

*Student  
Workbook  
for  
Discovering  
Design with  
Chemistry*

Property of:

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## **Student Workbook for Discovering Design with Chemistry**

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**A copy of this document can be made for any student who owns Discovering Design with Chemistry**

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## Daily Assignments

If your school year consists of 180 days, you can use this list of daily assignments to complete the course in one school year.

### Week 1:

- Read the introduction and check out the course website
- Read pages 1-5, solving “Comprehension Check” problems 1&2.
- Read from page 6 to “Converting Between Units” on page 9, completing “Comprehension Check” problems 3&4 and completing Experiment 1.1.
- Read from “Converting Between Units” on page 9 to “Scientific Notation” on page 13, completing “Comprehension Check” problems 5&6.
- Read from “Scientific Notation” on page 13 to Experiment 1.2 on page 17, completing “Comprehension Check” problems 7&8.

### Week 2:

- Complete Experiment 1.2 and read to “Measuring Mass” on page 19, completing “Comprehension Check” problems 9&10.
- Read from “Measuring Mass” on page 19 to “More on Density” on page 25, completing “Comprehension Check” problems 11-14.
- Read from “More on Density” on page 25 to the end of page 26, completing Experiment 1.3 and “Comprehension Check” problem 15
- Do problems 1-13 in the review.
- Do problems 14-22 in the review.

### Week 3:

- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 1.
- Read from page 37 to the end of Experiment 2.1 on page 39, completing Experiment 2.1.
- Read from the end of Experiment 2.1 on page 39 to “Mass Conservation: It’s Not Just a Good Idea, It’s the Law!” on page 42, completing “Comprehension Check” problems 1&2.
- Read from “Mass Conservation: It’s Not Just a Good Idea, It’s the Law!” on page 42 to “Elements and Compounds” on page 44, completing Experiment 2.2 and “Comprehension Check” problem 3.

### Week 4:

- Read from “Elements and Compounds” on page 44 to the end of Experiment 2.3 on page 47, completing Experiment 2.3 and “Comprehension Check” problems 4&5.
- Read from the top of page 48 to the end of page 50, completing “Comprehension Check” problems 6&7.
- Read from the beginning of page 51 to “What’s Wrong with Dalton’s Theory: Part One” on page 54, completing “Comprehension Check” problems 8&9.
- Read from “What’s Wrong with Dalton’s Theory: Part One” on page 54 to the end of page 60, completing “Comprehension Check” problems 10&11.
- Do all problems in the review.

**Week 5:**

- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 2.
- Read from the top of page 67 to “Defining the Atoms that Make Up an Element” on page 70, completing “Comprehension Check” problems 1&2.
- Read from “Defining the Atoms that Make Up an Element” on page 70 to the end of page 74, completing “Comprehension Check” problems 3&4.
- Read from the top of page 75 to the “The Electromagnetic Spectrum” on page 78, completing “Comprehension Check” problems 5&6.

**Week 6:**

- Read from the “The Electromagnetic Spectrum” on page 78 to the end of page 80, completing Experiment 3.1 and “Comprehension Check” problem 7.
- Read from the top of page 81 to the end of Experiment 3.2 on page 84, completing Experiment 3.2 and “Comprehension Check” problem 8.
- Read from the end of Experiment 3.2 on page 84 to “More on The Bohr Model” on page 87, completing “Comprehension Check” problems 9&10.
- Read from “More on The Bohr Model” on page 87 to the end of page 90, completing “Comprehension Check” problem 11.
- Do problems 1-12 in the review.

**Week 7:**

- Do problems 13-23 in the review.
- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 3.
- Read from the top of page 97 to “Out with Orbits, In with Orbitals” on page 100, completing Experiment 4.1 and “Comprehension Check” problem 1.
- Read from “Out with Orbits, In with Orbitals” on page 100 to the end of page 104, completing “Comprehension Check” problems 2&3.

**Week 8:**

- Read from the beginning of page 105 to the end of page 109, completing “Comprehension Check” problems 4-6.
- Read from the beginning of page 110 to “Metals, Nonmetals, and the In-Betweens” on page 112, completing “Comprehension Check” problems 7&8.
- Read from “Metals, Nonmetals, and the In-Betweens” on page 112 to “Ionic Compounds” on page 114, completing Experiment 4.2 and “Comprehension Check” problem 9.
- Read from “Ionic Compounds” on page 114 to “An Important Characteristic of Ionic Compounds” on page 118, completing “Comprehension Check” problems 10-13.
- Read from “An Important Characteristic of Ionic Compounds” on page 118 to the end of page 120, completing Experiment 4.3.

