CO_2(g) + 6\ H_2O(g) + 2502\ \text{kJ}$

2. How much energy is released when 1.00 t of sulfur trioxide is produced by the following reaction?
   $2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g) \qquad \Delta_r H = -192.8\ \text{kJ}$

3. In respiration, glucose is oxidized by oxygen gas to produce carbon dioxide gas, liquid water, and energy. What is the energy released when 18.0 g of glucose is consumed?

4. Methanol is burned in a bomb calorimeter. Liquid water is formed as a product. If 3.40 g of methanol reacts, what is the expected temperature change in a 6.00 kg calorimeter with a heat capacity of 1.125 J/(g•°C)?

5. A waste heat exchanger is used to absorb the energy from the complete combustion of hydrogen sulfide gas. What volume of water undergoing a temperature change of 64 °C is required to absorb all of the energy from the burning of 15 kg of hydrogen sulfide?

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Student Worksheet LSM 11.CS

Chapter Checklist

Have you mastered the concepts, applications, and skills associated with the following items? Check them off when you are confident in your understanding.

Knowledge

explain how the sun is a major source of stored chemical energy on Earth (11.1)	
apply the equation $Q = mc\Delta t$ to the analysis of energy transfer (11.2)	
define enthalpy and molar enthalpy for chemical reactions (11.2, 11.3)	
use calorimetry evidence to determine enthalpy changes in chemical reactions (11.2, 11.3, 11.4, 11.5)	
describe photosynthesis, cellular respiration, and hydrocarbon combustion reactions, and understand that combustion and cellular respiration are similar and the reverse of photosynthesis (11.3)	
write and interpret balanced chemical equations incorporating $\Delta_r H$ notation (11.3)	
classify chemical reactions as endothermic or exothermic (11.3, 11.4, 11.5)	
explain and use Hess' law to calculate energy changes for a net reaction from a series of reactions (11.4)	
predict the standard enthalpy change for chemical equations using standard molar enthalpies of formation (11.5)	

STS

state that a goal of technology is to solve practical problems (11.1, 11.2)	
recognize that solving technological problems may require various solutions and have both intended and unintended consequences (11.1, 11.3, 11.4)	

Skills

initiating and planning	design a method to compare molar enthalpy changes when burning fuels (11.2)	
	describe procedures for safe handling, storage, and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information (11.2, 11.4)	
performing and recording	perform calorimetry experiments to determine molar enthalpy changes and use thermometers appropriately (11.2, 11.4)	
analyzing and interpreting	compare energy changes by analyzing data and energy diagrams (11.3, 11.4, 11.5)	
communication and teamwork	work collaboratively using appropriate notation and units for enthalpy changes and molar enthalpies (11.2, 11.3, 11.4, 11.5)	

Key Terms

11.1	thermochemistry	
	isolated system	
	calorimetry	
	calorimeter	
	heat	
	thermal energy, Q	
	specific heat capacity, c	
	joule (J)	
	enthalpy change, ΔH or ΔrH	
	exothermic reaction	
	endothermic reaction	
	enthalpy of reaction	
	molar enthalpy of reaction, $\Delta_r H_m$	
11.3	standard molar enthalpy of reaction, $\Delta_r H_m^\circ$	
	chemical potential energy diagram	
11.4	Hess' law	
11.5	reference energy state	
	standard enthalpy of formation $\Delta_f H^\circ$	
	thermal stability	

Key STS

consider one's personal reliance on chemical energy (CS: Personal Use of Chemical Energy; Section 11.4 Question 66; Chapter 11 Review questions 22, 30)	
explore uses of and alternatives to fossil fuels, including consequences (CS: Personal Use of Chemical Energy; Section 11.1 Questions 7-9, 12, 14, 15; Section 11.2 Questions 5, 9, 12; DYK: Renewable Energy Sources; Section 11.3 Question 8; DYK: Carbon Monoxide Poisoning; DYK: Cogeneration; EI: Alternative Energy Sources and Technologies; Section 11.4 Questions 1-5; WA: WQ—Rocket Fuel Thermochemistry; Section 11.5 Questions 5, 6; Chapter 11 Review questions 20, 31, 32)	
use energy more efficiently (CS: Personal Use of Chemical Energy; Section 11.1 Questions; WA: CS—Thermal Insulation)	

Key Skills

communicate energy changes of chemical reactions (Section 11.3; Section 11.4; Section 11.5; 6; Chapter 11 Review)	
design and use simple calorimeters (Investigation 11.1; Investigation 11.2; Investigation 11.3)	

Activation Energy and Potential Energy

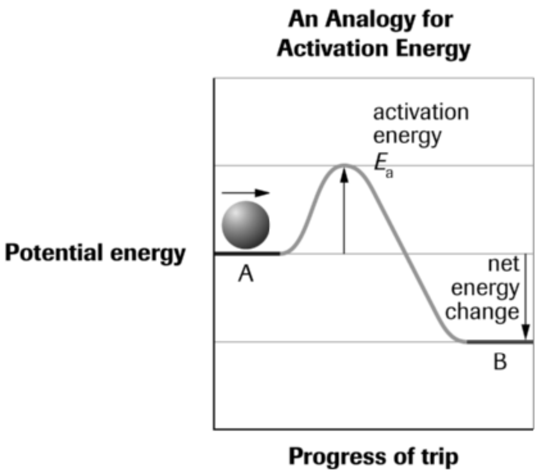
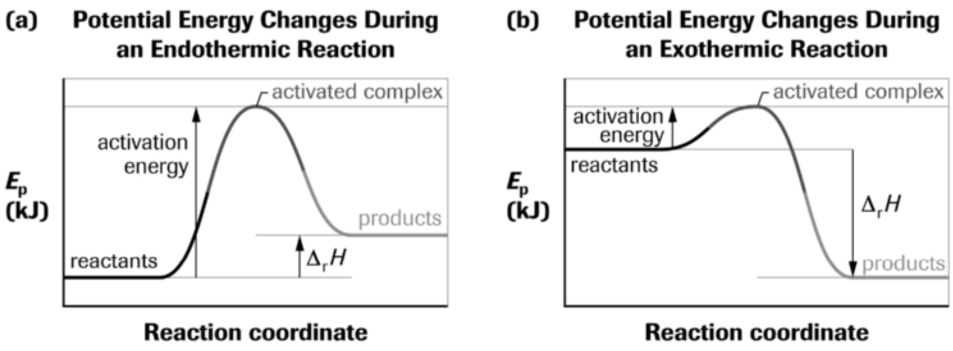


Figure 3
On a trip from A to B, there is a net decrease in potential energy, but there must be an initial increase in potential energy (activation energy) for the trip to be possible.



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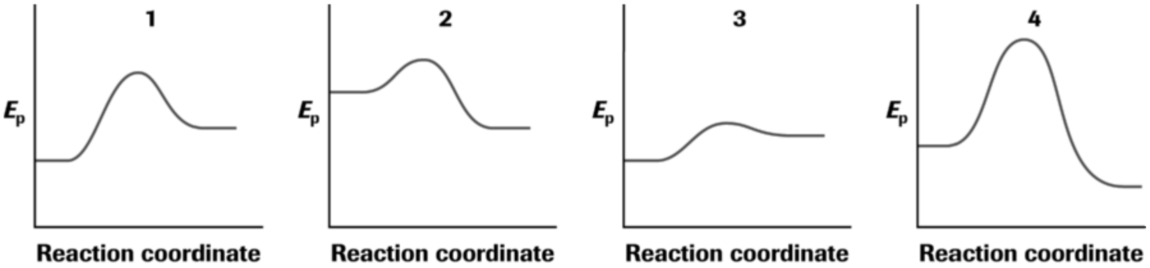
Student Worksheet LSM 12.1B

Energy Changes During a Chemical Reaction, Extra Exercises



For each of the diagrams above, explain the *reasoning* in complete sentence statements, in terms of the changes in potential energy and kinetic energy over the duration of the reaction.

1.

2.

3.

4.

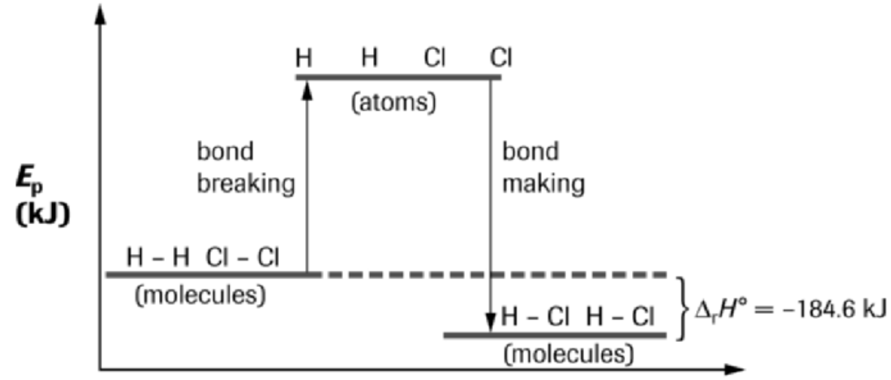
Bond Energy

SUMMARY Bond Energy and Enthalpy Changes

- Bond energy is the energy required to break a chemical bond; it is also the energy released when a bond is formed.
- The change in enthalpy represents the net effect from breaking and making bonds.
  $\Delta_r H$ = (energy released from bond making) − (energy required for bond breaking)
  Exothermic reaction: making > breaking ($\Delta_r H$ is negative.)
  Endothermic reaction: breaking > making ($\Delta_r H$ is positive.)

Example:

Energy Changes in the Formation of HCl



$H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g) + \text{energy}$

A Theoretical Explanation of Catalysis

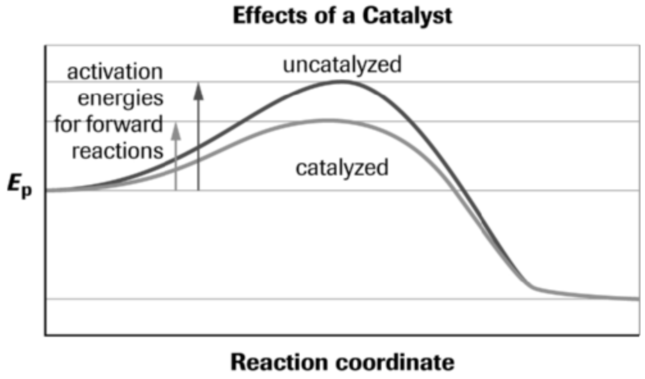
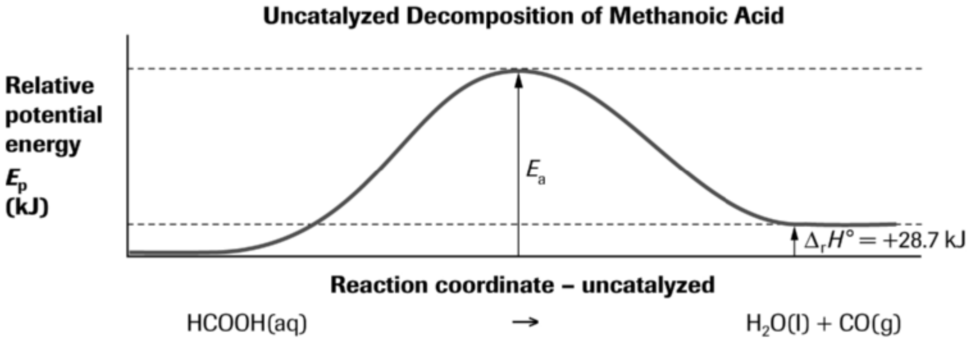
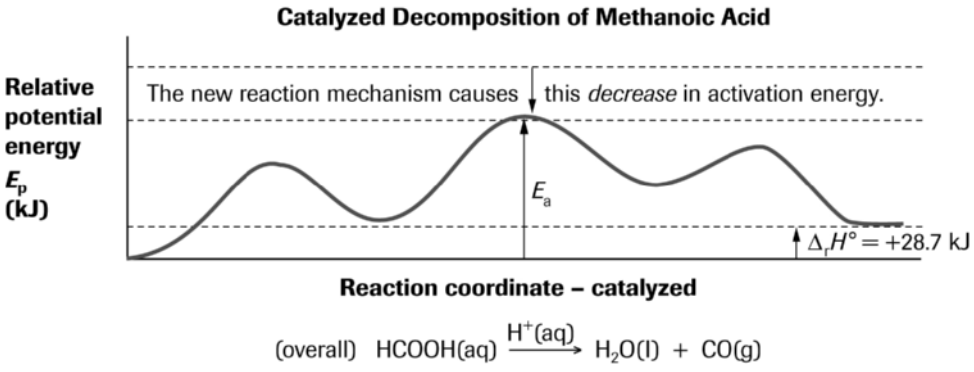


Figure 4
The catalyzed pathway has a lower activation energy, so more collisions lead to a successful reaction.





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Student Worksheet LSM 12.CS

Chapter Checklist

Have you mastered the concepts, applications, and skills associated with the following items? Check them off when you are confident in your understanding.

Knowledge

analyze and label energy diagrams for a chemical reaction, including reactants, products, enthalpy change, and activation energy (all sections)	
define activation energy as the energy barrier that must be overcome for a chemical reaction to occur (12.1)	
explain the energy changes that occur during chemical reactions referring to bonds breaking and forming and changes in potential and kinetic energy (12.2)	
explain that catalysts increase reaction rates by providing alternative pathways for changes without affecting the net energy involved (12.3)	

STS

recognize the values and limitations of technological products and processes (12.1, 12.3)	
state that a goal of technology is to solve practical problems (12.2)	
evaluate technologies from a variety of perspectives (12.3)	

Skills

initiating and planning	describe procedures for safe handling, storage, and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information (12.3)	
performing and recording	plot chemical potential energy diagrams, enthalpy diagrams, and energy pathway diagrams indicating changes in energy for chemical reactions (all sections)	
analyzing and interpreting	interpret energy diagrams for chemical reactions (all sections)	
communication and teamwork	work collaboratively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results by using appropriate SI notation, and fundamental and derived units for calculating and communicating enthalpy changes (all sections)	

Key Terms

12.1	activation energy, *Ea*	
12.2	bond energy	
12.3	catalysis	
	catalysis	
	enzyme	
	intermediate	

Key STS

give examples of the use and importance of catalysts to reduce air pollution from the burning of hydrocarbons	
give examples of the use and importance of enzymes in natural and technological systems	

Key Skills

design an experimental procedure to illustrate the effect of a catalyst on a chemical reaction *(Investigation 12.1, Section 12.3 Question 7)*	
create illustrated presentations (WA: WQ—Neurotransmitters and Nerve Agents; Section 12.1 Question 10; Unit 6 Review question 34	

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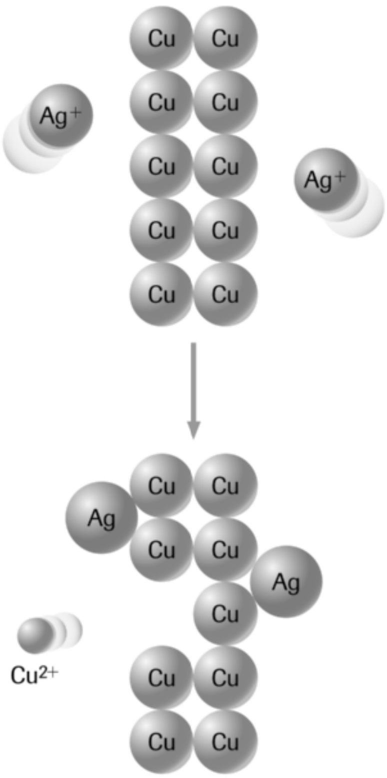
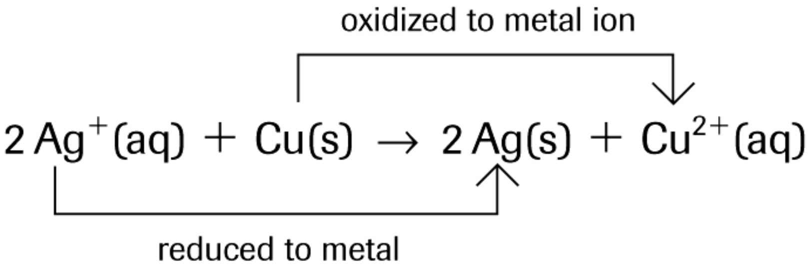
Instructional Master LSM13.1B

Oxidation and Reduction

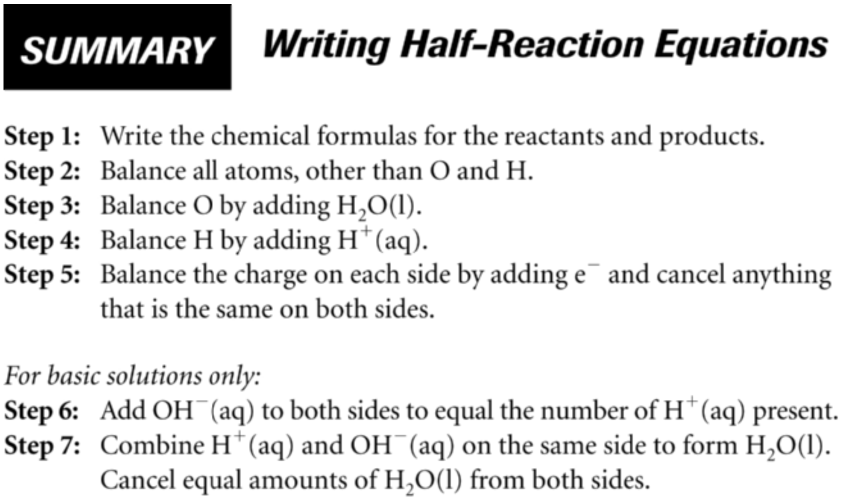




$$2\,Ag^+(aq) + Cu(s) \rightarrow 2\,Ag(s) + Cu^{2+}(aq)$$

"LEO the lion says "GER" is a mnemonic to help remember that
"Loss of Electrons is Oxidation" and "Gain of Electrons is Reduction."

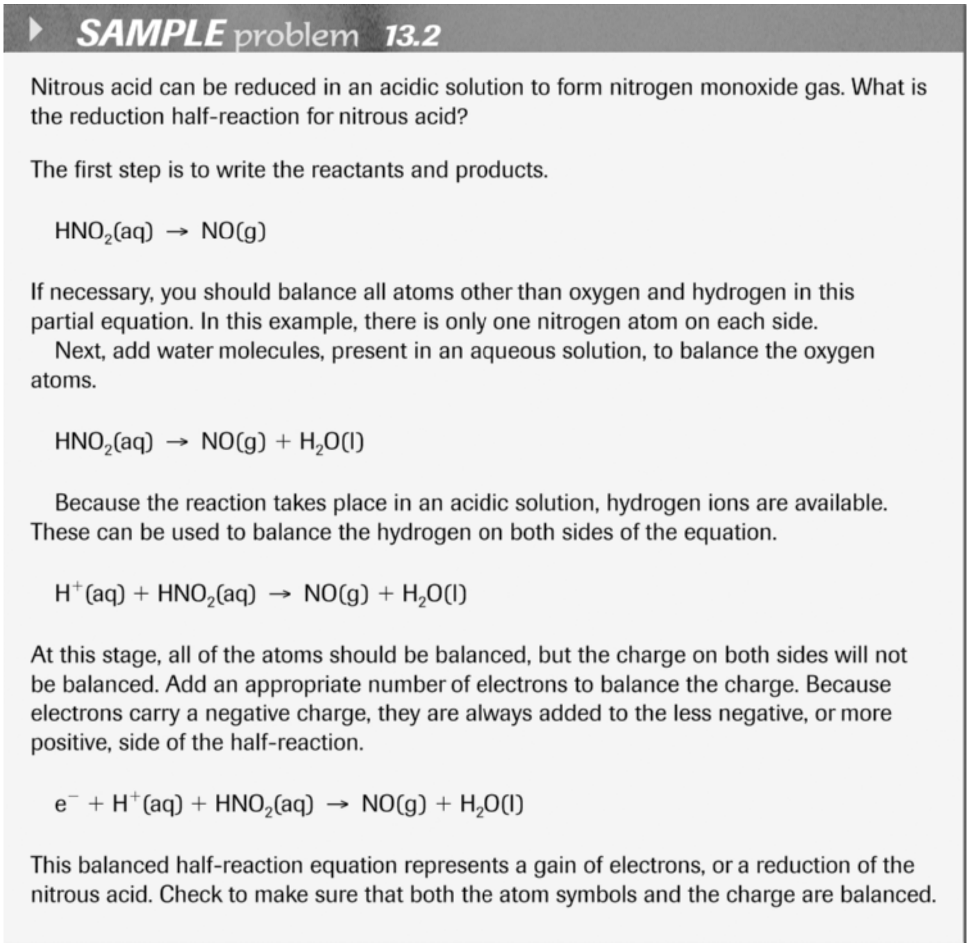
Writing Half-Reaction Equations



SUMMARY Writing Half-Reaction Equations

Step 1: Write the chemical formulas for the reactants and products.
Step 2: Balance all atoms, other than O and H.
Step 3: Balance O by adding $H_2O(l)$.
Step 4: Balance H by adding $H^+(aq)$.
Step 5: Balance the charge on each side by adding e^- and cancel anything that is the same on both sides.