**Week 9:**

- Do all problems in the review.
- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 4.

- Read from the top of page 127 to “More Complicated Lewis Structures” on page 131, completing “Comprehension Check” problems 1&2.
- Read from “More Complicated Lewis Structures” on page 131 to “Naming Covalent Compounds” on page 134, completing “Comprehension Check” problem 3.

**Week 10:**

- Read from “Naming Covalent Compounds” on page 134 to “A Consequence of Polar Covalent Bonds” on page 139, completing “Comprehension Check” problems 4-7.
- Read from “A Consequence of Polar Covalent Bonds” on page 139 to the definition of a polar covalent molecule on page 141, completing Experiment 5.1.
- Read from the definition of a polar covalent molecule on page 141 to Example 5.4 on page 146, completing “Comprehension Check” problems 8&9.
- Read from Example 5.4 on page 146 to the end of page 147, completing “Comprehension Check” problems 10-12.
- Read from the beginning of page 148 to “Why is Polarity Important?” on page 151, completing “Comprehension Check” problems 13-14.

**Week 11:**

- Read from “Why is Polarity Important?” on page 151 to the end of page 153, completing Experiment 5.2 and “Comprehension Check” problem 15.
- Do problems 1-6 in the review.
- Do all problems 7-14 in the review.
- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 5.

**Week 12:**

- Read from page 161 to Experiment 6.1 on page 165, completing “Comprehension Check” problems 1&2.
- Read from Experiment 6.1 on page 165 to “The Kinetic Theory of Matter” on page 167, completing Experiment 6.1 and “Comprehension Check” problem 3.
- Read from “The Kinetic Theory of Matter” on page 167 to the definition of the Kinetic Theory of Matter on page 169, completing Experiment 6.2.
- Read from the definition of the Kinetic Theory of Matter on page 169 to “Chemical Reactions” on page 172, completing “Comprehension Check” problems 4&5.
- Read from “Chemical Reactions” on page 172 to “Balancing Chemical Equations” on page 176, completing “Comprehension Check” problem 6.

**Week 13:**

- Read from “Balancing Chemical Equations” on page 176 to “The Mathematical Nature of Chemical Equations” on page 180, completing “Comprehension Check” problems 7-9.
- Read from “The Mathematical Nature of Chemical Equations” on page 180 to “Single and Double Displacement Reactions” on page 183, completing “Comprehension Check” problems 10-12.
- Read from “Single and Double Displacement Reactions” on page 183 to “Combustion Reactions” on page 185, completing Experiment 6.3 and “Comprehension Check” problem 13.
- Read from “Combustion Reactions” on page 185 to the end of page 188, completing Experiment 6.4 and “Comprehension Check” problem 14.
- Do problems 1-13 in the review.

**Week 14:**

- Do problems 14-16 in the review.
- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 6.
- Read from page 197 to the end of Example 7.1 on page 200, completing “Comprehension Check” problems 1&2.
- Read from the end of Example 7.1 on page 200 to Experiment 7.1 on page 202, completing “Comprehension Check” problems 3&4.

**Week 15:**

- Read from Experiment 7.1 on page 202 to the end of Experiment 7.1 on page 203, completing Experiment 7.1.
- Read from the end of Experiment 7.1 on page 203 to “There is a Limit!” on page 206, completing “Comprehension Check” problems 5&6.
- Read from “There is a Limit!” on page 206 to the definition of limiting reactant on page 208, completing Experiment 7.2.
- Read from the definition of limiting reactant on page 208 to “Stoichiometry Gets Massive” on page 212, completing “Comprehension Check” problems 7-9.
- Read from “Stoichiometry Gets Massive” on page 212 to “A Practical Application of Stoichiometry” on page 216, completing “Comprehension Check” problems 10-12.

**Week 16:**

- Read from “A Practical Application of Stoichiometry” on page 216 to the end of page 217, completing Experiment 7.3.
- Do problems 1-13 in the review.
- Do problems 14-20 in the review.
- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 7.