For basic solutions only:
Step 6: Add $OH^-(aq)$ to both sides to equal the number of $H^+(aq)$ present.
Step 7: Combine $H^+(aq)$ and $OH^-(aq)$ on the same side to form $H_2O(l)$. Cancel equal amounts of $H_2O(l)$ from both sides.



SAMPLE problem 13.2

Nitrous acid can be reduced in an acidic solution to form nitrogen monoxide gas. What is the reduction half-reaction for nitrous acid?

The first step is to write the reactants and products.

$$HNO_2(aq) \rightarrow NO(g)$$

If necessary, you should balance all atoms other than oxygen and hydrogen in this partial equation. In this example, there is only one nitrogen atom on each side.

Next, add water molecules, present in an aqueous solution, to balance the oxygen atoms.

$$HNO_2(aq) \rightarrow NO(g) + H_2O(l)$$

Because the reaction takes place in an acidic solution, hydrogen ions are available. These can be used to balance the hydrogen on both sides of the equation.

$$H^+(aq) + HNO_2(aq) \rightarrow NO(g) + H_2O(l)$$

At this stage, all of the atoms should be balanced, but the charge on both sides will not be balanced. Add an appropriate number of electrons to balance the charge. Because electrons carry a negative charge, they are always added to the less negative, or more positive, side of the half-reaction.

$$e^- + H^+(aq) + HNO_2(aq) \rightarrow NO(g) + H_2O(l)$$

This balanced half-reaction equation represents a gain of electrons, or a reduction of the nitrous acid. Check to make sure that both the atom symbols and the charge are balanced.

Balancing Redox Equations with Half-Reaction Equations, Extra Exercises

Balance the following reaction equations.

1. $HNO_2(aq) + In^{+}(aq) \rightarrow NO(g) + In^{3+}(aq)$ (acidic solution)

2. $TiO_2(s) + V^{2+}(aq) \rightarrow Ti(s) + V^{3+}(aq)$ (acidic solution)

3. $BrO_3^{-}(aq) + H_2C_2O_4(aq) \rightarrow Br^{-}(aq) + CO_2(g)$ (acidic solution)

4. $MnO_4^{-}(aq) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + MnO_2(s)$ (basic solution)

5. $OF_2(aq) + I^{-}(aq) \rightarrow F^{-}(aq) + I_3^{-}(aq)$ (acidic solution)

6. $CH_3OH(aq) + MnO_4^{-}(aq) \rightarrow MnO_4^{2-}(aq) + CO_3^{2-}(aq)$ (basic solution)

7. $Hg(l) + H_3IO_6^{2-} \rightarrow Hg_2O(s) + IO_3^{-}(aq)$ (basic solution)

Creating a Redox Table

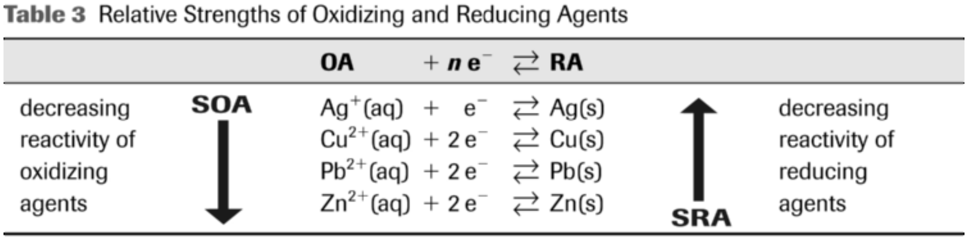


Table 3 Relative Strengths of Oxidizing and Reducing Agents

		OA + $n\,e^-$ ⇄ RA		
decreasing reactivity of oxidizing agents	SOA ↓	$Ag^+(aq) + e^- \rightleftarrows Ag(s)$	↑ SRA	decreasing reactivity of reducing agents
		$Cu^{2+}(aq) + 2\,e^- \rightleftarrows Cu(s)$		
		$Pb^{2+}(aq) + 2\,e^- \rightleftarrows Pb(s)$		
		$Zn^{2+}(aq) + 2\,e^- \rightleftarrows Zn(s)$		

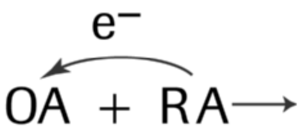
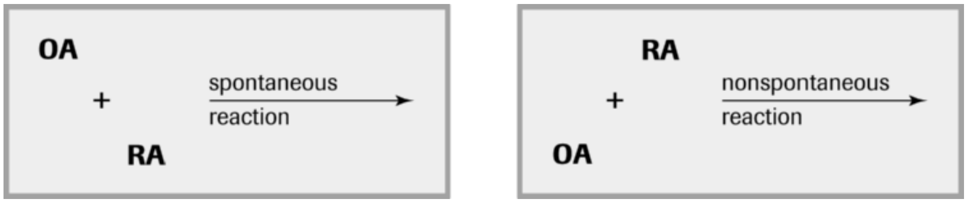
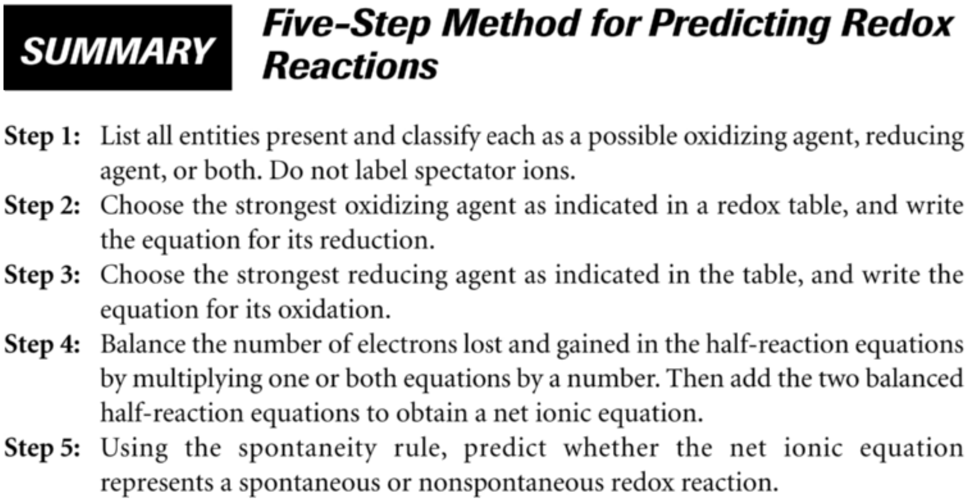


Figure 3
In all redox reactions, electrons are transferred from a reducing agent (RA) to an oxidizing agent (OA).

The redox spontaneity rule





SUMMARY **Five-Step Method for Predicting Redox Reactions**

Step 1: List all entities present and classify each as a possible oxidizing agent, reducing agent, or both. Do not label spectator ions.

Step 2: Choose the strongest oxidizing agent as indicated in a redox table, and write the equation for its reduction.

Step 3: Choose the strongest reducing agent as indicated in the table, and write the equation for its oxidation.

Step 4: Balance the number of electrons lost and gained in the half-reaction equations by multiplying one or both equations by a number. Then add the two balanced half-reaction equations to obtain a net ionic equation.

Step 5: Using the spontaneity rule, predict whether the net ionic equation represents a spontaneous or nonspontaneous redox reaction.

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Student Worksheet LSM13.2D

Creating a Redox Table, Extra Exercises

1. The following equations represent spontaneous reactions. From this evidence, set up a table of relative strengths of oxidizing and reducing agents. Write half-reaction equations and label the strongest oxidizing agent and reducing agent.

$Co(s) + Pd^{2+}(aq) \xrightarrow{\text{spont.}} Co^{2+}(aq) + Pd(s)$

$Pd(s) + Pt^{2+}(aq) \xrightarrow{\text{spont.}} Pd^{2+}(aq) + Pt(s)$

$Mg(s) + Co^{2+}(aq) \xrightarrow{\text{spont.}} Mg^{2+}(aq) + Co(s)$

2. The following equations are interpretations of the evidence from the reactions of four metals with various cation solutions. Make a table of half-reactions and arrange the four metallic ions and the hydrogen ion in order of their decreasing tendency to react. Label the strongest oxidizing agent and reducing agent.

$Cd(s) + 2\ H^{+}(aq) \xrightarrow{\text{spont.}} Cd^{2+}(aq) + H_2(g)$

$Hg(s) + 2\ H^{+}(aq) \xrightarrow{\text{spont.}} Hg^{2+}(aq) + H_2(g)$

$Be(s) + Cd^{2+}(aq) \xrightarrow{\text{spont.}} Be^{2+}(aq) + Cd(s)$

$Ca^{2+}(aq) + Be(s) \xrightarrow{\text{nonspont.}} Ca(s) + Be^{2+}(aq)$

3. Four metals were placed into test tubes containing various ion solutions. Their resulting behaviour is communicated by the equations below. Create a half-reaction table and order the metallic ions and the hydrogen ion according to their tendency to react. Label the strongest oxidizing agent and reducing agent.

   $Pt(s) + 2\ H^+(aq) \xrightarrow{\text{nonspont.}} Pt^{2+}(aq) + H_2(g)$

   $2\ Ce(s) + 3\ Ni^{2+}(aq) \xrightarrow{\text{spont.}} 2\ Ce^{3+}(aq) + 3\ Ni(s)$

   $3\ Sr(s) + 2\ Ce^{3+}(aq) \xrightarrow{\text{spont.}} 3\ Sr^{2+}(aq) + 2\ Ce(s)$

   $Ni(s) + 2\ H^+(aq) \xrightarrow{\text{spont.}} Ni^{2+}(aq) + H_2(g)$

4. A student is required to store an aqueous solution of iron(III) nitrate. She has a choice of a copper, tin, iron, or silver container. Use the table of relative strengths of oxidizing and reducing agents in Appendix I and an appropriate generalization to predict which container would be most suitable for storing the solution.

5. An analytical chemist reacts an unknown metal X with a copper(II) sulfate solution, plating out copper metal. Metal X does not react with aqueous zinc nitrate. What is the order for these metal ions in decreasing tendency to react? What groups of metals are eliminated as a possible identity of the unknown metal? What other solutions might next be chosen to help identify the unknown metal?

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Student Worksheet LSM13.2F

Predicting Redox Reactions, Extra Exercises

For each of the following questions, use the five-step method to predict and communicate the most likely redox reaction, and use the spontaneity rule to predict whether the reaction will occur spontaneously.

1. An aqueous solution of potassium permanganate was reacted with an acidic solution of sodium bromide and an orange–brown substance was formed.

2. A strip of silver metal is placed in a solution of aqueous nickel(II) chloride.

3. Liquid mercury is mixed with a paste of acidic manganese(IV) oxide.

4. Hydrogen peroxide and silver nitrate are mixed.

5. Potassium metal is placed into some water.

6. In a car battery, lead and lead(IV) oxide electrodes are exposed to a sulfuric acid electrolyte. (Assume that the sulfuric acid ionizes to hydrogen and sulfate ions.)

Instructional Master LSM13.3A

Oxidation Number

Table 1 Common Oxidation Numbers

Atom or ion	Oxidation number	Examples
all atoms in elements	0	Na is 0 Cl in Cl_2 is 0
hydrogen in all compounds, except hydrogen in hydrides	+1 −1	H in HCl is +1 H in LiH is −1
oxygen in all compounds, except oxygen in peroxides	−2 −1	O in H_2O is −2 O in H_2O_2 is −1
all monatomic ions	charge on ion	Na^+ is +1 S^{2-} is −2

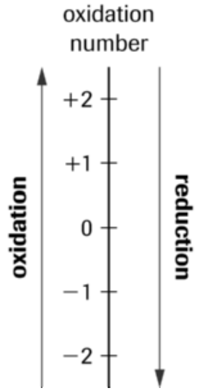

Figure 3
In a redox reaction, both oxidation and reduction occur.

Learning Tip

Oxidation numbers are simply positive or negative numbers assigned on the basis of a set of arbitrary rules. It is important for you to realize that these are not electric charges. For this reason, chemists use the term *oxidation number*. For example, we assign oxidation numbers of −2 and +1 to the oxygen and hydrogen atoms in a water molecule.

SUMMARY Determining Oxidation Numbers

Step 1: Assign common oxidation numbers (**Table 1** on page 583).

Step 2: The total of the oxidation numbers of atoms in a molecule or ion equals the value of the net electric charge on the molecule or ion.

- (a) The sum of the oxidation numbers for a compound is zero.
- (b) The sum of the oxidation numbers for a polyatomic ion equals the charge on the ion.

Step 3: Any unknown oxidation number is determined algebraically from the sum of the known oxidation numbers and the net charge on the entity.

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Student Worksheet LSM13.3B

Oxidation Numbers, Extra Exercises

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1. Assign oxidation numbers to chlorine in each of the following chemicals.

   $HCl(aq)$, $Cl_2(g)$, $NaClO(s)$, $Cl^-(aq)$, $HClO_3(aq)$,

   $ClO_3^-(aq)$, $KClO_2(s)$, $ClO_2(g)$, $HClO_4(aq)$

2. Assign oxidation numbers to manganese in each of the following chemicals.

   $MnO_2(s)$, $KMnO_4(s)$, $Mn(s)$, $MnO_4^{2-}(aq)$, $MnCl_2(s)$,

   $Mn_2O7(s)$, $Mn^{2+}(aq)$

For the following reaction equations, use oxidation numbers to identify the oxidation (the atom oxidized) and the reduction (the atom reduced).

3. $AsO_3^{3-}(aq) + IO_3^-(aq) \rightarrow AsO_4^{3-}(aq) + I^-(aq)$

4. $CuO(s) + NH_3(g) \rightarrow N_2(g) + H_2O(l) + Cu(s)$

5. $MnO_4^-(aq) + H_2Se(g) + H^+(aq) \rightarrow Se(s) + Mn^{2+}(aq) + H_2O(l)$

6. $PbO_2(s) + Pb(s) + H_2SO_4(aq) \rightarrow PbSO_4(s) + H_2O(l)$

7. $Cl_2(g) + OH^-(aq) \rightarrow ClO_3^-(aq) + Cl^-(aq) + H_2O(l)$

Balancing Redox Equations Using Oxidation Numbers

SUMMARY Balancing Redox Equations Using Oxidation Numbers

Step 1: Assign oxidation numbers and identify the atoms/ions whose oxidation numbers change.

Step 2: Using the change in oxidation numbers, write the number of electrons transferred per atom.

Step 3: Using the chemical formulas, determine the number of electrons transferred per reactant. (Use the formula subscripts to do this.)

Step 4: Calculate the simplest whole number coefficients for the reactants that will balance the total number of electrons transferred. Balance the reactants and products.

Step 5: Balance the O atoms using $H_2O(l)$, and then balance the H atoms using $H^+(aq)$.

For basic solutions only

Step 6: Add $OH^-(aq)$ to both sides equal in number to the number of $H^+(aq)$ present.

Step 7: Combine $H^+(aq)$ and $OH^-(aq)$ on the same side to form $H_2O(l)$, and cancel the same number of $H_2O(l)$ on both sides.

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Student Worksheet LSM 13.3E

Balancing Redox Equations Using Oxidation Numbers, Extra Exercises

Balance the following reaction equations.

1. $Sn^{2+}(aq) + Fe^{3+}(aq) \rightarrow Fe^{2+}(aq) + Sn^{4+}(aq)$ (neutral solution)

2. $CH_3OH(aq) + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + CH_2O(aq)$ (acidic solution)

3. $Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$ (acidic solution)

4. $CH_3OH(aq) + MnO_4^-(aq) \rightarrow MnO_4^{2-}(aq) + CO_3^{2-}(aq)$ (basic solution)

5. $MnO_4^-(aq) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + MnO_2(s)$ (basic solution)

Student Worksheet LSM 13.CS

Chapter Checklist

Have you mastered the concepts, applications, and skills associated with the following items? Check them off when you are confident in your understanding.

Knowledge

define oxidation and reduction operationally (historically) and theoretically (13.1, 13.2, 13.3)	
define the following terms: oxidizing agent, reducing agent, oxidation number, half-reaction, disproportionation (13.1, 13.2, 13.3)	
differentiate between redox reactions and other reactions by identifying half-reactions and changes in oxidation number (13.1, 13.2, 13.3)	
identify electron transfer, oxidizing agents, and reducing agents in redox reactions that occur in everyday life in both living and nonliving systems (all sections)	
compare the relative strengths of oxidizing and reducing agents from empirical data (13.2)	
predict the spontaneity of a redox reaction based on a redox table, and compare predictions to experimental results (13.2)	
write and balance equations for redox reactions in acidic, basic, and neutral solutions, including disproportionation reactions, by using half-reaction equations, developing simple half-reaction equations, and assigning oxidation numbers (13.2, 13.3, 13.4)	
perform calculations to determine quantities of substances involved in redox titrations (13.4)	

STS

state that a goal of technology is to solve practical problems (all sections)	
recognize that technological problems may require various solutions and have both intended and unintended consequences (13.1, 13.2, 13.3)	

Skills

initiating and planning	design an experiment to determine the reactivity of various metals (13.1, 13.2)	
	describe procedures for safe handling, storing, and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information (13.1, 13.2, 13.4)	
performing and recording	select and correctly use the appropriate equipment to perform a redox titration (13.4)	
	use a standard reduction potential table as a tool in predicting the spontaneity of redox reactions and their products (13.2, 13.4)	
	create charts, tables or spreadsheets that present the results of redox experiments (13.1, 13.2, 13.4)	
analyzing and interpreting	analyze evidence from an experiment to derive a simple reduction table (13.2)	
	interpret patterns and trends in data derived from redox reactions (all sections)	
	evaluate redox experiments, including identifying the limitations of the evidence (13.1, 13.2, 13.4)	

communication and teamwork	work collaboratively in addressing problems, and select and use appropriate modes of representation for redox reactions and answers to problems related to redox titrations (all sections)	

Key Terms

13.1	half-reaction	
	reduction	
	oxidation	
	redox reaction	
13.2	reducing agent	
	oxidizing agent	
	redox spontaneity rule	
	disproportionation	
13.3	oxidation number	
	oxidation	
	reduction	

Key STS

identify solutions to practical problems in early metallurgy, various cleaners, food preservation, and medical analyses *(Exploration: Cleaning Silver; CS: Early Metallurgy; Investigation 13.4: Analyzing a Hydrogen Peroxide Solution)*	
analyze the implications of redox processes in various applications, such as pulp and paper, corrosion, antioxidants, and blood analysis technologies *(WA: Piercings—A Rash Decision; WA Catalytic Converters; CS: Bleaching Wood Pulp; WA: Imants Lauks)*	
research a variety of careers *(CC: Materials/Metallurgical Engineer; WA: Henry Taube and Rudolph Marcus; CC: Conservator; WA: Imants Lauks)*	

Key Skills

create charts, tables, or spreadsheets that present the results of redox experiments *(Investigation 13.1: Single Replacement Reactions; Investigation 13.2 Spontaneity of Redox Reactions; LE 13.A Building a Redox Table; Investigation 13.3: Predicting the Reaction of Sodium Metal; Investigation 13.4: Analyzing a Hydrogen Peroxide Solutions)*	
interpret patterns and trends in data derived from redox reactions *(Investigation 13.1: Single Replacement Reactions; WA: Redox Reaction; Investigation 13.2 Spontaneity of Redox Reactions; LE 13.A Building a Redox Table; Investigation 13.3: Predicting the Reaction of Sodium Metal; LE 13.B: Oxidation States of Vanadium; LE 13.C: Analyzing for Tin; LE 13.D: Analyzing for Chromium in Steel; Investigation 13.4: Analyzing a Hydrogen Peroxide Solutions)*	
identify the limitations of data collected *(Investigation 13.1: Single Replacement Reactions; Investigation 13.2 Spontaneity of Redox Reactions; LE 13.A Building a Redox Table; Investigation 13.3: Predicting the Reaction of Sodium Metal; Investigation 13.4: Analyzing a Hydrogen Peroxide Solutions)*	

Instructional Master LSM 14.1B

Electric Cells

Figure 1
A version of Volta's first battery. Each beaker contains two different metals, copper and zinc, in an electrolyte, salt water. A series of beakers forms a series of cells (a battery) whose total voltage is the sum of the individual voltages of all cells.