**Week 17:**

- Read from page 227 to the end of Experiment 8.1 on page 228, completing Experiment 8.1.
- Read from the top of page 229 to “Empirical and Molecular Formulas” on page 232, completing “Comprehension Check” problems 1-3.
- Read from “Empirical and Molecular Formulas” on page 232 to “Determining Empirical Formulas of Metal Oxides” on page 235, completing “Comprehension Check” problems 4-6.
- Read from “Determining Empirical Formulas of Metal Oxides” on page 235 to “More complicated Combustion Analysis” on page 239, completing “Comprehension Check” problems 7&8.
- Read from “More complicated Combustion Analysis” on page 239 to “Polyatomic Ions” on page 243, completing “Comprehension Check” problems 9-11.

**Week 18:**

- Read from “Polyatomic Ions” on page 243 to the end of page 247, completing “Comprehension Check” problems 12-14.
- Do problems 1-12 in the review.
- Do problems 13-20 in the review.



- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 8.

**Week 19:**

- Read from page 257 to Experiment 9.1 on page 260, completing “Comprehension Check” problem 1.
- Read from Experiment 9.1 on page 260 to the paragraph that starts “As a side note” on page 262, completing Experiment 9.1.
- Read from the paragraph that starts “As a side note” on page 262 to the end of page 264, completing “Comprehension Check” problems 2&3 and Experiment 9.2.
- Read from page 265 to the end of Experiment 9.3 on page 267, completing “Comprehension Check” problem 4 and Experiment 9.3.
- Read from the end of Experiment 9.3 on page 267 to “Using Molarity in Stoichiometry” on page 270, completing “Comprehension Check” problems 5&6.

**Week 20:**

- Read from “Using Molarity in Stoichiometry” on page 270 to “This is Depressing!” on page 274, completing “Comprehension Check” problems 7-10.
- Read from “This is Depressing!” on page 274 to Example 9.7 on page 276, completing Experiment 9.4.
- Read from Example 9.7 on page 276 to the end of page 279, completing “Comprehension Check” problems 11-13.
- Do problems 1-13 in the review.
- Do problems 14-23 in the review.

**Week 21:**

- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 9.
- Read from page 289 to the end of page 291, completing “Comprehension Check” problem 1.
- Read from page 292 to “Charles’s Law” on page 294, completing Experiment 10.1.
- Read from “Charles’s Law” on page 294 to Example 10.1 on page 296, completing Experiment 10.2.

**Week 22:**

- Read from Example 10.1 on page 296 to “This Law is Ideal!” on page 301, completing “Comprehension Check” problems 2-5.
- Read from “This Law is Ideal!” on page 301 to “Dalton’s Law of Partial Pressures” on page 306, completing “Comprehension Check” problems 6-9.
- Read from “Dalton’s Law of Partial Pressures” on page 306 to Experiment 10.3 on page 312, completing “Comprehension Check” problem 10.
- Read from Experiment 10.3 on page 312 to Experiment 10.4 on page 314, completing “Comprehension Check” problem 11 and Experiment 10.3.
- Read from Experiment 10.4 on page 314 to the end of page 316, completing Experiment 10.4. Also, do problems 1-9 in the review.

**Week 23:**

- Do problems 10-23 in the review.
- Correct any of your errors in the review and study for the test.

- Take the test for Chapter 10.
- Read from page 325 to Experiment 11.1 on page 327.
- Read from Experiment 11.1 on page 327 to “The Chemical Definition of Acids and Bases” on page 329, completing Experiment 11.1 and “Comprehension Check” problems 1&2.

**Week 24:**

- Read from “The Chemical Definition of Acids and Bases” on page 329 to the end of page 334, completing “Comprehension Check” problems 3-5.
- Read from page 335 to Experiment 11.2 on page 339, completing “Comprehension Check” problems 6&7.
- Read from Experiment 11.2 on page 339 to “Acid/Base Neutralization” on page 341, completing Experiment 11.2 and “Comprehension Check” problem 8.
- Read from “Acid/Base Neutralization” on page 341 to Experiment 11.3 on page 345, completing “Comprehension Check” problems 9&10.
- Read from Experiment 11.3 on page 345 to “Diluting Acid and Bases” on page 347, completing Experiment 11.3

**Week 25:**

- Read from “Diluting Acid and Bases” on page 347 to the end of page 348, completing “Comprehension Check” problem 11. Also, do problems 1-10 in the review.
- Do problems 11-21 in the review.
- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 11.
- Read from page 357 to Example 12.2 on page 361, completing “Comprehension Check” problem 1.

**Week 26:**

- Read from Example 12.2 on page 361 to Experiment 12.1 on page 366, completing “Comprehension Check” problems 2&3.
- Read from Experiment 12.1 on page 366 to the end of page 367, completing Experiment 12.1.
- Read from page 368 to the end of Experiment 12.2 on page 369, completing Experiment 12.2.
- Read from the end of Experiment 12.2 on page 369 to “Counting Electrons and Balancing Simple Redox Equations” on page 374, completing “Comprehension Check” problems 4-6.
- Read from “Counting Electrons and Balancing Simple Redox Equations” on page 374 to “Electroplating” on page 378, completing “Comprehension Check” problems 7-9.

**Week 27:**

- Read from “Electroplating” on page 378 to the end of page 380, completing Experiment 12.3 and “Comprehension Check” problem 10.
- Do all problems in the review.
- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 12.
- Read from page 387 to “Specific Heat Capacity” on page 391, completing “Comprehension Check” problems 1-3.

**Week 28:**

- Read from “Specific Heat Capacity” on page 391 to “Measuring Heat” on page 393, completing Experiment 13.1 and “Comprehension Check” problem 4.
- Read from “Measuring Heat” on page 393 to Experiment 13.2 on page 397, completing “Comprehension Check” problems 5&6.
- Read from Experiment 13.2 on page 397 to the end of the experiment on page 399, completing Experiment 13.2.
- Read from the end of Experiment 13.2 on page 399 to Experiment 13.3 on page 403, completing “Comprehension Check” problems 7-9.
- Read from Experiment 13.3 on page 403 to “The Heat Associated with Chemical Reactions” on page 406, completing Experiment 13.3 and “Comprehension Check” problem 10.

**Week 29:**

- Read from “The Heat Associated with Chemical Reactions” on page 406 to the end of page 408, completing Experiment 13.4 and “Comprehension Check” problem 11.
- Do problems 1-13 in the review.
- Do problems 14-23 in the review.
- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 13.

**Week 30:**

- Read from page 417 to the end of Experiment 14.1 on page 419, completing Experiment 14.1.
- Read from the end of Experiment 14.1 on page 419 to “Enthalpy Change and Hess’s Law” on page 422, completing “Comprehension Check” problems 1&2.
- Read from “Enthalpy Change and Hess’s Law” on page 422 to “Activation Energy” on page 427, completing “Comprehension Check” problems 3-5.
- Read from “Activation Energy” on page 427 to “Thermodynamics” on page 430, completing “Comprehension Check” problem 6.
- Read from “Thermodynamics” on page 430 to “Changes in Entropy” on page 434, completing “Comprehension Check” problems 7&8.

**Week 31:**

- Read from “Changes in Entropy” on page 434 to “The Gibbs Free Energy” on page 439, completing “Comprehension Check” problems 9-11.
- Read from “The Gibbs Free Energy” on page 439 through “Comprehension Check” problem 12 on page 442, completing “Comprehension Check” problem 12.
- Read from the end of “Comprehension Check” problem 12 in page 442 to the end of page 445, completing Experiment 14.2 and “Comprehension Check” problem 13.
- Do problems 1-10 in the review.
- Do problems 11-20 in the review.

**Week 32:**

- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 14.
- Read from page 453 to “What Things Affect the Rate of a Chemical Reaction?” on page 455, completing “Comprehension Check” problem 1.

- Read from “What Things Affect the Rate of a Chemical Reaction?” on page 455 to “Collision Theory” on page 457, completing Experiment 15.1 and “Comprehension Check” problem 2.
- Read from “Collision Theory” on page 457 to “Determining Reaction Orders” on page 462, completing “Comprehension Check” problem 3.

**Week 33:**

- Read from “Determining Reaction Orders” on page 462 to “Activation Energy, Temperature, and the Rate Constant” on page 467, completing “Comprehension Check” problems 4&5.
- Read from “Activation Energy, Temperature, and the Rate Constant” on page 467 to step 12 of Experiment 15.2 on page 470, completing “Comprehension Check” problems 6&7.
- Read from step 12 of Experiment 15.2 on page 470 to “How Catalysts Work” on page 471, completing Experiment 15.2 and “Comprehension Check” problem 8.
- Read “How Catalysts Work” on page 471 to the end of page 475, completing “Comprehension Check” problems 9-11.
- Do problems 1-13 in the review.