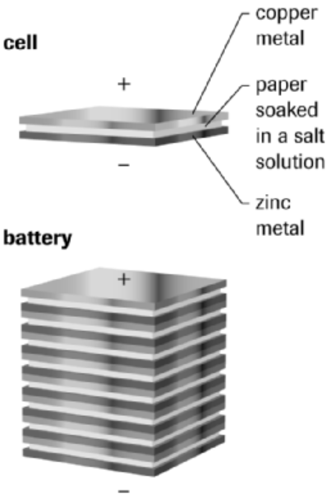

Figure 2
Volta's revised cell design, simpler than the first, consisted of a sandwich of two metals separated by paper soaked in salt water (the electrolyte). A cell consisted of a layer of zinc metal separated from a layer of copper metal by the brine-soaked paper. A large pile of cells could be constructed to give more electrical energy per unit charge.

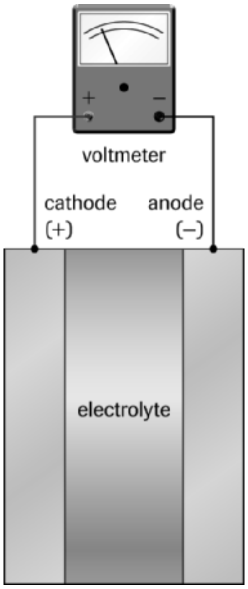

Figure 3
An electric cell always contains two electrodes—an anode and a cathode—and an electrolyte. When testing the voltage of a cell or battery, the red (+) lead of the voltmeter is connected to the positive electrode (cathode), and the black (–) lead is connected to the negative electrode (anode).

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Instructional Master LSM 14.1C

Primary, Secondary, and Fuel Cells

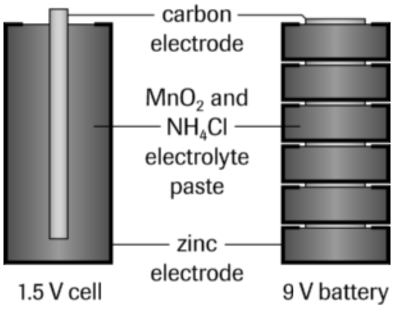
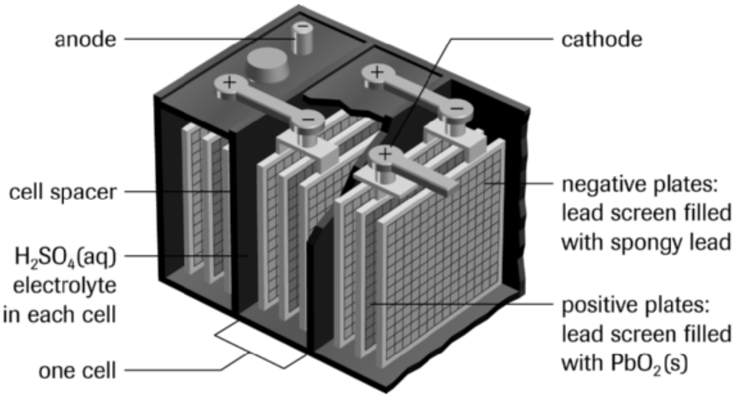


Figure 6
Like a flashlight D cell, the zinc chloride dry cell on the left has a voltage of 1.5 V. The 9 V battery on the right is made up of six 1.5 V dry cells in series.

Figure 7
The anodes of a lead–acid car battery are composed of spongy lead and the cathodes are composed of lead(IV) oxide on a metal screen. The large electrode surface area is designed to deliver sufficient current to start a car engine.



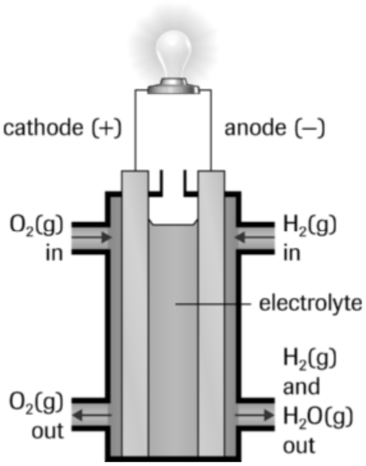
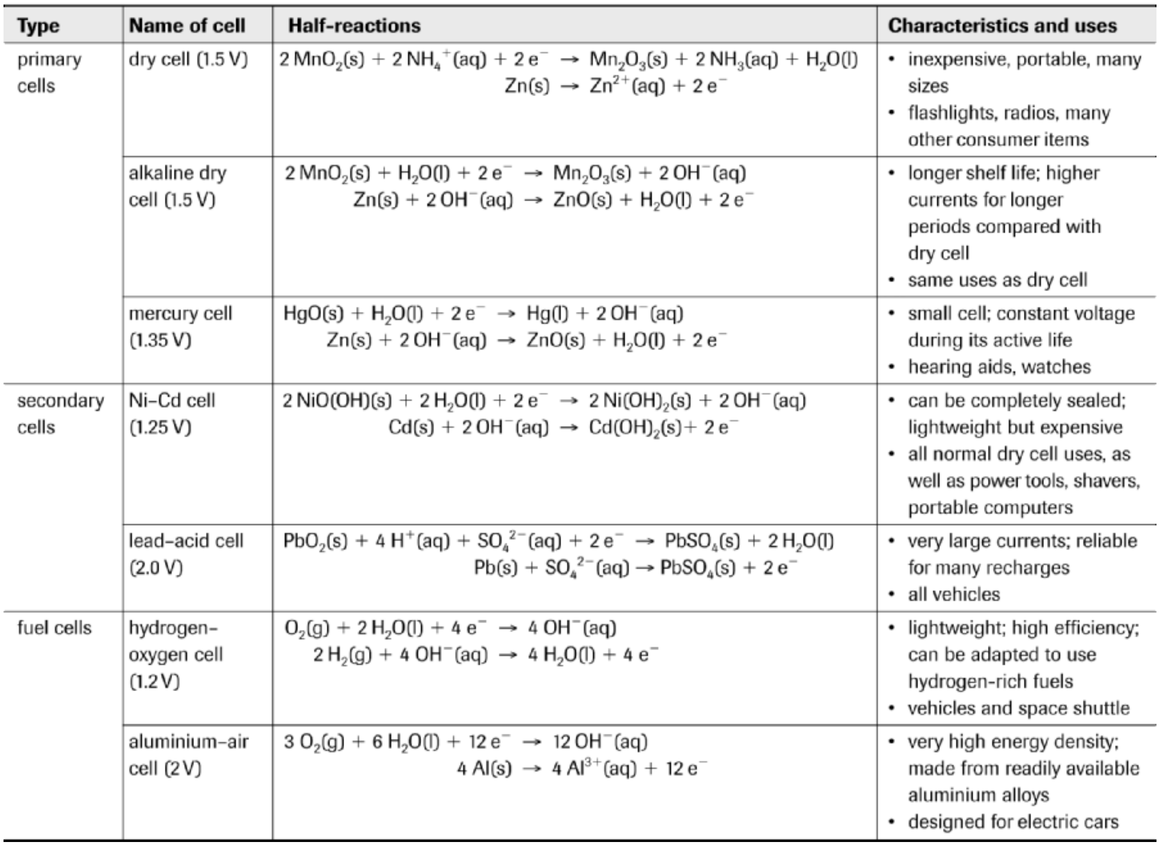


Figure 8
A hydrogen–oxygen fuel cell. Hydrogen and oxygen gases are continuously pumped into the cell, and each reacts at a different electrode. Unused gases are removed, filtered, and then recycled.

Primary, Secondary, and Fuel Cell Characteristics

Table 2 Primary, Secondary, and Fuel Cells



Type	Name of cell	Half-reactions	Characteristics and uses
primary cells	dry cell (1.5 V)	$2\,MnO_2(s) + 2\,NH_4^+(aq) + 2\,e^- \rightarrow Mn_2O_3(s) + 2\,NH_3(aq) + H_2O(l)$ $Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$	• inexpensive, portable, many sizes • flashlights, radios, many other consumer items
	alkaline dry cell (1.5 V)	$2\,MnO_2(s) + H_2O(l) + 2\,e^- \rightarrow Mn_2O_3(s) + 2\,OH^-(aq)$ $Zn(s) + 2\,OH^-(aq) \rightarrow ZnO(s) + H_2O(l) + 2\,e^-$	• longer shelf life; higher currents for longer periods compared with dry cell • same uses as dry cell
	mercury cell (1.35 V)	$HgO(s) + H_2O(l) + 2\,e^- \rightarrow Hg(l) + 2\,OH^-(aq)$ $Zn(s) + 2\,OH^-(aq) \rightarrow ZnO(s) + H_2O(l) + 2\,e^-$	• small cell; constant voltage during its active life • hearing aids, watches
secondary cells	Ni–Cd cell (1.25 V)	$2\,NiO(OH)(s) + 2\,H_2O(l) + 2\,e^- \rightarrow 2\,Ni(OH)_2(s) + 2\,OH^-(aq)$ $Cd(s) + 2\,OH^-(aq) \rightarrow Cd(OH)_2(s) + 2\,e^-$	• can be completely sealed; lightweight but expensive • all normal dry cell uses, as well as power tools, shavers, portable computers
	lead–acid cell (2.0 V)	$PbO_2(s) + 4\,H^+(aq) + SO_4^{2-}(aq) + 2\,e^- \rightarrow PbSO_4(s) + 2\,H_2O(l)$ $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2\,e^-$	• very large currents; reliable for many recharges • all vehicles
fuel cells	hydrogen–oxygen cell (1.2 V)	$O_2(g) + 2\,H_2O(l) + 4\,e^- \rightarrow 4\,OH^-(aq)$ $2\,H_2(g) + 4\,OH^-(aq) \rightarrow 4\,H_2O(l) + 4\,e^-$	• lightweight; high efficiency; can be adapted to use hydrogen-rich fuels • vehicles and space shuttle
	aluminium–air cell (2 V)	$3\,O_2(g) + 6\,H_2O(l) + 12\,e^- \rightarrow 12\,OH^-(aq)$ $4\,Al(s) \rightarrow 4\,Al^{3+}(aq) + 12\,e^-$	• very high energy density; made from readily available aluminium alloys • designed for electric cars

Instructional Master LSM 14.2C

Voltaic Cell Design

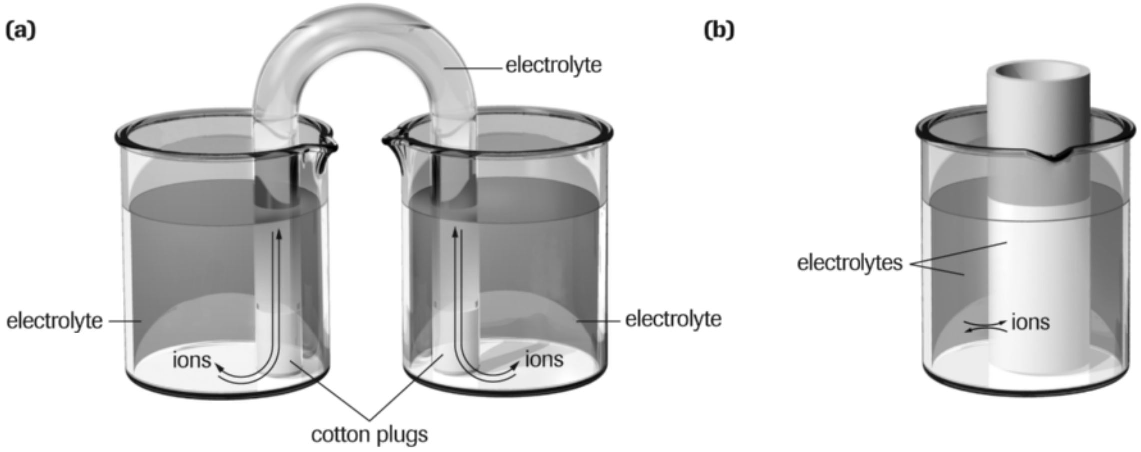


Figure 1
(a) A salt bridge is a U-shaped tube containing an inert (unreactive) aqueous electrolyte such as sodium sulfate.
(b) An unglazed porcelain (porous) cup containing one electrolyte sits in a container of a second electrolyte.

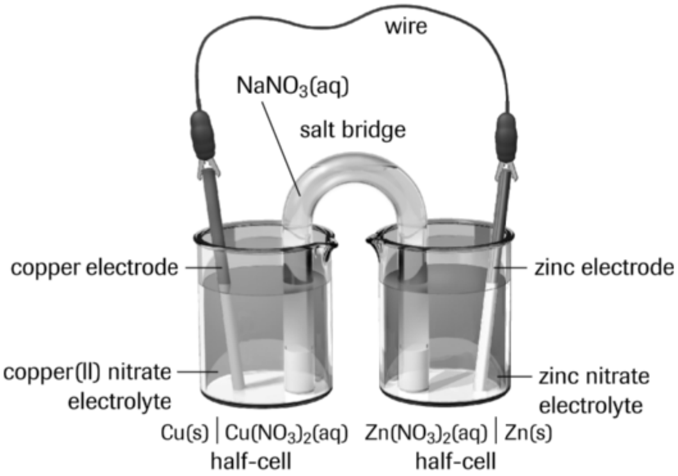


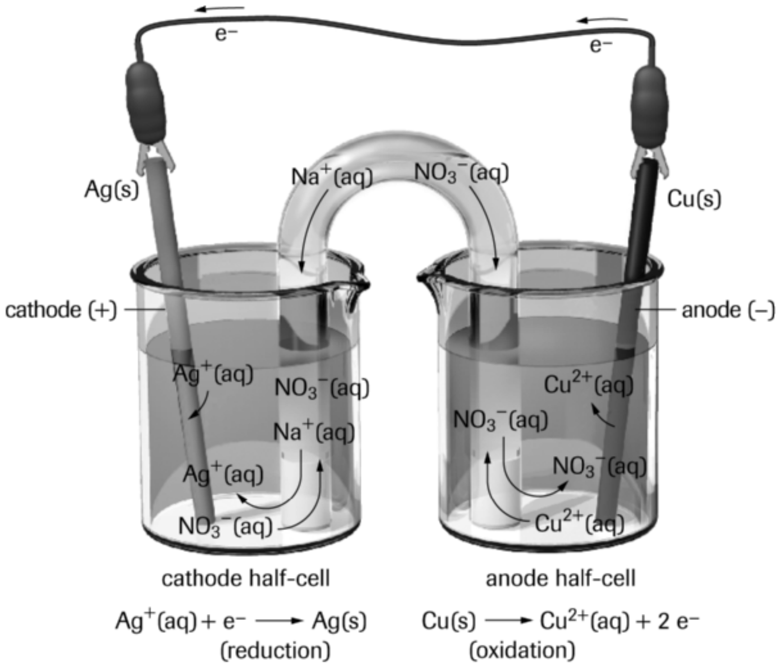
Figure 2
Each electrode is in its own electrolyte, forming a half-cell. The two half-cells are connected by a salt bridge (containing $NaNO_3(aq)$) and by an external wire to make a complete circuit.

Instructional Master LSM 14.2D

Silver–Copper Cell

Table 1 Evidence and Interpretations of the Silver–Copper Cell

Evidence	Interpretation
The copper electrode decreases in size and the intensity of the blue colour of the electrolyte increases.	Oxidation of copper metal is occurring: $\underset{}{\mathrm{Cu(s)}} \rightarrow \underset{\text{blue}}{\mathrm{Cu^{2+}(aq)}} + 2\,e^-$
The silver electrode increases in size as long, silver-coloured crystals grow.	Reduction of silver ions is occurring: $\mathrm{Ag^+(aq)} + e^- \rightarrow \mathrm{Ag(s)}$
A blue colour slowly moves up the U-tube from the copper half-cell to the silver half-cell.	Copper(II) ions (cations) move toward the cathode.
An ammeter shows that the electric current flows along a wire between the copper electrode and the silver electrode.	Electrons move through the wire from the copper electrode to the silver electrode.
A voltmeter indicates that the silver electrode cathode (positive) and the copper electrode is the anode (negative).	Electrons have a tendency to leave the copper half-cell and enter the silver half-cell.