**Week 34:**

- Do problems 14-22 in the review.
- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 15.
- Read from page 483 to “The Equilibrium Constant” on page 486, completing Experiment 16.1 and “Comprehension Check” problem 1.
- Read from “The Equilibrium Constant” on page 486 to “Ignoring Things” on page 490, completing “Comprehension Check” problems 2&3.

**Week 35:**

- Read from “Ignoring Things” on page 490 to “Why Salt Melts Ice” on page 494, completing “Comprehension Check” problems 4-6.
- Read from “Why Salt Melts Ice” on page 494 to “More on Le Chatelier’s Principle” in page 497, completing “Comprehension Check” problem 7.
- Read from “More on Le Chatelier’s Principle” in page 497 to “Le Chatelier’s Principle and Temperature” on page 500, completing Experiment 16.2 and “Comprehension Check” problem 8.
- Read from “Le Chatelier’s Principle and Temperature” on page 500 to “Le Chatelier’s Principle and Pressure” on page 503, completing Experiment 16.3 and “Comprehension Check” problem 9.
- Read from “Le Chatelier’s Principle and Pressure” on page 503 to the end of page 507, completing “Comprehension Check” problems 10&11.

**Week 36:**

- Do problems 1-8 in the review.
- Do problems 9-18 in the review.
- Correct any of your errors in the review and study for the test.
- Take the test for Chapter 16.
- Celebrate the fact that you are done with chemistry!

## Worksheets

### Chapter 1 Comprehension Check Questions

1. Which is longer: an object that is 1 inch long or an object that is 1 centimeter long?
2. How many significant figures are in each of the numbers below?
  - a. 0.01020 in
  - b. 12,007 cm
  - c. 609,000 cm
  - d. 89.10400 in
3. What is the correct answer when 3.1 in are added to 8.991 inches?
4. What is the correct answer when a measurement of 45.5 cm is multiplied by 9 cm?

5. A laptop computer is 33.56 cm wide. How many inches wide is the laptop?

6. The diameter of a U.S. dime is 17.9 millimeters. How many meters is that?

7. Convert the following numbers to scientific notation so they each have five significant figures:

a. 53,000,000,000

b. 0.0094230

8. Convert the following numbers into decimal notation:

a.  $8.612 \times 10^{-8}$

b.  $6.965 \times 10^4$

9. Convert 7.4 liters into  $\text{cm}^3$ .

10. A box is measured to have a volume of 11.7 liters. What is its volume in  $\text{m}^3$ ?

11. Convert 7.4 kg into cg.

12. Convert 11,230 mL into kL.



13. A chunk of copper has a volume of 20.0 mL. What is its mass? (The density of copper is 8.96 g/mL)
14. The density of silver is 10.49 g/mL. What is the volume of a 9.43-kg sample?
15. The mass of a full can of Coke is 394 g. Because it uses an artificial sweetener, the mass of a full can of Diet Coke is lower: only 355 g. The volume of both cans (including the outside of the can) is 359 mL. Will either can float in water when it is full? (The density of water is 1.0 g/mL.)



4. You are measuring the volume of an object using a scale that is marked off with lines that represent 10 mL each. To what level of precision (one mL, tenths of an mL, hundredths of an mL, etc.) should you report your measurement?
5. An object has a mass of 123.4 kg, which is the same as 8.456 slugs. Which measures more mass: 1 slug or 1 kg?

6. A box has a volume of  $1.01 \text{ m}^3$ . Two students measure the box. The first says the volume is  $1 \text{ m}^3$ , while the second says the volume is  $1.23 \text{ m}^3$ . Assuming they are reporting the correct number of significant figures for the measuring devices used, which student used the more precise device? Which student provided the more accurate answer?

7. How many significant figures are in the following measurements?

a.  $1.06 \times 10^4 \text{ mL}$

b.  $12,000 \text{ cm}$

c.  $0.0340 \text{ kg}$

d.  $1.0 \times 10^1 \text{ in}$

8. When you put ice cubes in a glass of water, the ice cubes float. What does that tell you about the density of ice compared to the density of water?



11. A carpet-layer measures a room to be 4.6 m long and 3.2 m wide. He multiplies the two measurements and reports the area of the room to be 14.72. What two things are wrong with his answer?

12. What is the correct answer to the following equation?

$$21.0234 \text{ g} - 12 \text{ g}$$

13. A U.S. dime has a diameter of 17.9 mm. What is the diameter in m?

14. A soup can has a mass of 490 grams. What is its mass in kg?

15. The volume of a soccer ball is 6,080 mL. How many kL is that?

16. An object has a length of 15.0 inches. What is its length in meters? (2.54 cm = 1 inch)

17. Convert the following measurements into scientific notation with 4 significant figures each.

a. 0.0001214 kg

b. 34,500 m

c. 123,500,000 mg

d. 0.01010 km

18. Convert the following measurements into decimal notation.

a.  $1.2 \times 10^{-2}$  km

b.  $7.82 \times 10^7$  mL

c.  $9.1 \times 10^{-8}$  kg

d.  $8.912 \times 10^6$  m



19. A bowl has a volume of 1.1 L. What is its volume in cubic centimeters?

20. A sample of liquid has a volume of 143.6 L. What is the volume in  $\text{m}^3$ ?

21. Aluminum has a density of 2.70 g/mL. What is the volume of an aluminum block that has a mass of 55.67 kg?

22. A sample of what looks like silver has a mass of 1.7 kg and a volume of 0.164 liters. Is it really silver? (The density of silver is 10.49 g/mL.)

**Chapter 2 Comprehension Check Questions**

1. Classify the following as a mixture or a pure substance:
  - a. A bowl of cereal and milk
  - b. An ice cube
  - c. A drink made by dissolving a powder in water
  - d. A sample of oxygen
  
2. Classify the following mixtures as heterogeneous or homogeneous:
  - a. A bowl of cereal and milk
  - b. A clump of dirt pulled from a lawn
  - c. A drink made by dissolving a powder in water
  - d. A clear sample of sea water



- b. The sodium, hydrogen, and oxygen cannot be made into simpler substances, but the carbon dioxide can be made into two simpler substances. Identify each of the chemicals involved (baking soda, sodium, hydrogen, oxygen, and carbon dioxide) as compounds or elements.
5. Laughing gas is often used as an anesthetic by dentists. It is a pure substance made up of two elements: nitrogen and oxygen. The air you are breathing is mostly a mixture of nitrogen and oxygen. Why can't air be used as an anesthetic? After all, it contains the same elements as laughing gas.
6. A chemist reacts 10.00 grams of copper and 3.21 grams of sulfur to make 9.56 grams of a blue powder. He also has some leftover copper. How much of each element should he react to make 100.4 grams of the blue powder with no leftovers?
7. A chemist reacts 50.0 grams of sulfur with 50.0 grams of oxygen to make 83.4 grams of a gas, along with some leftover sulfur. In a different experiment, he reacts 20.0 grams of sulfur with 30.0 grams of oxygen to make 49.9 grams of a gas, along with some leftover oxygen. Are the two gases he made the same?

8. Which of the following are not consistent with Dalton's Atomic Theory?
- a. Atoms cannot be created or destroyed in a chemical reaction.
  - b. Compounds are composed of identical atoms.
  - c. Chemical reactions change the atoms that are involved.
  - d. When two compounds react, the atoms change how they are associated with one another.

9. If a chemist reacts 6.4 g of copper with 1.6 grams of oxygen, cupric oxide is made. It is composed of one copper atom and one oxygen atom. However, copper and oxygen can also combine to make cuprous oxide, which is made of two copper atoms and one oxygen atom. Suppose you react 1.6 grams of oxygen with copper to make cuprous oxide. How many grams of copper would you have to use? (HINT: Think about the number of copper atoms in each molecule.)
10. A gold atom has 79 protons in its nucleus. An atom of carbon has only 12 protons in its nucleus. Suppose Rutherford did his experiment with a carbon foil instead of a gold foil. Would he see more deflected particles, fewer deflected particles, or the same number of deflected particles?
11. Consider the following four atoms. Which pair could be called isotopes?