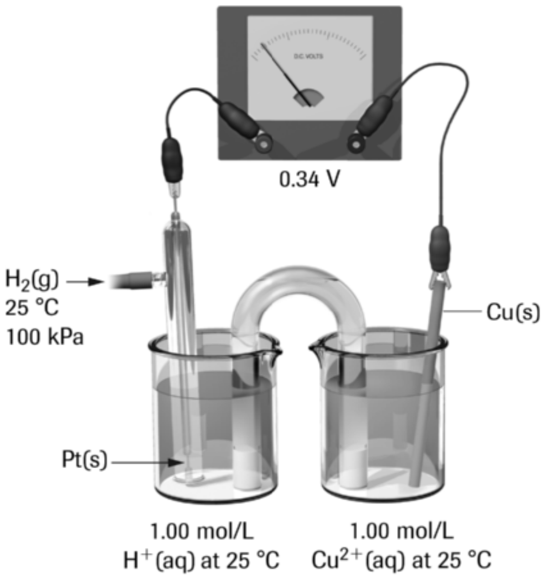


$\mathrm{Ag^+(aq)} + e^- \longrightarrow \mathrm{Ag(s)}$

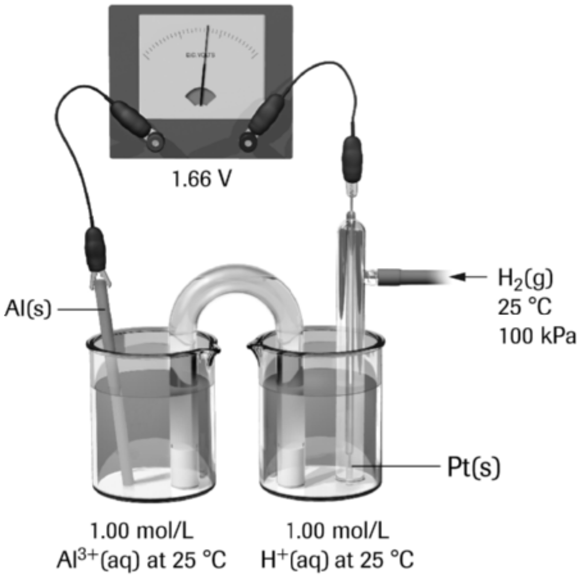
$\mathrm{Cu(s)} \longrightarrow \mathrm{Cu^{2+}(aq)} + 2\,e^-$

Figure 3
A theoretical interpretation of the silver–copper cell

Measuring E_{r}°



A copper–hydrogen standard cell



An aluminium–hydrogen standard cell

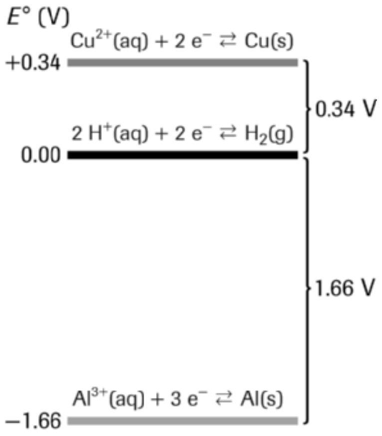


Figure 11
Measurements of standard cell potentials show that the reduction potential of $Cu^{2+}(aq)$ is 0.34 V greater than that of $H^+(aq)$, which is 1.66 V greater than that of $Al^{3+}(aq)$. If you set up a standard cell using copper and aluminium, what would be the cell potential, $E°$cell. (Answer: 2.00 V)

Analyzing Standard Cells

SUMMARY Rules for Analyzing Standard Cells

You can analyze a standard cell knowing the contents of both half-cells using one or more of the following steps:

- Determine which electrode is the cathode. The cathode is the electrode where the strongest oxidizing agent present in the cell reacts, i.e., the oxidizing agent that is closest to the top on the left side of the redox table. If required, copy the reduction half-reaction for the strongest oxidizing agent and its reduction potential.
- Determine which electrode is the anode. The anode is the electrode where the strongest reducing agent present in the cell reacts, i.e., the reducing agent that is closest to the bottom on the right side of the redox table. If required, copy the oxidation half-reaction (reverse the half-reaction by reading from right to left) for the strongest reducing agent and its reduction potential.
- Determine the overall cell reaction. Balance the electrons for the two half-reaction equations (but do not change the E°_r) and add the half-reaction equations.
- Determine the standard cell potential, E°_{cell}, using the equation:

$$E^\circ_{cell} = \underset{\text{cathode}}{E^\circ_r} - \underset{\text{anode}}{E^\circ_r}$$

SUMMARY Standard Cell Potential

- A standard cell is one in which all entities shown in the half-reaction equation are present and at SATP. The concentration of aqueous entities is 1.0 mol/L.
- The standard cell potential, E°_{cell}, is the maximum electric potential difference between the cathode and anode of a voltaic cell at standard conditions.

$$E^\circ_{cell} = \underset{\text{cathode}}{E^\circ_r} - \underset{\text{anode}}{E^\circ_r}$$

- A positive standard cell potential ($E^\circ_{cell} > 0$) indicates that the overall cell reaction is spontaneous.
- The standard reduction potential E°_r, represents the ability of a standard half-cell to attract electrons, relative to the reference half-cell.
- The reference half-cell is $Pt(s) \mid H_2(g) \mid H^+(aq)$, (refer to page 627). It has, by definition, a standard reduction potential of exactly zero volts.

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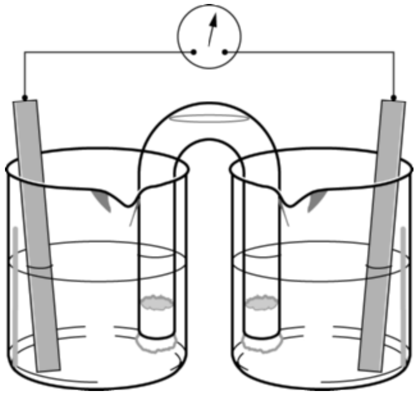
Student Worksheet LSM 14.2G

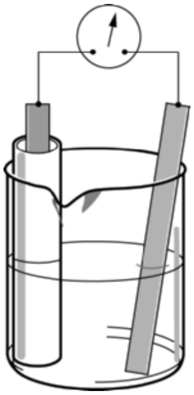
Voltaic Cells, Extra Exercises

For each of the following cells, use the given cell notation to identify the strongest oxidizing and reducing agents. Write chemical equations to represent the cathode, anode, and net cell reactions. Label electrodes, electrolytes, electron flow, and ion movement. Predict the cell potential.

1. Cd(s) | $Cd(NO_3)_2$(aq) || AgNO3(aq) | Ag(s)



2. Pt(s) | IO_3^-(aq), H^+(aq) || Zn^{2+}(aq) | Zn(s)



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Student Worksheet LSM 14.2I

Cell Potentials, Extra Exercises

1. Assume that the reference half-cell is changed to a standard mercury–mercury(II) half-cell.
   (a) What would be the reduction potential of a standard chlorine half-cell?
   (b) What would be the reduction potential of a standard nickel half-cell?
   (c) What would be the net cell potential of a standard chlorine–nickel cell?
   (d) Why is the answer to part c the same as the answer obtained using the standard hydrogen half-cell as the reference?

2. For each of the following standard cells, write the cell notation, label electrodes, and determine the cell potential.
   (a) cobalt–hydrogen standard cell
   (b) zinc–aluminium standard cell
   (c) tin(IV)–zinc standard cell

3. For each of the following cell notations, write the cathode, anode, and net cell reaction equations and calculate the cell potential.
   (a) $Pb(s) \mid Pb^{2+}(aq) \parallel Ni^{2+}(aq) \mid Ni(s)$
   (b) $Pt(s) \mid SO_4^{2-}(aq), H^+(aq), H_2SO_3(aq) \parallel Ag^+(aq) \mid Ag(s)$
   (c) $Cd(s) \mid Cd^{2+}(aq) \parallel ClO_4^-(aq), H^+(aq), Cl^-(aq) \mid C(s)$

4. Use the standard cell described below to determine the standard reduction potential of the gallium half-cell.

   $\underset{\text{cathode}}{\mathrm{Cu(s)}} \mid \mathrm{Cu^{2+}(aq)} \parallel \mathrm{Ga^{3+}(aq)} \mid \underset{\text{anode}}{\mathrm{Ga(s)}} \quad \Delta E° = +0.90\ \mathrm{V}$

Instructional Master LSM 14.2K

Corrosion of Iron

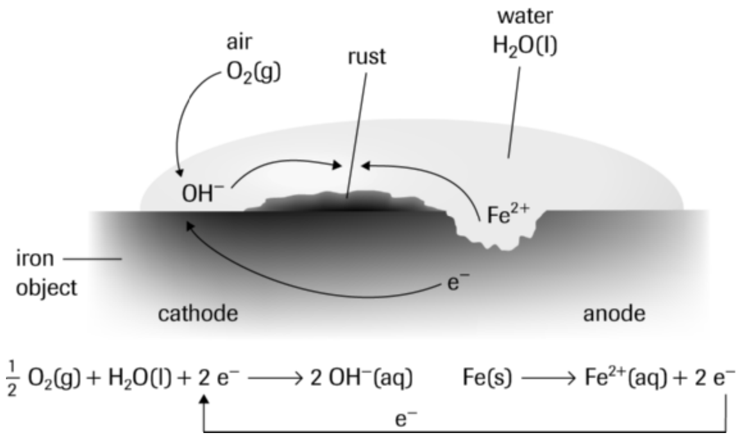


Figure 14
The corrosion of iron is a small electrochemical cell with iron oxidation at one location (the anode) and oxygen reduction at another location (the cathode).

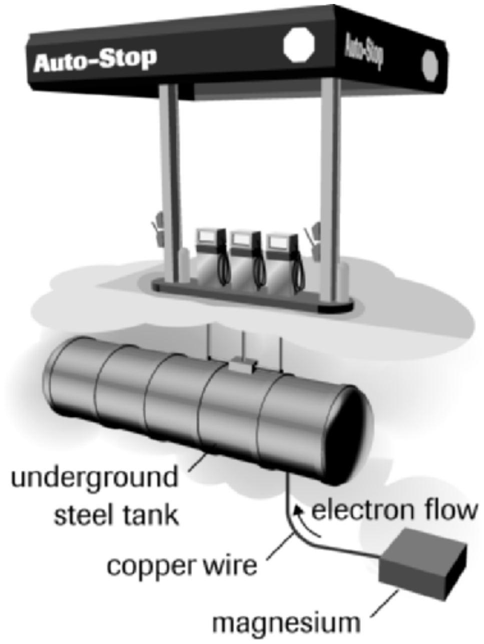


Figure 16
Corrosion of iron involves the oxidation of iron at the anode of a cell. If the iron is attached directly or connected electrically to a metal that is more easily oxidized (a sacrificial anode), then a spontaneous cell develops in which iron is the cathode. The electrolyte of the cell is the moisture in the ground.

Instructional Master LSM 14.3C

Voltaic and Electrolytic Cells

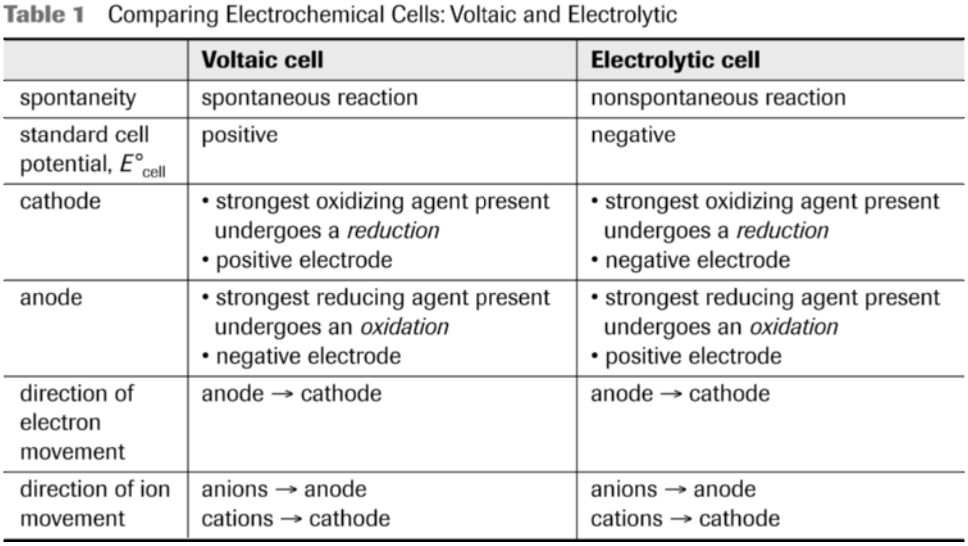
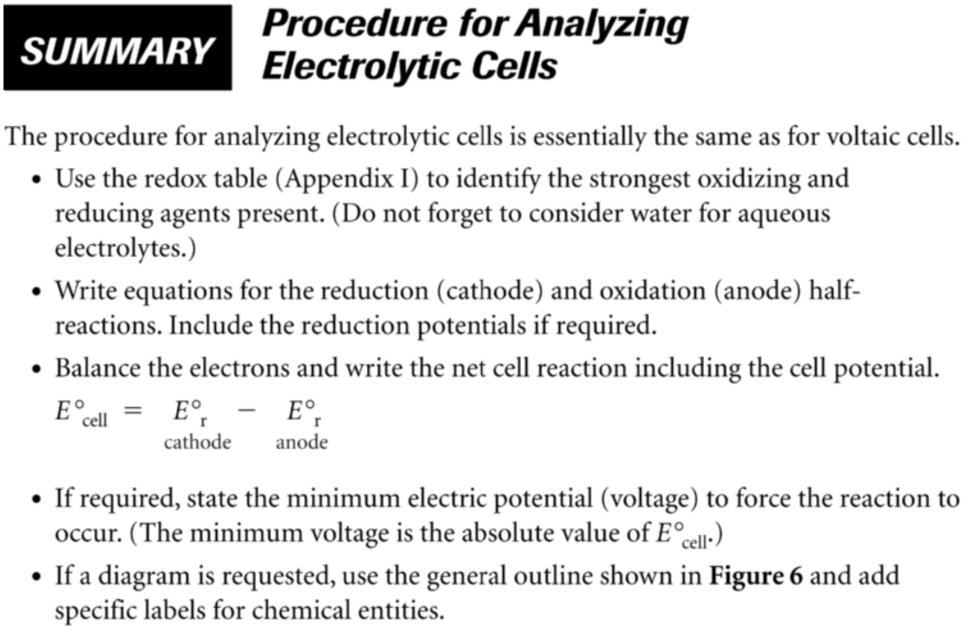


Table 1 Comparing Electrochemical Cells: Voltaic and Electrolytic

	Voltaic cell	**Electrolytic cell**
spontaneity	spontaneous reaction	nonspontaneous reaction
standard cell potential, E°_{cell}	positive	negative
cathode	• strongest oxidizing agent present undergoes a *reduction* • positive electrode	• strongest oxidizing agent present undergoes a *reduction* • negative electrode
anode	• strongest reducing agent present undergoes an *oxidation* • negative electrode	• strongest reducing agent present undergoes an *oxidation* • positive electrode
direction of electron movement	anode → cathode	anode → cathode
direction of ion movement	anions → anode cations → cathode	anions → anode cations → cathode

Instructional Master LSM 14.3D

Analyzing Electrolytic Cells



SUMMARY

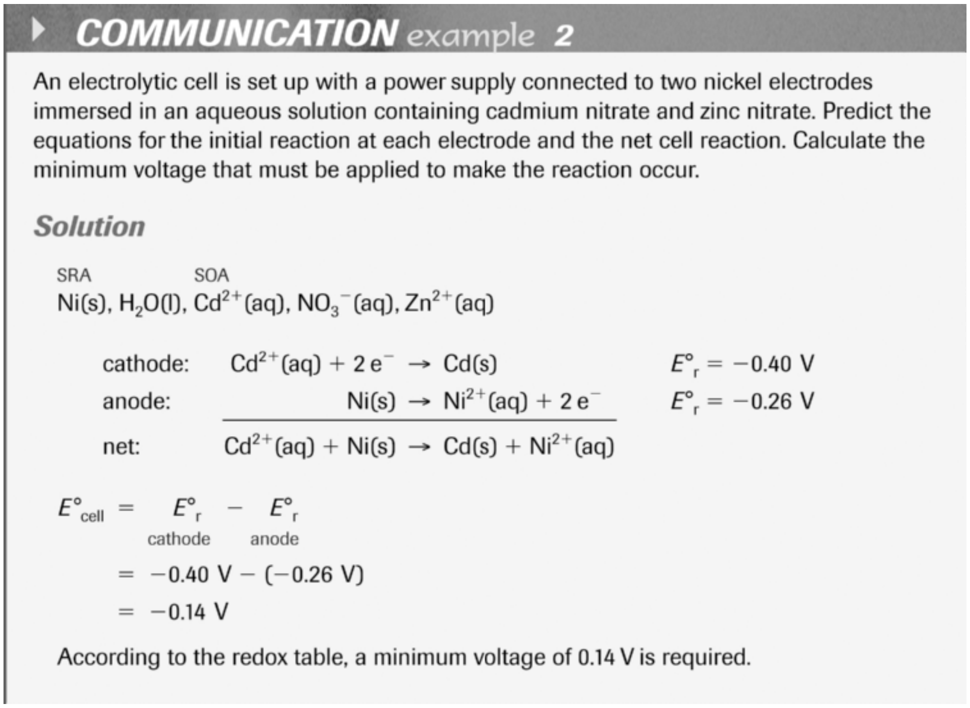
Procedure for Analyzing Electrolytic Cells

The procedure for analyzing electrolytic cells is essentially the same as for voltaic cells.

- Use the redox table (Appendix I) to identify the strongest oxidizing and reducing agents present. (Do not forget to consider water for aqueous electrolytes.)
- Write equations for the reduction (cathode) and oxidation (anode) half-reactions. Include the reduction potentials if required.
- Balance the electrons and write the net cell reaction including the cell potential.

  $$E^\circ_{\text{cell}} = \underset{\text{cathode}}{E^\circ_{\text{r}}} - \underset{\text{anode}}{E^\circ_{\text{r}}}$$

- If required, state the minimum electric potential (voltage) to force the reaction to occur. (The minimum voltage is the absolute value of E°_{cell}.)
- If a diagram is requested, use the general outline shown in **Figure 6** and add specific labels for chemical entities.



COMMUNICATION example 2

An electrolytic cell is set up with a power supply connected to two nickel electrodes immersed in an aqueous solution containing cadmium nitrate and zinc nitrate. Predict the equations for the initial reaction at each electrode and the net cell reaction. Calculate the minimum voltage that must be applied to make the reaction occur.

Solution

$$\underset{}{\overset{\text{SRA}}{\text{Ni(s)}}},\ \text{H}_2\text{O(l)},\ \overset{\text{SOA}}{\text{Cd}^{2+}\text{(aq)}},\ \text{NO}_3^{-}\text{(aq)},\ \text{Zn}^{2+}\text{(aq)}$$

$$\begin{array}{lrcll}
\text{cathode:} & \text{Cd}^{2+}\text{(aq)} + 2\,\text{e}^- & \rightarrow & \text{Cd(s)} & E^\circ_{\text{r}} = -0.40\ \text{V} \\
\text{anode:} & \text{Ni(s)} & \rightarrow & \text{Ni}^{2+}\text{(aq)} + 2\,\text{e}^- & E^\circ_{\text{r}} = -0.26\ \text{V} \\
\hline
\text{net:} & \text{Cd}^{2+}\text{(aq)} + \text{Ni(s)} & \rightarrow & \text{Cd(s)} + \text{Ni}^{2+}\text{(aq)} &
\end{array}$$

$$\begin{aligned}
E^\circ_{\text{cell}} &= \underset{\text{cathode}}{E^\circ_{\text{r}}} - \underset{\text{anode}}{E^\circ_{\text{r}}} \\
&= -0.40\ \text{V} - (-0.26\ \text{V}) \\
&= -0.14\ \text{V}
\end{aligned}$$

According to the redox table, a minimum voltage of 0.14 V is required.

Student Worksheet LSM 14.3E

Electrolytic Cells, Extra Exercises

1. Write the equations for the anode, cathode, and net reactions for the electrolysis of molten aluminium oxide. One half-reaction will involve conversion of liquid oxide ions into oxygen gas.

2. Write equations for the anode, cathode, and net reactions for the electrolysis of molten potassium bromide.

3. For the following electrolytic cells, write equations for half-reactions and the net cell reaction, and calculate the minimum electrical potential difference that would be required to cause the electrolysis to occur.