- a. An atom made from six protons, six electrons, and six neutrons
- b. An atom made from eight protons, eight electrons, and nine neutrons
- c. An atom made from ten protons, ten electrons, and nine neutrons
- d. An atom made from eight protons, eight electrons, and eight neutrons



**Chapter 2 Review Questions**

1. Define the following terms:
  - a. Pure substance
  - b. Mixture
  - c. Homogeneous mixture
  - d. Heterogeneous mixture
  - e. Law of Mass Conservation
  - f. Element
  - g. Compound
  - h. Law of Definite Proportions

i. Molecule

j. Law of Multiple Proportions

k. Isotopes

2. Classify the following as an element, compound, heterogeneous mixture, or homogeneous mixture.

a. A bowl of fruit covered with yogurt

b. A sample of helium gas, which cannot be broken down into simpler substances

c. A sample of sugar thoroughly dissolved in water

- d. A sample of sodium bicarbonate (baking soda), which can be broken down into hydrogen, carbon, oxygen, and sodium.
  
  - e. Several grams of magnesium, which is one of the two simplest substances produced when magnesium oxide breaks down.
  
  - f. The magnesium oxide from which the magnesium discussed above was produced.
  
  - g. A strawberry
  
  - h. A cup of tea with no leaves in it.
3. A student does a chemical reaction with two chemicals. The total mass of the two chemicals is 45.0 grams. When she is done, she finds that the mass of all the chemicals she has collected is now only 34.5 grams. Has she collected all the products of the reaction? How do you know?
4. In the situation discussed in problem #3, what is the mass of the chemical or chemicals the student didn't collect?

5. Water is a compound that can be broken down into hydrogen gas and oxygen gas. However, a mixture of hydrogen gas and oxygen gas looks and behaves nothing like water. Why?

6. A 75.0-gram sample of a white powder is chemically broken down into 29.86 grams of copper, 15.06 grams of sulfur, and an unknown amount of oxygen gas.
- How much oxygen gas was made?
  - Suppose you want to make 11.2 grams of the white powder with no leftovers. How much copper, sulfur, and oxygen would you have to use?
7. A chemist makes 86.94 grams of a black powder by reacting 54.94 grams of manganese with 32.00 grams of oxygen. If a student breaks down 144.9 grams of that powder, what mass of manganese and what mass of oxygen will be made?





13. A chemist is making two different compounds from the same two elements: nitrogen and oxygen. To make the first gas, she reacts 10.0 grams of nitrogen and 11.42 grams of oxygen to make 21.42 grams of the first gas. Suppose she starts with 10.0 grams of nitrogen again but wants to make a completely different gas with no leftover oxygen. Should she use 20.50 grams of oxygen, 22.84 grams of oxygen, or 35.12 grams of oxygen?





- a. An atom made of 13 protons, 14 neutrons, and 13 electrons
- b. An atom made of 14 protons, 14 neutrons, and 14 electrons
- c. An atom made of 13 protons, 12 neutrons, and 13 electrons
- d. An atom made of 15 protons, 16 neutrons, and 15 electrons

20. Which of the two isotopes you found in #19 would be the heaviest?

21. When you burn a fuel, what besides the fuel gets used up?

22. What particles do you find in the nucleus of an atom?

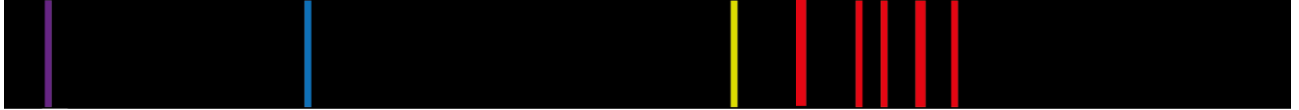
**Chapter 3 Comprehension Check Questions**

1. How many protons are in an atom of Aluminum?
2. What is the mass of the average chlorine atom in grams?
3. How many protons, electrons, and neutrons are found in the following atoms?
  - a.  $^{18}\text{F}$
  - b.  $^{40}\text{Ar}$
  - c. sodium-23
  - d. Hg-200
4. What is the most abundant isotope of sulfur?



8. What is the wavelength of a photon with an energy of  $6.81 \times 10^{-17}$  J? ( $c = 3.00 \times 10^8$  m/sec,  $h = 6.63 \times 10^{-34}$  J·sec.)
9. An electron moves from the  $n = 2$  orbit of a hydrogen atom to the  $n = 1$  orbit. The wavelength of light it emits to do this is 122 nm. Suppose the electron then gains enough energy to go to the  $n = 5$  orbit. When it moves directly back to the  $n = 1$  orbit, will the wavelength of light it emits be longer or shorter than 122 nm?
10. An electron jumps to the  $n = 4$  orbit. How many wavelengths of light could it possibly emit to get back to its ground state?