   (a) an aqueous solution of nickel (II) chloride is electrolyzed

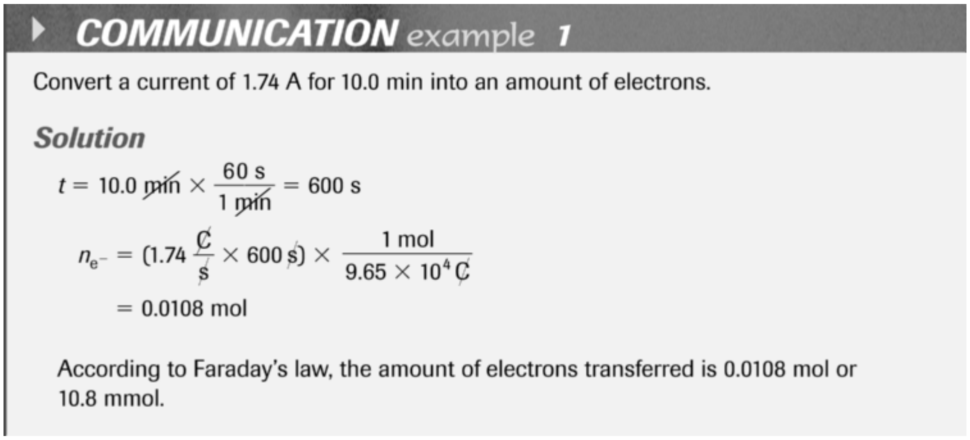
   (b) an aqueous solution of potassium sulfate is electrolyzed

   (c) an aqueous solution of lead (II) nitrate is electrolyzed

Faraday's Constant

$$F = 9.65 \times 10^4 \, \frac{\text{C}}{\text{mol e}^-}$$

$$n_{\text{e}^-} = \frac{It}{F}$$



COMMUNICATION *example* **1**

Convert a current of 1.74 A for 10.0 min into an amount of electrons.

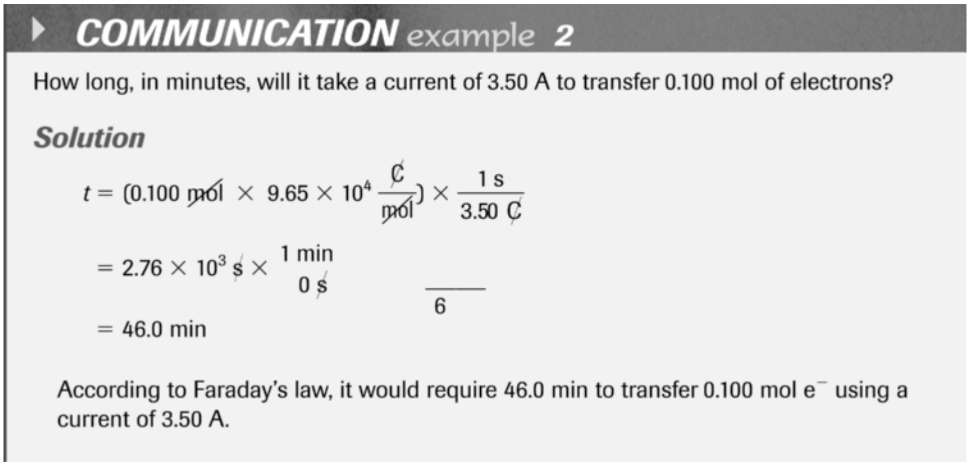
Solution

$$t = 10.0\ \cancel{\text{min}} \times \frac{60\ \text{s}}{1\ \cancel{\text{min}}} = 600\ \text{s}$$

$$n_{\text{e}^-} = \left(1.74\ \frac{\cancel{\text{C}}}{\cancel{\text{s}}} \times 600\ \cancel{\text{s}}\right) \times \frac{1\ \text{mol}}{9.65 \times 10^4\ \cancel{\text{C}}}$$

$$= 0.0108\ \text{mol}$$

According to Faraday's law, the amount of electrons transferred is 0.0108 mol or 10.8 mmol.



COMMUNICATION *example* **2**

How long, in minutes, will it take a current of 3.50 A to transfer 0.100 mol of electrons?

Solution

$$t = \left(0.100\ \cancel{\text{mol}} \times 9.65 \times 10^4\ \frac{\cancel{\text{C}}}{\cancel{\text{mol}}}\right) \times \frac{1\ \text{s}}{3.50\ \cancel{\text{C}}}$$

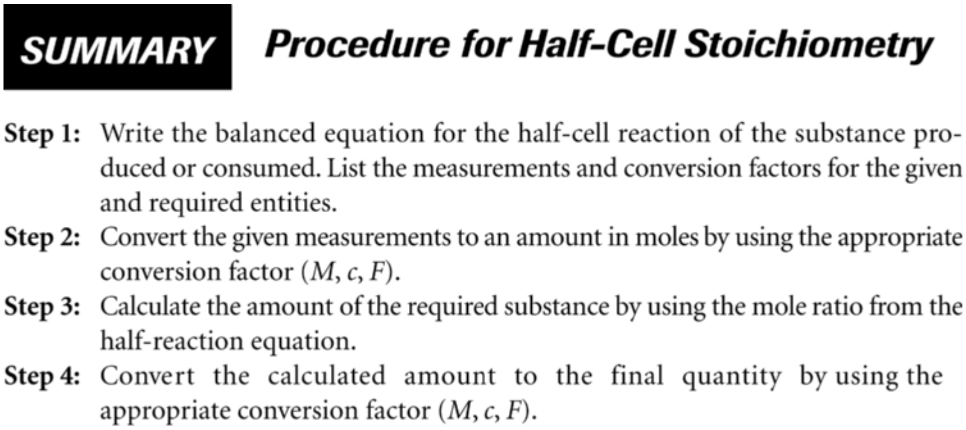
$$= 2.76 \times 10^3\ \cancel{\text{s}} \times \frac{1\ \text{min}}{0\ \cancel{\text{s}}}$$

6

$$= 46.0\ \text{min}$$

According to Faraday's law, it would require 46.0 min to transfer 0.100 mol e^- using a current of 3.50 A.

Half-Cell Stoichiometry



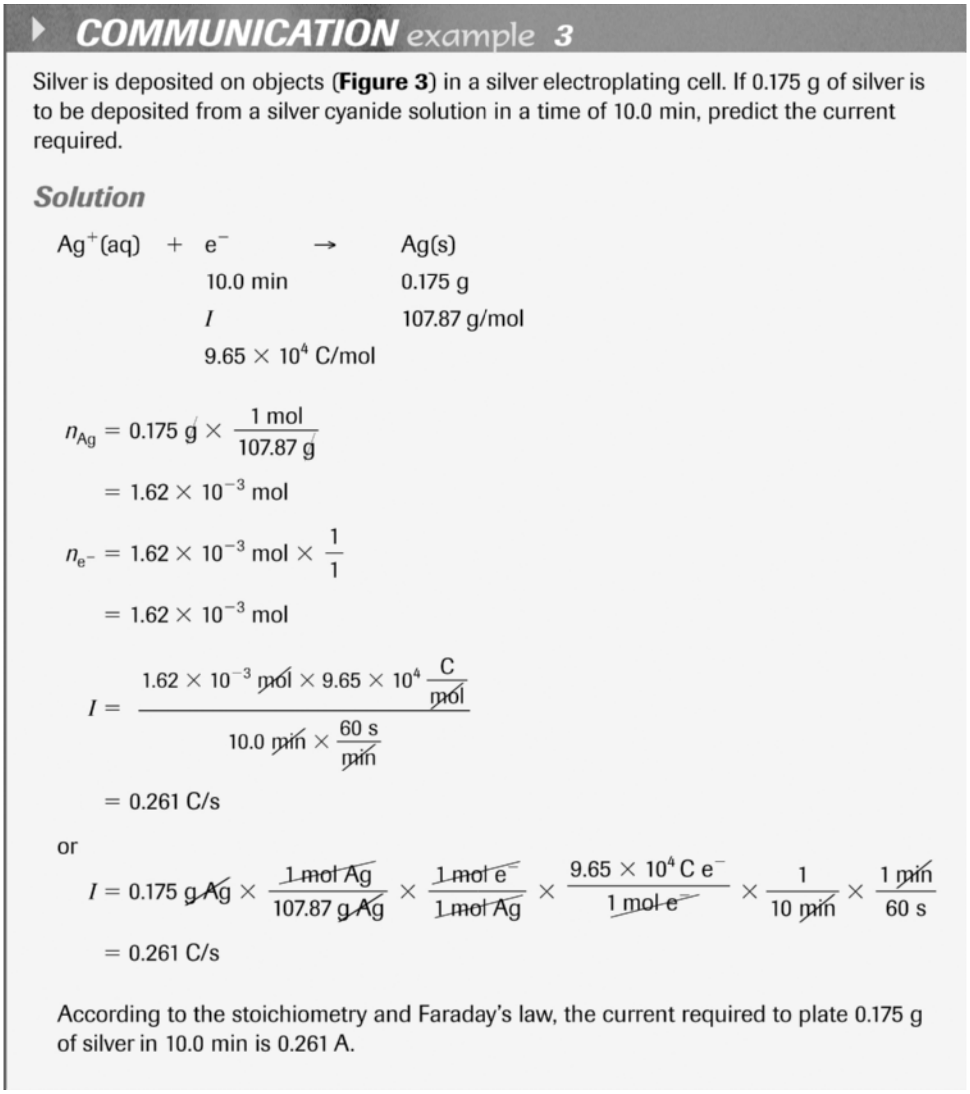
SUMMARY Procedure for Half-Cell Stoichiometry

Step 1: Write the balanced equation for the half-cell reaction of the substance produced or consumed. List the measurements and conversion factors for the given and required entities.

Step 2: Convert the given measurements to an amount in moles by using the appropriate conversion factor (M, c, F).

Step 3: Calculate the amount of the required substance by using the mole ratio from the half-reaction equation.

Step 4: Convert the calculated amount to the final quantity by using the appropriate conversion factor (M, c, F).



COMMUNICATION example 3

Silver is deposited on objects (**Figure 3**) in a silver electroplating cell. If 0.175 g of silver is to be deposited from a silver cyanide solution in a time of 10.0 min, predict the current required.

Solution

$Ag^+(aq)$	+	e^-	→	$Ag(s)$
		10.0 min		0.175 g
		I		107.87 g/mol
		9.65×10^4 C/mol		

$$n_{Ag} = 0.175\ \cancel{g} \times \frac{1\ \text{mol}}{107.87\ \cancel{g}}$$

$$= 1.62 \times 10^{-3}\ \text{mol}$$

$$n_{e^-} = 1.62 \times 10^{-3}\ \text{mol} \times \frac{1}{1}$$

$$= 1.62 \times 10^{-3}\ \text{mol}$$

$$I = \frac{1.62 \times 10^{-3}\ \cancel{\text{mol}} \times 9.65 \times 10^4\ \frac{\text{C}}{\cancel{\text{mol}}}}{10.0\ \cancel{\text{min}} \times \frac{60\ \text{s}}{\cancel{\text{min}}}}$$

$$= 0.261\ \text{C/s}$$

or

$$I = 0.175\ \cancel{\text{g Ag}} \times \frac{\cancel{1\ \text{mol Ag}}}{107.87\ \cancel{\text{g Ag}}} \times \frac{\cancel{1\ \text{mol e}^-}}{\cancel{1\ \text{mol Ag}}} \times \frac{9.65 \times 10^4\ \text{C e}^-}{\cancel{1\ \text{mol e}^-}} \times \frac{1}{10\ \cancel{\text{min}}} \times \frac{1\ \cancel{\text{min}}}{60\ \text{s}}$$

$$= 0.261\ \text{C/s}$$

According to the stoichiometry and Faraday's law, the current required to plate 0.175 g of silver in 10.0 min is 0.261 A.

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Student Worksheet LSM 14.4C

Stoichiometry of Cell Reactions, Extra Exercises

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1. A student wishes to set up an electrolytic cell to plate copper onto a belt buckle. Predict the length of time it will take to plate out 2.5 g of copper from a copper(II) nitrate solution using a 2.5 A current. At which electrode should the buckle be attached?

2. Determine the mass of chlorine produced when a 200 A current flows for 24.0 h through a cell containing molten sodium chloride (a Downs cell). At which electrode is the chlorine produced?

3. How long would it take a 500 A current to produce 1.00 kg of aluminium from aluminium oxide dissolved in molten cryolite (Hall–Héroult cell)?

4. A trophy company is setting up a nickel-plating cell using an electrolyte containing nickel(II) ions. Predict the current required to produce nickel metal at the rate of 5.00 g/min.

Student Worksheet LSM 14.CS

Chapter Checklist

Have you mastered the concepts, applications, and skills associated with the following items? Check them off when you are confident in your understanding.

Knowledge

define anode, cathode, anion, cation, salt bridge/porous cup, electrolyte, voltaic cell, and electrolytic cell (14.2, 14.3)	
predict and write the half-reaction equation that occurs at each electrode in an electrochemical cell (14.2, 14.3, 14.4)	
explain that the values of standard reduction potential are all relative to E°_{r} = 0.00 V set for the hydrogen electrode at standard conditions (14.2)	
calculate the standard cell potential for electrochemical cells (14.2, 14.3)	
predict the spontaneity of redox reactions based on standard cell potentials (14.2, 14.3)	
identify the similarities and differences between a voltaic cell and an electrolytic cell (14.3)	
recognize that predicted reactions do not always occur (14.3)	
calculate mass, amounts, current, and time in single voltaic and electrolytic cells by applying Faraday's law and stoichiometry (14.4)	

STS

state examples of science leading technology and technology leading science (14.1, 14.3)	
recognize the values and limitations of technological products and processes (14.1, 14.3)	
describe the interactions of science and technology (14.1, 14.2, 14.3)	

Skills

initiating and planning	design an experiment, including a labelled diagram, to test predictions for reactions occurring in electrochemical cells (14.1, 14.2, 14.3)	
	describe procedures for safe handling, storage, and disposal of materials used in the laboratory (14.1, 14.2)	
	develop a plan to build a cell (battery) using a trial-and-error procedure (14.1)	
performing and recording	construct and observe electrochemical cells (14.1, 14.2, 14.3)	
	investigate the issue of disposal of batteries (14.1);	
	compile and display information about electrochemical cells (all sections)	
analyzing and interpreting	identify the products of electrochemical cells (all sections)	
	compare predictions with observations (all sections)	
	identify the limitations of evidence collected (14.1, 14.2, 14.3)	
	explain discrepancies between predicted and measured cell potentials (14.2, 14.3)	
	assess the practicalities of different cell designs (14.1)	
	evaluate experimental designs for cells (14.1, 14.2, 14.3)	

communication and teamwork	work collaboratively in addressing problems and use appropriate SI notation and significant digits (all sections)	
	select and integrate information from various sources about technological applications of cells (all sections)	

Key Terms

4.1	electrode	
	electrolyte	
	electric potential difference	
	electric current	
	fuel cell	
14.2	porous boundary	
	half-cell	
	voltaic cell	
	cathode	
	anode	
	inert electrode	
	standard cell	
	standard cell potential	
	standard reduction potential	
	reference half-cell	
	corrosion	
	cathodic protection	
14.3	electrolytic cell	
	electrolysis	
14.4	Faraday's law	
	Faraday constant	

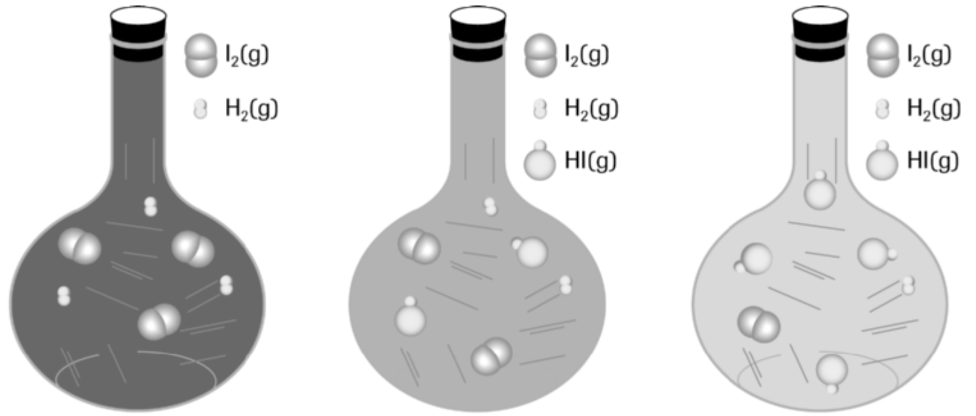
Key STS

appreciate the relationship of scientific knowledge and technological development in the applications of electrochemical cells (*Exploration: A Simple Electric Cell; Investigation 14.1: Designing an Electric Cell; WA: Hydrogen: Wonderfuel or Hype?; CS: The Ballard Fuel Cell*)	
give examples of the use of technology to solve practical problems related to corrosion (*MI: Home Corrosion Experiment; WA: Galvanizing Steel*)	
investigate the economic importance of electrochemical cells (*WA: Lewis Urry; CS: The Ballard Fuel Cell*)	
research a variety of careers (CC: Materials Engineering Technologist; WA: Lewis Urry; CC: Chemical Technologist)	

Key Terms

design, construct, observe, analyze, and evaluate an electric cell from a technological perspective and electrochemical cells from a scientific perspective (*Exploration: A Simple Electric Cell; Investigation 14.1: Designing an Electric Cell; Investigation 14.2: A Voltaic Cell (Demonstration); WA: Voltaic Cells Under Standard Conditions; Investigation 14.3: Testing Voltaic Cells; Investigation 14.4: A Potassium Iodide Electrolytic Cell; 14.5: Electrolysis (Demonstration*)	
create a presentation in a variety of formats of scientific and technological information about electrochemical cells *(Exploration: A Simple Electric Cell; Investigation 14.1: Designing an Electric Cell; CS: The Ballard Fuel Cell; Investigation 14.2: A Voltaic Cell (Demonstration); WA: Voltaic Cells Under Standard Conditions; Investigation 14.3: Testing Voltaic Cells; Investigation 14.4: A Potassium Iodide Electrolytic Cell; 14.5: Electrolysis (Demonstration)*	
investigate issues related to consumer, commercial, and industrial cells *(WA: Hydrogen: Wonderfuel or Hype? WA: Lewis Urry; CS: The Ballard Fuel Cell; WA: Electrolytic Cell Stoichiometry)*	

The Hydrogen–Iodine Reaction System



Initially, hydrogen (in excess) and iodine are added to the flask. The colour of the iodine vapour is the only easily observable property.

Early in the reaction, hydrogen and iodine form hydrogen iodide faster than hydrogen iodide forms hydrogen and iodine. Overall, the amount of iodine decreases, so the colour of the flask contents appears to lighten. Both hydrogen and hydrogen iodide are colourless.

At equilibrium, analysis shows that the flask contains all three substances. The purple colour shows that some iodine remains. The constancy of the colour is evidence that equilibrium exists. Forward and reverse reactions are occurring at equal rates.

Figure 4
When hydrogen and iodine are added to the flask, the colour of the iodine vapour is the only easily observable (empirical) property. An equilibrium equation describes this evidence theoretically. $H_2(g) + I_2(g) \rightarrow 2\ HI(g)$, $t = 448\ °C$

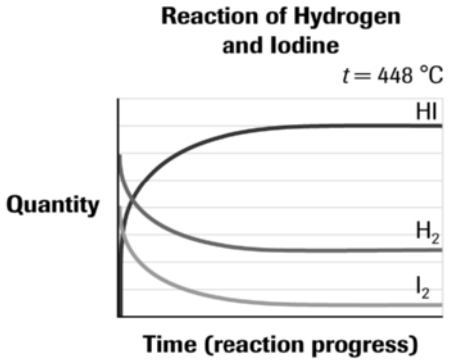


Figure 5
The graph of the quantity of each substance against time shows that the rate of reaction of the reactants decreases as the number of reactant molecules decreases, and the rate at which the product changes back to reactants increases as the number of product molecules increases. These two rates must become equal at some point, after which the quantity of each substance present will not change.

Equilibrium Law

SUMMARY Writing Equilibrium Law Expressions

Write an equilibrium law expression based on a balanced equation for the reaction system. Use single whole-number coefficients, written in net ionic form, and ignore concentrations of pure solid or liquid phases:

If: $aA + bB \rightleftarrows cC + dD$

then: $K_c = \dfrac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b}$

An equilibrium constant value

- always depends on the system temperature
- is independent of the reagent concentrations
- is independent of any catalyst present
- is independent of the time taken to reach equilibrium
- is normally stated as a numerical value, ignoring any units
- is greater, the more the system favours the formation of products

Student Worksheet LSM 15.1D

Chemical Equilibrium, Extra Exercises

---

1. For each of the following, write the chemical reaction equation with appropriate equilibrium arrows.
   (a) pH measurements indicate that acetic acid in vinegar is approximately 1% ionized into hydrogen ions and acetate ions.

   (b) Quantitative analysis of the reaction of sodium sulfate and calcium chloride solutions shows that the products are favoured.

   (c) Aluminium sulfate solution reacts quantitatively with a sodium hydroxide solution.

2. Chlorine and carbon monoxide gases are mixed in a 1.00 L container and the following equilibrium is established.

   $CO(g) + Cl_2(g) \leftrightarrows COCl_2(g)$

   Initially, 1.50 mol of chlorine was present with 2.55 mol of carbon monoxide. At equilibrium, 0.80 mol of $COCl_2(g)$ was found.

Concentration	**[ ] (mol/L)**	**[ ] (mol/L)**	**[ ] (mol/L)**
Initial			
Change			
Equilibrium			

   (a) Create an ICE table and calculate the percent reaction of chlorine.

   (b) Write the equilibrium law for this reaction.

   (c) At equilibrium, 1.75 mol of carbon monoxide and 0.70 mol of chlorine were present. Calculate the equilibrium constant.

3. Write the equilibrium law for each of the following chemical reaction equations.
   (a) $2\ SO_2(g) + O_2(g) \leftrightarrows 2\ SO_3(g)$

   (b) $2\ NO_2(g) \leftrightarrows 2\ NO(g) + O_2(g)$

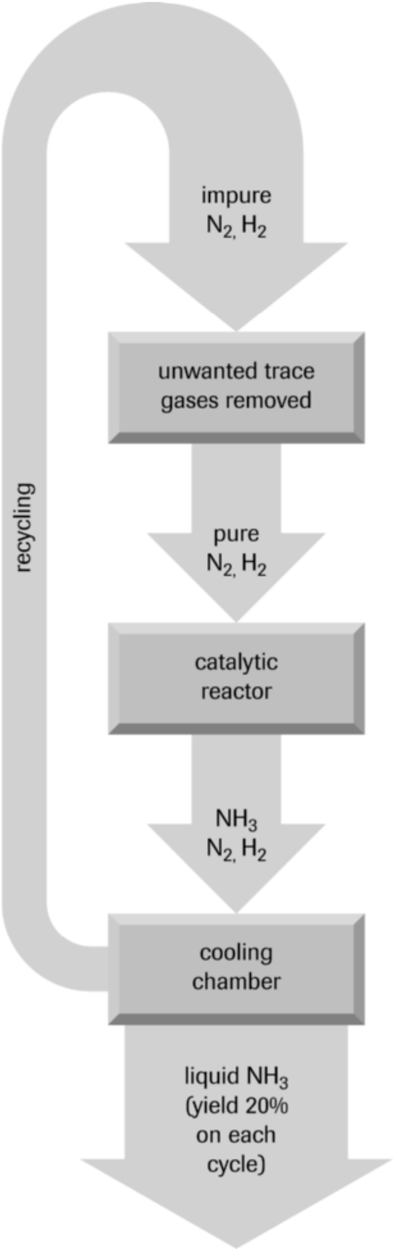
   (c) $N_2(g) + 3\ H_2(g) \leftrightarrows 2\ N\ H_3(g)$

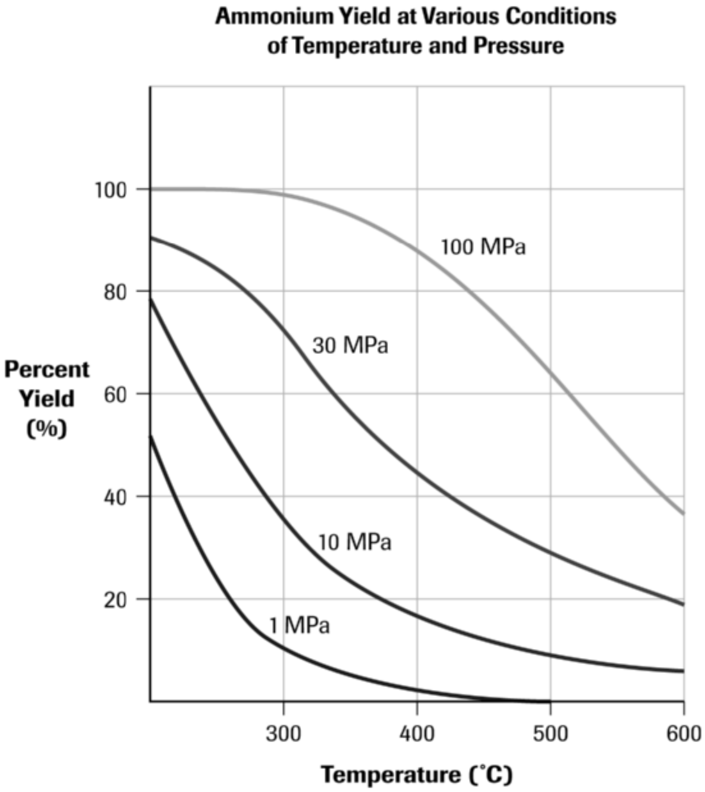
4. In an experiment at a high temperature, 0.500 mol/L of hydrogen bromide gas is placed into a sealed container and decomposes into hydrogen and bromine gases.
   (a) Write the equilibrium equation and law for this reaction.

   (b) The equilibrium concentration for $Br_2(g)$ in this system is found from calorimetry to be 0.130 mol/L. Use an ICE table to help calculate the equilibrium constant.

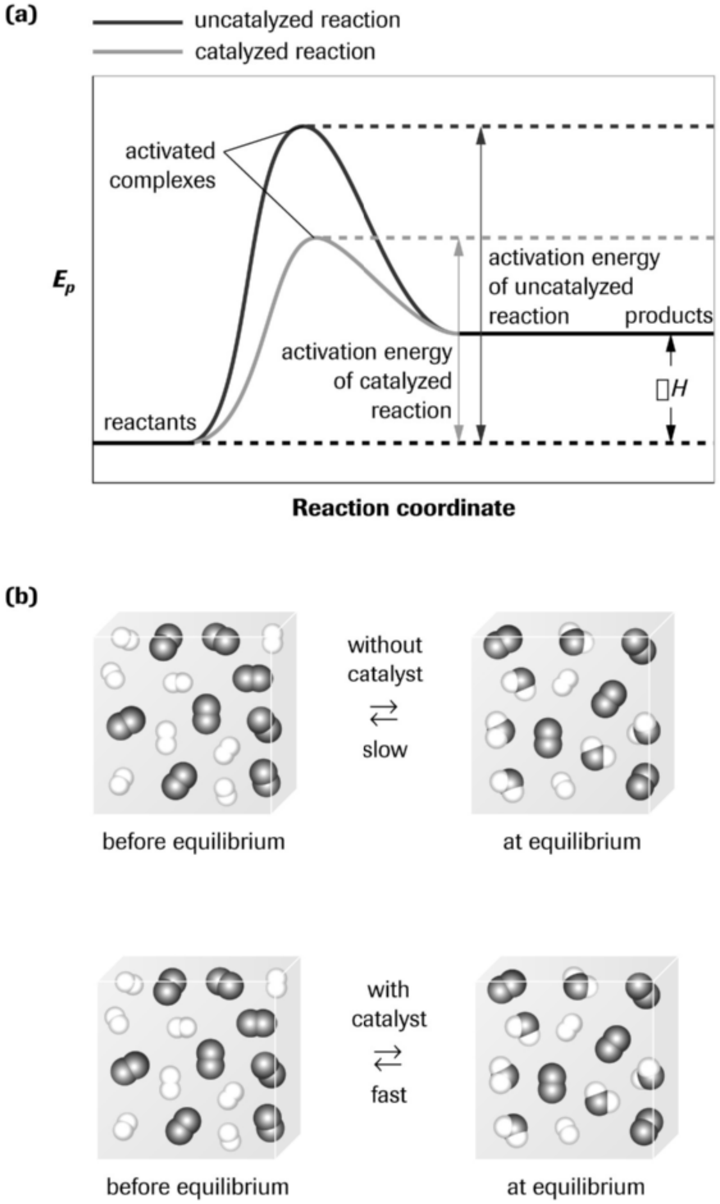
Concentration	**[ ] (mol/L)**	**[ ] (mol/L)**	**[ ] (mol/L)**
Initial			
Change			
Equilibrium			

The Production of Ammonia by the Haber Process





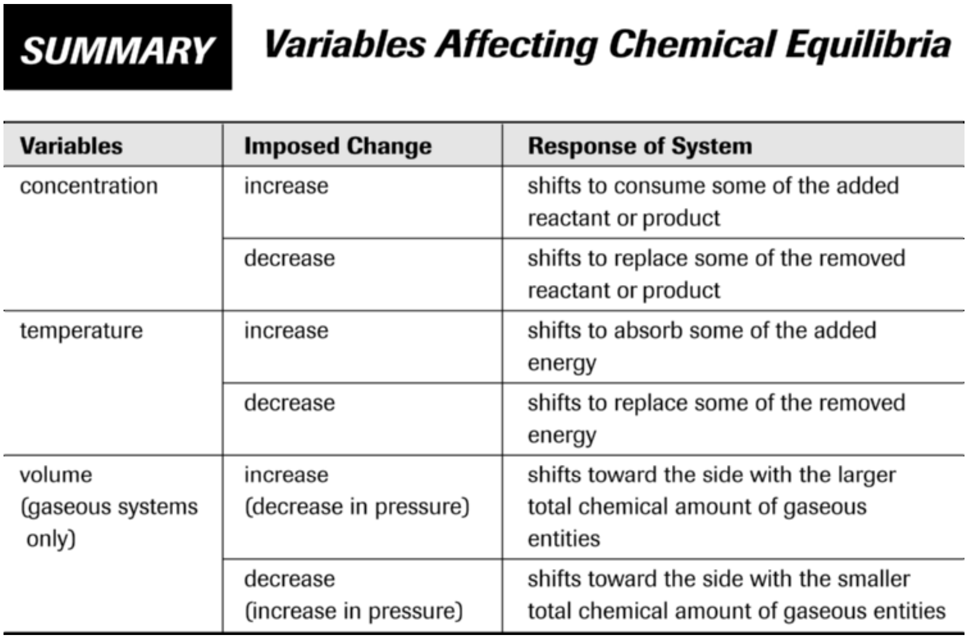
Catalysts and Equilibrium



(a) A catalyst reduces the activation energy by the same amount whether the reaction proceeds to the right or to the left.
(b) It does not affect the relative concentrations of entities.

Instructional Master LSM 15.2D

Variables Affecting Chemical Equilibria

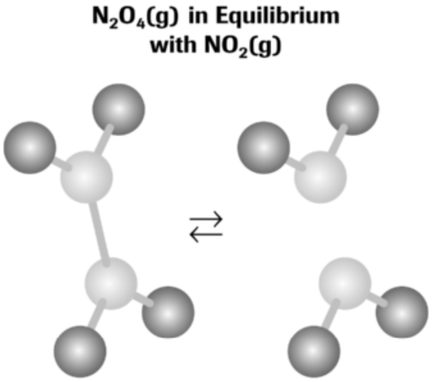


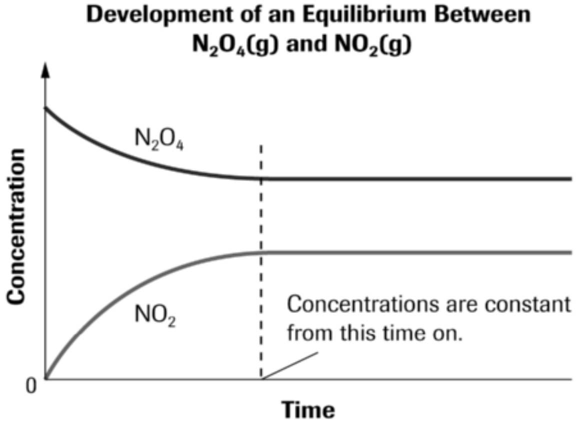
SUMMARY ***Variables Affecting Chemical Equilibria***

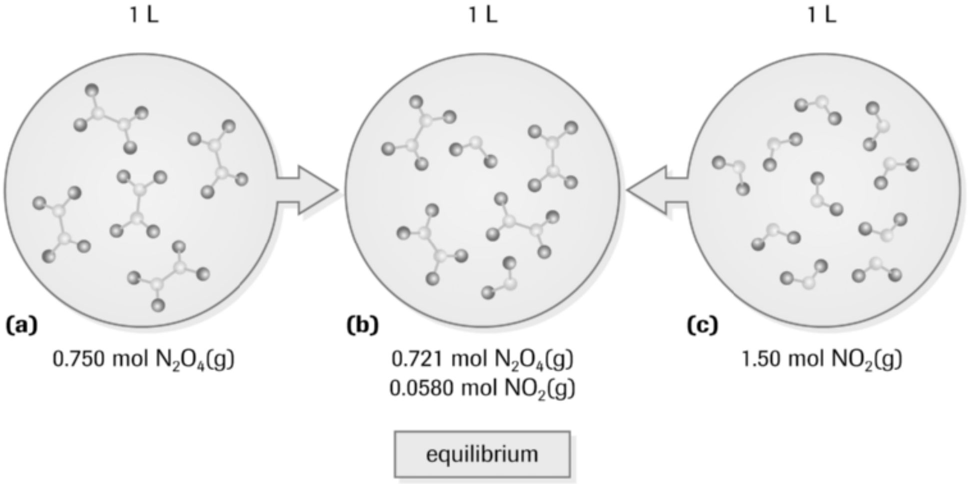
Variables	Imposed Change	Response of System
concentration	increase	shifts to consume some of the added reactant or product
	decrease	shifts to replace some of the removed reactant or product
temperature	increase	shifts to absorb some of the added energy
	decrease	shifts to replace some of the removed energy
volume (gaseous systems only)	increase (decrease in pressure)	shifts toward the side with the larger total chemical amount of gaseous entities
	decrease (increase in pressure)	shifts toward the side with the smaller total chemical amount of gaseous entities

Instructional Master LSM 15.2F

N_2O_4(g) in Equilibrium with NO_2(g)

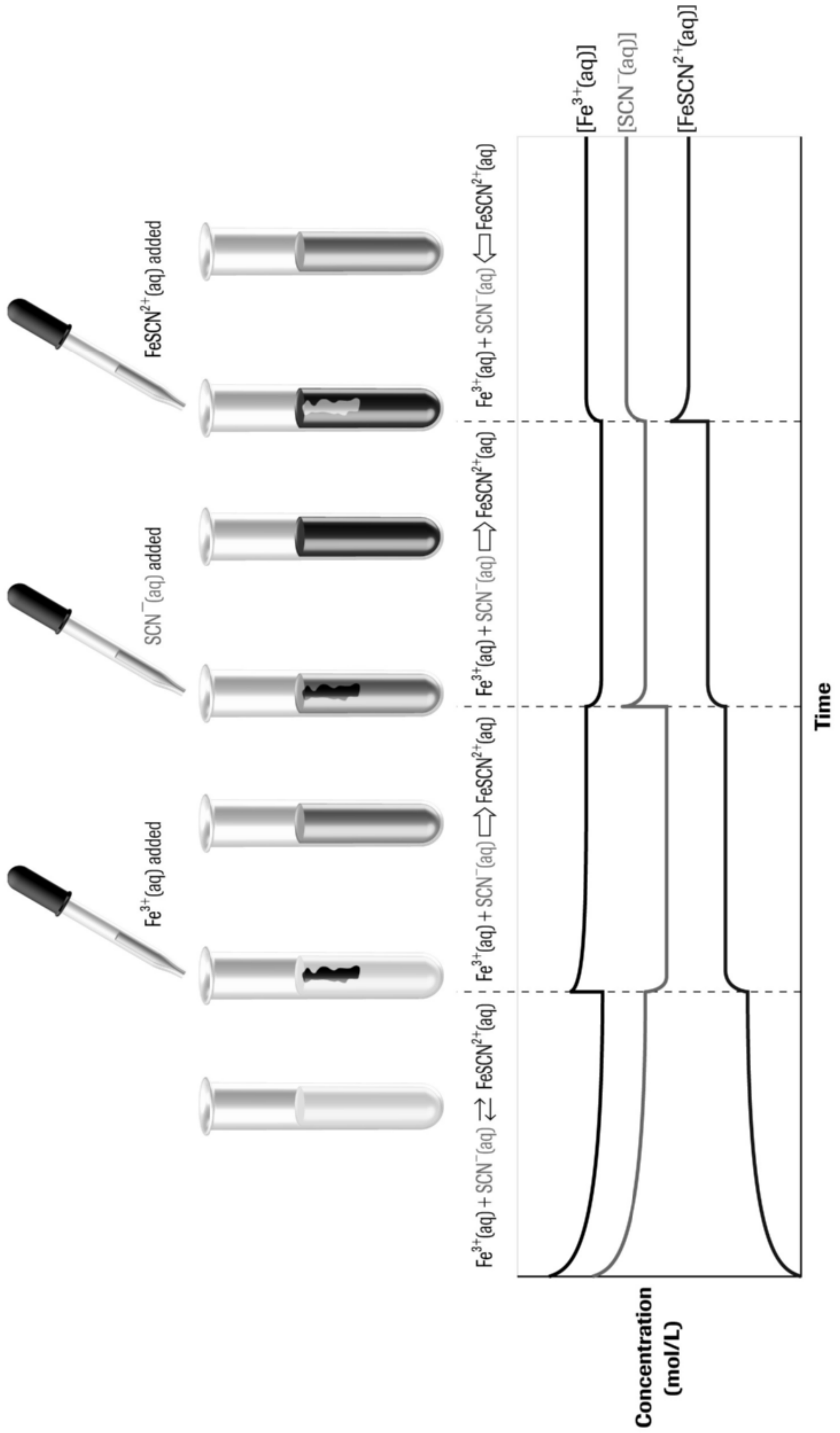






The same dynamic equilibrium composition is reached whether we start from pure N_2O_4(g), pure NO_2(g), or a mixture of the two, provided that environment, system and total mass remain the same.

Disturbing the Iron(III)–Thiocyanate Equilibrium



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Student Worksheet LSM 15.2J

Le Châtelier's Principle, Extra Exercises

---

1. Nitrogen monoxide, a major air pollutant, is formed in automobile engines from the endothermic reaction of nitrogen and oxygen gases.
   (a) Write the equilibrium reaction equation including the term "energy" in the equation.

   (b) Describe the direction of the equilibrium shift if the concentration of oxygen is increased.

   (c) Describe the direction of the equilibrium shift if the pressure is increased.

   (d) Gasoline burns better at higher temperatures. What are some disadvantages of the operation of automobile engines at higher temperatures?

2. In a sealed container, nitrogen monoxide and oxygen gases are in equilibrium with nitrogen dioxide gas. The reaction of nitrogen monoxide and oxygen is exothermic.

   $$2\ NO(g) + O_2(g) \leftrightharpoons 2\ NO_2(g) + \text{energy}$$

   Predict the equilibrium shift when the following changes are made.
   (a) the temperature is decreased

   (b) the concentration of NO(g) is decreased

   (c) the concentration of $NO_2(g)$ is increased

   (d) the volume of the system is decreased

3. The equilibrium of the iron(III)–thiocyanate system is convenient to study.

   $$\underset{\text{almost colourless}}{Fe^{3+}(aq)} \quad + \quad \underset{\text{colourless}}{SCN^{-}(aq)} \quad \leftrightharpoons \quad \underset{\text{red}}{FeSCN^{2+}(aq)}$$

   Predict the colour change in the equilibrium mixture when each of the following changes is made.
   (a) A crystal of KSCN(s) is added to the system.

   (b) A crystal of $CeCl_3(s)$ is added to the system.

   (c) A crystal of NaOH(s) is added to the system.

Solutions for Le Châtelier's Principle, Extra Exercises

---

1. Nitrogen monoxide, a major air pollutant, is formed in automobile engines from the endothermic reaction of nitrogen and oxygen gases.
   (a) Write the equilibrium reaction equation including the term "energy" in the equation.
   $N_2(g) + O_2(g) + \text{energy} \leftrightarrows 2\ NO(g)$
   (b) Describe the direction of the equilibrium shift if the concentration of oxygen is increased.
   The equilibrium shifts to the right.
   (c) Describe the direction of the equilibrium shift if the pressure is increased.
   There will be no shift in the equilibrium because the number of gas molecules is the same on both sides of the equilibrium equation.
   (d) Gasoline burns better at higher temperatures. What are some disadvantages of the operation of automobile engines at higher temperatures?
   At high temperatures, the equilibrium shifts to the right, producing more of the pollutant nitrogen monoxide. There may also be additional technical and safety problems as the temperature is increased.

2. In a sealed container, nitrogen monoxide and oxygen gases are in equilibrium with nitrogen dioxide gas. The reaction of nitrogen monoxide and oxygen is exothermic.
   $2\ NO(g) + O_2(g) \leftrightarrows 2\ NO_2(g) + \text{energy}$
   Predict the equilibrium shift when the following changes are made.
   (a) the temperature is decreased
   The equilibrium shifts to the right.
   (b) the concentration of $NO_{(g)}$ is decreased
   The equilibrium shifts to the left.
   (c) the concentration of $NO_{2(g)}$ is increased
   The equilibrium shifts to the left.
   (d) the volume of the system is decreased
   The equilibrium shifts to the right.

3. The equilibrium of the iron(III)–thiocyanate system is convenient to study.

$Fe^{3+}(aq)$	+	$SCN^{-}(aq)$	$\leftrightarrows$	$FeSCN^{2+}(aq)$
almost colourless		colourless		red

   Predict the colour change in the equilibrium mixture when each of the following changes is made.
   (a) A crystal of $KSCN(s)$ is added to the system.
   The red colour becomes more intense as the equilibrium shifts to the right.
   (b) A crystal of $CeCl_3(s)$ is added to the system.
   The red colour becomes more intense as the equilibrium shifts to the right.
   (c) A crystal of $NaOH(s)$ is added to the system.
   The red colour becomes less intense as the equilibrium shifts to the left. (Iron(III) ions precipitate as iron(III) hydroxide.)

Student Worksheet LSM 15.CS

Chapter Checklist

Have you mastered the concepts, applications, and skills associated with the following items? Check them off when you are confident in your understanding.

Knowledge

define equilibrium and state the criteria that apply to a chemical system in equilibrium (15.1)	
identify, write, and interpret chemical equations for systems at equilibrium (15.1, 15.2)	
predict, qualitatively, using Le Châtelier's principle, shifts in equilibrium caused by changes in temperature, pressure, volume, concentration, or the addition of a catalyst, and describe how these changes affect the equilibrium constant (15.2)	
define *Kc* and write equilibrium law expressions for given chemical equations, using lowest whole-number coefficients (15.1)	
calculate equilibrium constants and concentrations for homogeneous systems when concentrations at equilibrium are known, when initial concentrations and one equilibrium concentration are known, and when the equilibrium constant and one equilibrium concentration are known (15.1)	

STS

state that the goal of science is knowledge about the natural world (15.1, 15.2)	
list the characteristics of empirical and theoretical knowledge (15.2)	
state that a goal of technology is to solve practical problems (15.2)	

Skills

initiating and planning	state that a goal of technology is to solve practical problems (15.2)	
	design an experiment to show equilibrium shifts (15.2)	
	describe procedures for safe handling, storage, and disposal of materials used in the laboratory (15.1, 15.2)	
performing and recording	perform an experiment to test, qualitatively, predictions of equilibrium shifts (15.2)	
analyzing and interpreting	write the equilibrium law expression for a given equation (15.1)	
	analyze, qualitatively, the changes in concentrations of reactants and products after an equilibrium shift (15.2)	
	interpret data from a graph to determine when equilibrium is established, and determine the cause of a stress on the system (15.2)	
communication and teamwork	work collaboratively in addressing problems and communicate effectively (15.1, 15.2)	

Key Terms

15.1	closed system	
	equilibrium	
	phase equilibrium	
	solubility equilibrium	
	chemical reaction equilibrium	
	dynamic equilibrium theory	
	forward reaction	
	reverse reaction	
	ICE table	
	equilibrium constant, *Kc*	
	equilibrium law	
15.2	Le Châtelier's principle	
	equilibrium shift	

Key STS

recognize chemical equilibria in the natural world and in household products *(DYK: Diabetes: Blood Sugar Equilibrium; Section 15.2, Fig.5; Biology Connection: CO_2 Transport; WA: WQ—Poison Afloat; Section 15.2 Question 8; Chapter 15 Review questions 14, 24, 40, 43, 44)*	
appreciate how chemical equilibrium is used to control important industrial reactions *(Section 15.2—nitric acid, hydrochloric acid, sulfur trioxide, ammonia (the Haber process), methanol; CC: Chemical Process Engineer; CS: Urea Production in Alberta; Section 15.2 Questions 1, 2, 6, 7; Chapter 15 Review questions 38, 39, 41, 42)*	
appreciate how equilibrium theories developed *(Section 15.1; DYK: Lavoiser and Closed Systems; LE 15.A: The Synthesis of an Equilibrium Law; DYK: Related Interests; Section 15.1 Question 11; Section 15.2; Inv.15.2; Equilibrium Shifts (Demonstration); Inv.15.3: Testing Le Châtelier's Principle)*	

Key Skills

express the equilibrium law and calculate equilibrium constants from experimental data *(LE 15.A: The Synthesis of an Equilibrium Law; LE15.B: Determining an Equilibrium Constant; Section 15.1 Questions 5, 6, 7, 8, 9; Chapter 15 Review questions 26, 28)*	
create ICE tables for quantity concentration *changes* in reactions *(Sample Problem 15.2; Section 15.1 Questions 3, 6, 7, 8, 9, 10; Chapter 15 Review question 10; Chapter 15 Review questions 26, 27)*	
draw and interpret reaction progress graphs of concentration versus time for equilibrium reactions *(Section 15.1 Question 3; Section 15.2; Section 15.2 Questions 4; Chapter 15 Review question 26)*	

pH and Water Ionization

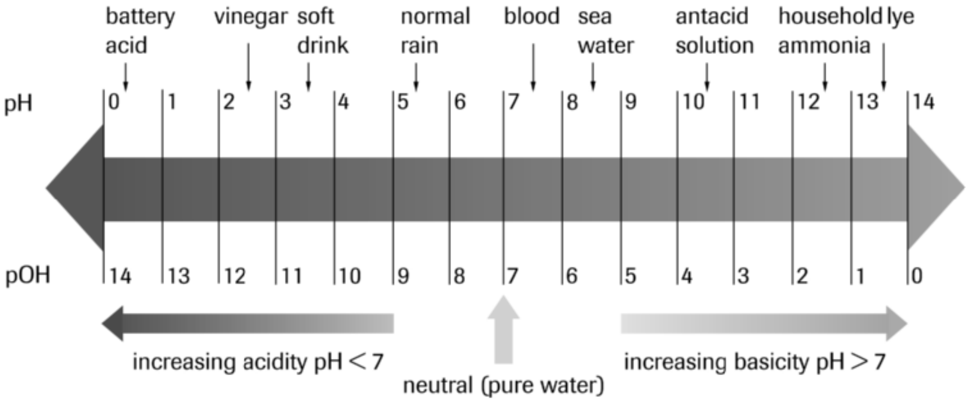

Figure 5
The pH scale

SUMMARY Water Ionization Conversions and Values (at SATP)

$$K_w = [H_3O^+(aq)][OH^-(aq)] = 1.0 \times 10^{-14}\ (mol/L)^2$$

$$pH = -\log[H_3O^+(aq)] \qquad pOH = -\log[OH^-(aq)]$$

$$[H_3O^+(aq)] = 10^{-pH} \qquad [OH^-(aq)] = 10^{-pOH}$$

$$pH + pOH = 14.00$$

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Student Worksheet LSM 16.1C

K_w, pH, and pOH Calculations, Extra Exercises

---

1. Calculate the $[\mathrm{OH}^-(aq)]$ in limes which have a $[\mathrm{H_3O}^+(aq)]$ of 1.3×10^{-2} mol/L.

2. Calculate the $[\mathrm{H_3O}^+(aq)]$ in lemons which have a $[\mathrm{OH}^-(aq)]$ of 2.0×10^{-12} mol/L.

3. A sodium hydroxide solution is prepared by dissolving 2.50 g to make 2.00 L of solution. Calculate the hydroxide and hydronium ion concentrations.

4. A 0.728 g sample of hydrogen chloride gas is dissolved in 200 mL of solution. Calculate the hydronium and hydroxide ion concentrations.

5. Vinegar has a hydronium ion concentration of 1.5×10^{-3} mol/L. Calculate the pH.

6. An ammonia solution has a pOH of 2.92. What is the concentration of hydroxide ions in the solution?

7. Calculate the pOH and pH of a solution made by dissolving 7.50 g of strontium hydroxide to make 500 mL of solution.

8. Complete the following table.

Table 1 pH of Common Substances

Substance	**$[\mathrm{H_3O}^+(aq)]$ (mol/L)**	**pH**	**$[\mathrm{OH}^-(aq)]$ (mol/L)**	**pOH**	**Acidic, basic, or neutral**
milk			3.2×10^{-4}		
pure water		7.0			
blood	4.0×10^{-8}				
cleaner				3.20	

Brønsted−Lowry Definitions and Predicting Reactions

SUMMARY Brønsted–Lowry Definitions

- An acid is a proton donor and a base is a proton acceptor, in a specific reaction.
- An acid–base reaction involves a single proton transfer from one entity (the acid) to another (the base).
- An amphiprotic entity (amphoteric substance) is one that acts as a Brønsted–Lowry acid in some reactions and as a Brønsted–Lowry base in other reactions.
- A conjugate acid–base pair consists of two entities with formulas that differ only by a proton.
- A strong acid has a very weak attraction for protons. A strong base has a very strong attraction for protons.
- The stronger an acid, the more weakly it holds its proton. The stronger a base, the more it attracts another proton.
- The stronger an acid, the weaker is its conjugate base. The stronger a base, the weaker is its conjugate acid.

SUMMARY A Five-Step Method for Predicting the Predominant Acid–Base Reaction

1. List all entities (ions, atoms, or molecules including $H_2O(l)$) initially present as they exist in aqueous solution. (Refer to Table 1, page 728.)
2. Identify and label all possible aqueous acids and bases, using the Brønsted–Lowry definitions.
3. Identify the strongest acid and the strongest base present, using the table of Relative Strengths of Aqueous Acids and Bases (Appendix I).
4. Write an equation showing a transfer of one proton from the strongest acid to the strongest base, and predict the conjugate base and the conjugate acid to be the products.
5. Predict the approximate position of the equilibrium, using the generalization developed in Lab Exercise 16.B on page 727, the Learning Tip on page 729, and the table of Relative Strengths of Aqueous Acids and Bases (Appendix I).

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Student Worksheet LSM 16.2E

Brønsted–Lowry Definitions, Extra Exercises

1. According to the Brønsted–Lowry concept, define an acid and a base in terms of an acid–base reaction.

2. (a) What is meant by an “acidic” solution?

   (b) Does a Brønsted–Lowry acid have to form an acidic solution? Explain.

3. There are many species that are classified as amphiprotic.
   (a) What does amphiprotic mean?

   (b) What general type of species is amphiprotic?

4. $HOOCCOO^{-}(aq) + PO_4^{3-}(aq) \leftrightarrows HPO_4^{2-}(aq) + OOCCOO^{2-}(aq)$

   (a) Label each species as A or B for both forward and reverse reactions.

   (b) Identify all conjugate acid–base pairs.

   (c) According to the Brønsted–Lowry concept, what determines the position of this equilibrium?

5. What is the generalization about the strength of an acid relative to its conjugate base?

Instructional Master LSM 16.3A

Summary: Calculating the pH of a Solution of a Weak Monoprotic Acid, HA(aq), Given the Value of K_a

---

Step 1 List the major entities in solution.

Step 2 Write balanced equations for all entities that may produce $H^+(aq)$.

Step 3 Identify the dominant equilibrium, and write the equilibrium constant equation for the dominant equilibrium.

Step 4 Use the coefficients in the balanced equation of the dominant equilibrium to determine the changes in initial concentrations that will occur as the dominant reaction proceeds to equilibrium. Record all values in an ICE table.

Step 5 Substitute the equilibrium concentrations of all entities (from the "Equilibrium" line in the ICE table) into the acid ionization constant equation.

Step 6 Assume that $[\mathrm{HA}]_{\text{initial}} - x \doteq [\mathrm{HA}]_{\text{initial}}$, but only if $\frac{[\mathrm{HA}]_{\text{initial}}}{K_a} \geq 100$.

Step 7 Solve for x (which, for a monoprotic acid, equals $[H^+(aq)]$).

Step 8 Use the 5% rule to validate any assumption made in Step 6.

Step 9 Calculate pH from $[H^+(aq)]$ (the value of x).

Instructional Master LSM 16.3B

Summary: Calculating the pH of a Solution of a Weak Base, HB(aq), Given the Value of K_b

---

Step 1 List the major entities in solution.

Step 2 Write balanced equations for all entities that may produce $OH^-(aq)$.

Step 3 Identify the dominant equilibrium, and write the equilibrium constant equation for the dominant equilibrium.

Step 4 Use the coefficients in the balanced equation of the dominant equilibrium to determine the changes in initial concentrations that will occur as the dominant reaction proceeds to equilibrium. Record all values in an ICE table.

Step 5 Substitute the equilibrium concentrations of all entities (from the “Equilibrium” line in the ICE table) into the base ionization constant equation.

Step 6 Assume that $[\mathrm{B}]_{\text{initial}} - x \doteq [\mathrm{B}]_{\text{initial}}$, but only if $\dfrac{[\mathrm{B}]_{\text{initial}}}{K_\mathrm{b}} \geq 100$, where $[\mathrm{B}]_{\text{initial}}$ represents the concentration of the base before the ionization process begins.

Step 7 Solve for x (which equals $[OH^-(aq)]$).

Step 8 Use the 5% rule to validate assumptions made in Step 6.

Step 9 Calculate pOH from $[OH^-(aq)]$ (the value of x).

Step 10 Calculate the pH by substituting the value of pOH into the equation pH + pOH = 14.

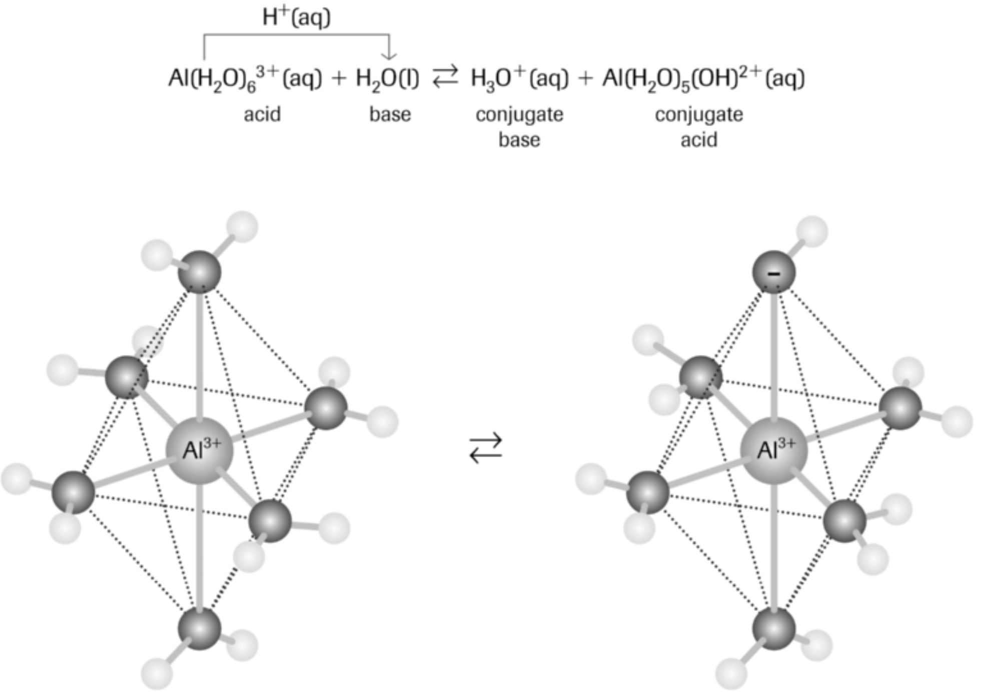
Instructional Master LSM 16.3C

The K_a for Some Metal Ions at SATP

Table 1 K_a for Some Metal Ions at SATP

Metal cation*	K_a
$Zr^{4+}(aq)$	2.1
$Sn^{2+}(aq)$	2.0×10^{-2}
$Fe^{3+}(aq)$	1.5×10^{-3}
$Cr^{3+}(aq)$	1.0×10^{-4}
$Al^{3+}(aq)$	9.8×10^{-6}
$Be^{2+}(aq)$	3.2×10^{-7}
$Fe^{2+}(aq)$	1.8×10^{-7}
$Pb^{2+}(aq)$	1.6×10^{-8}
$Cu^{2+}(aq)$	1.0×10^{-8}

* The aqueous metal ion is a hydrated complex ion (e.g., $Cu(H_2O)_4{}^{2-}(aq)$). Aqueous ions of transition metals are usually written in a simplified form, without showing the number of water molecules present in the actual hydrated complex ion, as shown in the table.



$Al(H_2O)_6{}^{3+}(aq) \rightleftarrows Al(H_2O_5)_5(OH)^{2+}(aq)$

Student Worksheet LSM 16.3D

Strengths of Acids, Extra Exercises

1. List three empirical properties that may be measured to distinguish among acids of different strengths.

2. Calculate the hydronium ion concentration and pH of a 0.10 mol/L solution of nitrous acid.

3. Calculate the hydrogen ion concentration and pH of a solution prepared by dissolving 10.70 g of ammonium chloride to make 2.00 L of solution.

4. Use the percent reaction value to determine the mass of sodium hydrogen sulfate required to prepare 500 mL of solution with a pH of 1.57.

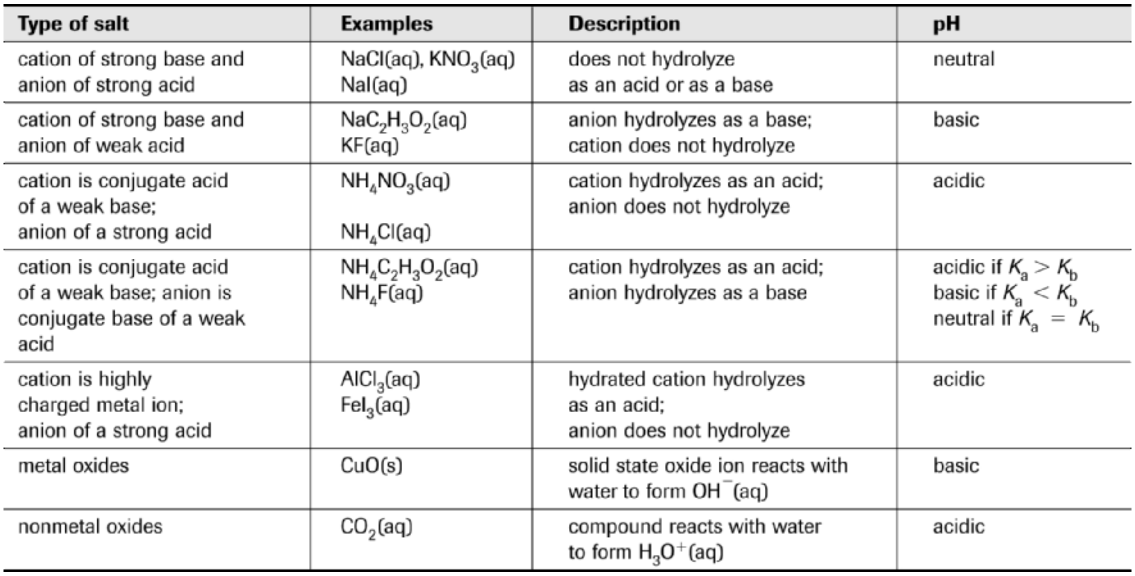
5. A 0.80 mol/L solution of an unknown acid, HX(aq), has a pH of 3.75.
   (a) Calculate the percent reaction.

   (b) Calculate the acid ionization constant.

6. Calculate the pH of a solution containing 0.25 mol/L of an acid with an acid ionization constant of 3.2×10^{-6} mol/L.

Instructional Master LSM 16.3F

Summary: Acid–Base Characteristics of Salts



Type of salt	Examples	Description	pH
cation of strong base and anion of strong acid	$NaCl(aq)$, $KNO_3(aq)$ $NaI(aq)$	does not hydrolyze as an acid or as a base	neutral
cation of strong base and anion of weak acid	$NaC_2H_3O_2(aq)$ $KF(aq)$	anion hydrolyzes as a base; cation does not hydrolyze	basic
cation is conjugate acid of a weak base; anion of a strong acid	$NH_4NO_3(aq)$ $NH_4Cl(aq)$	cation hydrolyzes as an acid; anion does not hydrolyze	acidic
cation is conjugate acid of a weak base; anion is conjugate base of a weak acid	$NH_4C_2H_3O_2(aq)$ $NH_4F(aq)$	cation hydrolyzes as an acid; anion hydrolyzes as a base	acidic if $K_a > K_b$ basic if $K_a < K_b$ neutral if $K_a = K_b$
cation is highly charged metal ion; anion of a strong acid	$AlCl_3(aq)$ $FeI_3(aq)$	hydrated cation hydrolyzes as an acid; anion does not hydrolyze	acidic
metal oxides	$CuO(s)$	solid state oxide ion reacts with water to form $OH^-(aq)$	basic
nonmetal oxides	$CO_2(aq)$	compound reacts with water to form $H_3O^+(aq)$	acidic

Titration Curves for Strong and Weak Acids

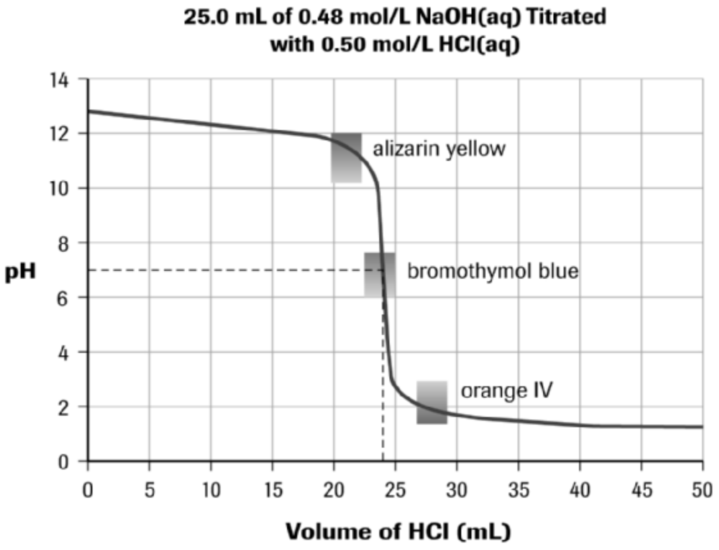


Figure 3
Alizarin yellow is not a suitable indicator because it will change colour long before the equivalence point of this strong acid–strong base reaction, which theoretically has a pH of exactly 7. Orange IV is also unsuitable; its colour change would occur too late. The pH at the middle of the colour change range for bromothymol blue is 6.8, which very closely matches the equivalence point pH; so, a titration analysis endpoint for this reaction, as indicated by bromothymol blue, should give accurate results.

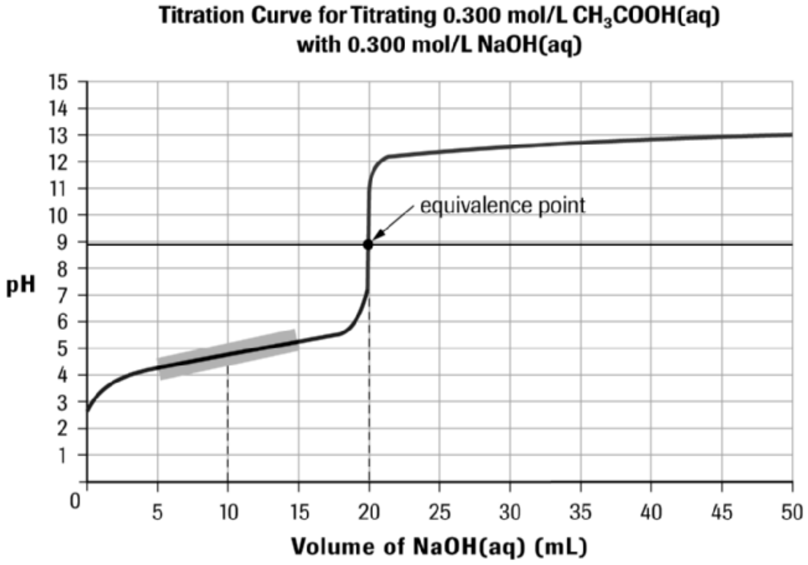


Figure 15
The highlighted plateau shows an effective buffering region during this titration of aqueous acetic acid with sodium hydroxide. At the point halfway to the equivalence point, the solution is a 1:1 ratio mix of an acetic acid–acetate ion buffer.

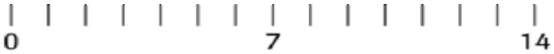
SUMMARY *pH Curve Reaction Information*

Empirical pH curves provide a wealth of information:
- initial pH of sample solutions
- pH when excess titrant is added
- number of quantitative reactions
- non-quantitative (equilibrium) reactions
- equivalence point(s) for indicator selection
- buffering regions

Student Worksheet LSM 16.4D

Indicators, Extra Exercises

1. Use the following pH scale to label the colours for bromothymol blue over the 0–14 pH range. Identify the form of the indicator for each distinct colour using conventional symbols.



2. **Problem**

   What is the approximate pH of an unknown solution?

   Evidence

   Separate samples of the unknown solution turned blue litmus to red, congo red to blue, and orange IV to yellow.

   Analysis

3. Three unknown solutions in unlabelled beakers have pH values of 5.8, 7.8, and 9.8. Write two diagnostic tests using indicators to identify the pH of each solution.

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Student Worksheet LSM 16.4F

pH Curves, Extra Exercises

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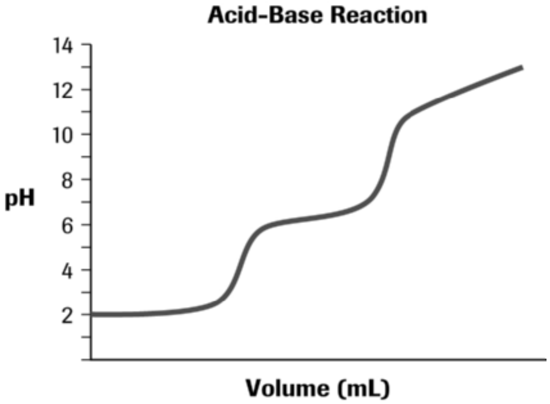
1. Use the accompanying sketch of a pH curve for a titration to answer the following questions.



   (a) Does the burette contain the acid or the base?

   (b) Is the sample reacted an acid or a base?

   (c) How many equivalence points are represented? Estimate the pH of each equivalence point.

   (d) How many quantitative reactions have occurred?

   (e) Choose the best indicator for each equivalence point?

2. Sketch a pH curve for the reaction of sulfuric acid with sodium hydroxide solution. All reaction steps are known to be quantitative. Include reaction equations.

3. A sodium hydrogen phosphate solution is to be titrated with hydrochloric acid. Only one quantitative reaction is observed. Sketch the pH curve and write equilibrium equations.

4. What are two advantages of detecting an endpoint using an indicator as opposed to identifying it with a pH meter?

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Student Worksheet LSM 16.4H

Acid–Base Stoichiometry, Extra Exercises

---

1. In a chemical analysis, 25.0 mL of sulfuric acid solution was titrated to the second endpoint with 0.358 mol/L KOH(aq). In the titration, an average volume of 18.2 mL was required. Calculate the molar concentration of the sulfuric acid.

2. Several 10.0 mL vinegar samples were titrated with a standardized 0.582 mol/L solution of sodium hydroxide. An average volume of 13.8 mL of sodium hydroxide was required to reach the phenolphthalein endpoint. What is the concentration of the vinegar solution?

3. A sodium borate solution was titrated to the second endpoint with 0.241 mol/L hydrobromic acid. An average volume of 15.2 mL of hydrobromic acid was required to react with 20.0 mL samples of sodium borate. Calculate the molar concentration of sodium borate.

4. **Problem**

What is the molar concentration of a hydrochloric acid solution?

Design

100.0 mL of a standard solution of sodium oxalate was prepared using 1.85 g of the dry solid. Using the second endpoint, 10.0 mL samples were titrated with hydrochloric acid.

Evidence

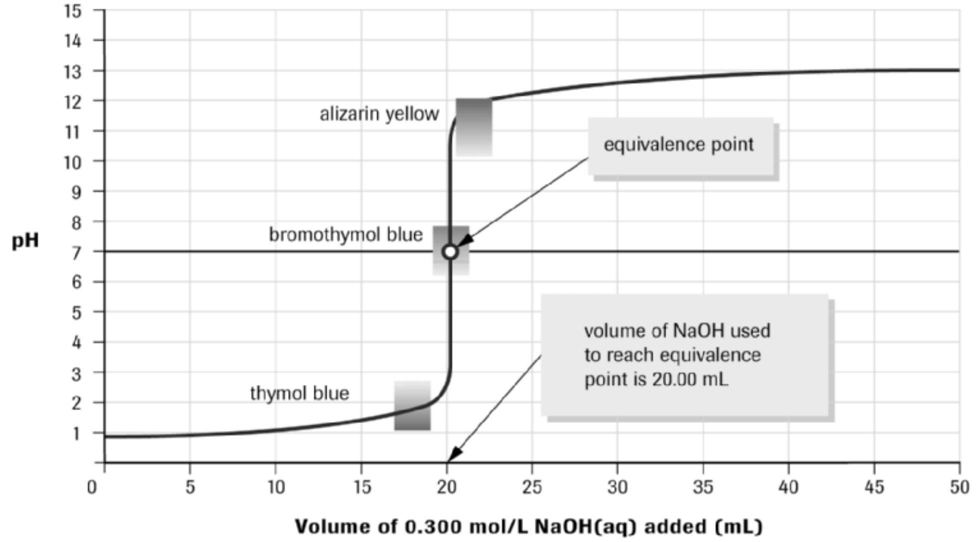
Table 1 Titration of 10.0 mL Samples of $Na_2OOCCOO(aq)$ with $HCl(aq)$

Trial	1	2	3	4
Final burette reading (mL)	16.1	31.5	46.9	16.9
Initial burette reading (mL)	0.3	16.1	31.5	1.5
Comment on endpoint	poor	good	good	good

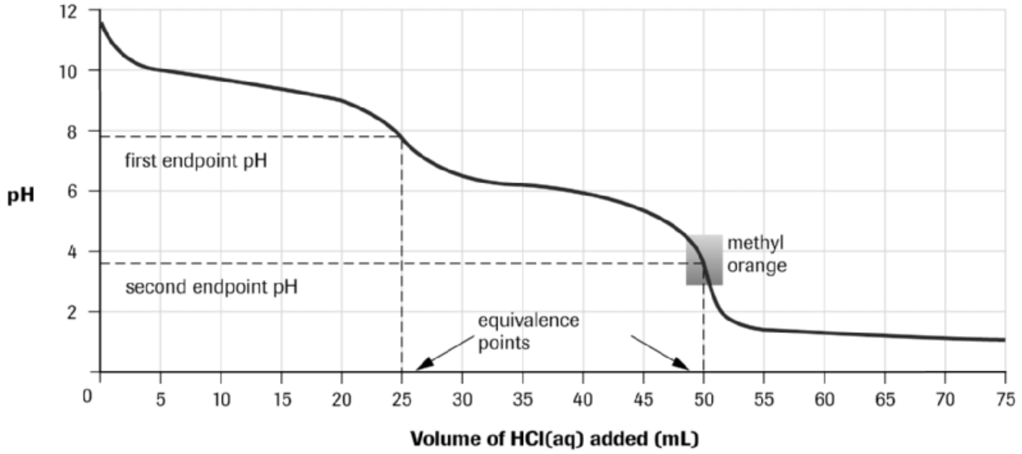
Analysis

Selecting an Indicator for a Titration

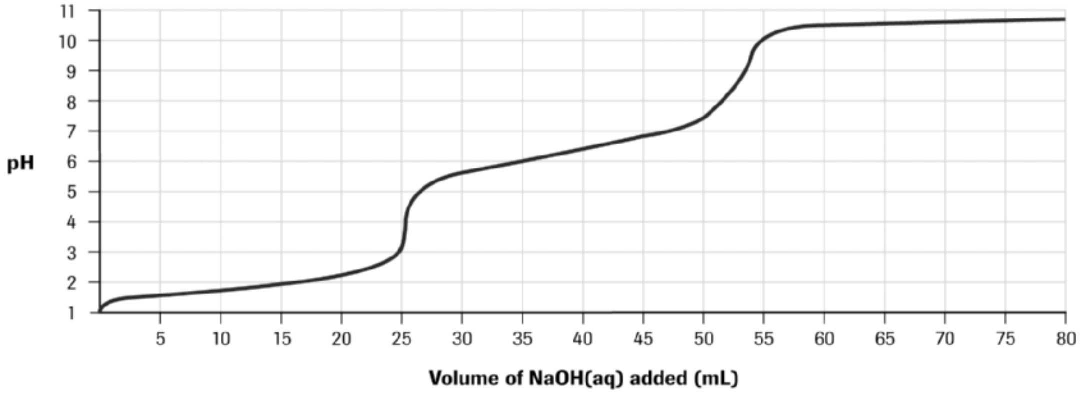
Titration Curve for Titrating 20.00 mL of 0.300 mol/L HCl(aq) with 0.300 mol/L NaOH(aq)



Titration Curve for Titrating 25.0 mL of 0.50 mol/L Na_2CO_3(aq) with 0.50 mol/L HCl(aq)

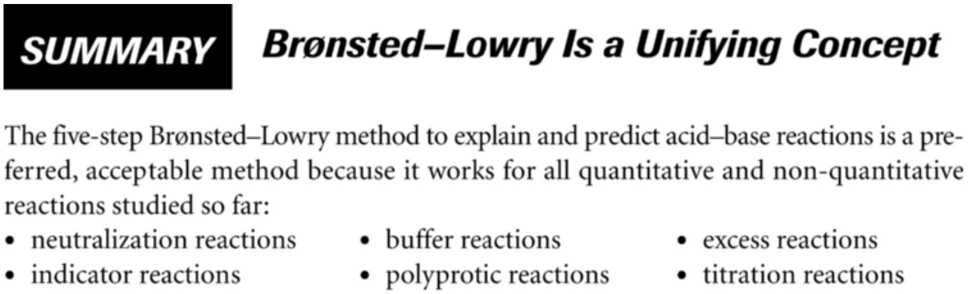


Titration Curve for Titrating 25.0 mL of 0.50 mol/L H_3PO_4(aq) with 0.50 mol/L NaOH(aq)



Instructional Master LSM 16.4L

Brønsted–Lowry Concept



SUMMARY **Brønsted–Lowry Is a Unifying Concept**

The five-step Brønsted–Lowry method to explain and predict acid–base reactions is a preferred, acceptable method because it works for all quantitative and non-quantitative reactions studied so far:

- neutralization reactions
- indicator reactions
- buffer reactions
- polyprotic reactions
- excess reactions
- titration reactions

Student Worksheet LSM 16.CS

Chapter Checklist

Have you mastered the concepts, applications, and skills associated with the following items? Check them off when you are confident in your understanding.

Knowledge

describe Brønsted–Lowry acids as proton donors and bases as proton acceptors (16.2)	
write Brønsted–Lowry equations and predict whether reactants or products are favoured for acid–base equilibrium reactions (including indicators and polyprotic acids and bases) (16.2, 16.4)	
identify polyprotic acids, polyprotic bases, conjugate pairs, and amphiprotic entities (16.2, 16.4)	
define a buffer as relatively large amounts of a conjugate acid–base pair in equilibrium that maintain a relatively constant pH when small amounts of acid or base are added (16.4)	
sketch and qualitatively interpret titration curves of monoprotic and polyprotic acids and bases, identifying equivalence points and regions of buffering for weak acid–strong base, strong acid–weak base, and strong acid–strong base reactions (16.4)	
define Kw , Ka , and Kb and use them to determine pH, pOH, $[H_3O^+]$, and $[OH^-]$ of acidic and basic solutions (16.1, 16.3)	
calculate equilibrium constants and concentrations for homogeneous systems and Brønsted–Lowry acids and bases (excluding buffers) when concentrations at equilibrium are known, when initial concentrations and one equilibrium concentration are known, and when the equilibrium constant and one equilibrium concentration are known (16.1, 16.3)	

STS

state that the goal of science is knowledge about the natural world (16.1, 16.2, 16.3, 16.4)	
state that a goal of technology is to solve practical problems (16.2, 16.3, 16.4)	

Skills

initiating and planning	design an experiment to show quantitative equilibrium shifts in concentration under a given set of conditions (16.4)	
	describe procedures for safe handling, storage, and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information (16.2, 16.4)	
performing and recording	prepare a buffer to investigate the relative abilities of a buffer and a control to resist a pH change when a small amount of strong acid or strong base is added (16.4)	
analyzing and interpreting	use experimental data to calculate equilibrium constants (16.3)	
communication and teamwork	work cooperatively in addressing problems and apply the skills and conventions of science in communicating information and ideas and in assessing results (16.2, 16.4)	

LSM 16.CS (cont'd)

Key Terms

16.1	ionization constant for water, K_w	
16.2	Brønsted–Lowry acid	
	Brønsted–Lowry base	
	Brønsted–Lowry reaction equation	
	amphiprotic	
	amphoteric	
	conjugate acid–base pair	
16.3	acid ionization constant, K_a	
	base ionization constant, K_b	
16.4	pH curve	
	buffering	
	buffer	
	buffer capacity	

Key STS

provide examples of buffers in living systems *(Section 16.4; Biology Connection: Homeostasis)*	
appreciate the development of equilibrium and acid–base theories and principles *(Section 16.2; Investigation 16.1: Creating an Acid–Base Strength Table; CS: Changing Ideas on Acids and Bases—The Evolution of a Scientific Theory; Section 16.2 Question 3; Sample Problem 16.3; Chapter 15 Review question 5; Unit 8 Review questions 2, 15, 16, 17, 18, 19, 20, 30, 31)*	

Key Skills

prepare a buffer to investigate the relative abilities of a buffer and a control to resist a pH change when a small amount of strong acid or strong base is added *(WA: Simulation—Preparation of Buffer Solutions; Investigation 16.4: Testing a Buffer Effect)*	