11. A sample containing two elements is excited, and the light it emits is analyzed. The wavelengths emitted look like this:



Use the illustration on page 84 to determine which two elements are in the sample.



4. How many protons, neutrons, and electrons are in a magnesium-25 atom?
5. An atom has 5 protons, 5 neutrons, and 5 electrons. What is its symbol (including the mass number)?
6. What is the most abundant isotope of phosphorus?
7. What is the most difficult thing to do when it comes to making a nuclear bomb?





13. To what does the term “electromagnetic waves” refer?

14. Name three categories of electromagnetic waves, other than visible light.

15. A photon has a frequency of  $1.94 \times 10^{16}$  Hz. What is its energy?

16. A photon has an energy of  $1.19 \times 10^{-16}$  J. What is its frequency?

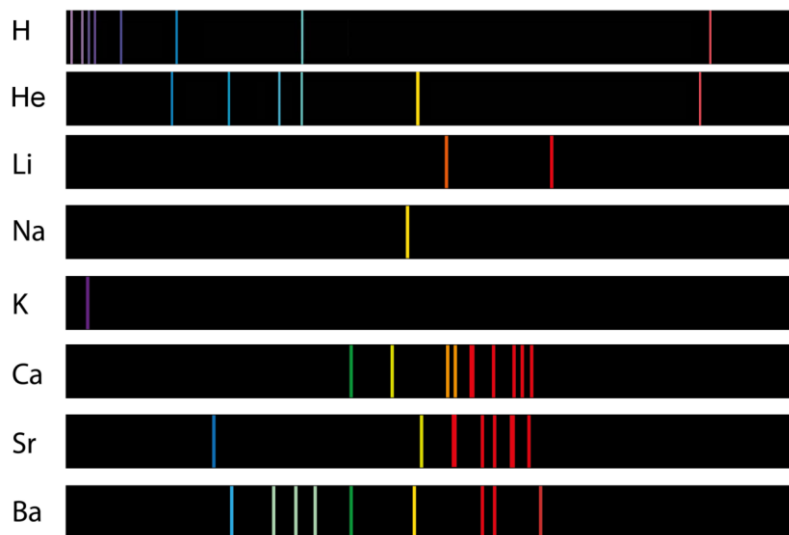
17. Orange light has a wavelength of 612 nm. What energy does a photon of this light have?

18. A photon has an energy of  $2.17 \times 10^{-18}$  J. What is its wavelength?



22. Fill in the blanks: Atoms are more than 99.999% \_\_\_\_\_.

23. Consider the emission spectra shown below:



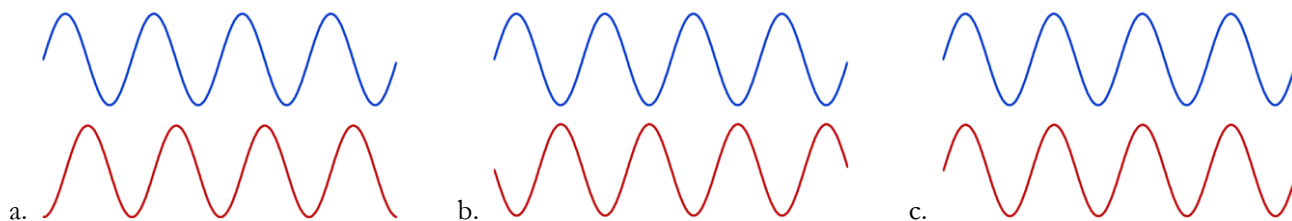
If a mixture of two elements has the following emission spectrum:



what elements are present in the mixture?

**Chapter 4 Comprehension Check Questions**

1. Two waves overlap as shown in the three pictures below. Which will result in constructive interference? Which will result in destructive interference? Which will result in something in-between?



2. An electron is in the second energy level. What types of orbitals are available to it? How many total orbitals are available to it?
3. An electron occupies a d orbital. What is its minimum energy level (1, 2, 3, 4, 5, etc.)?
4. a. Give the full electron configuration for oxygen in its ground state.
- b. Give the full electron configuration for rubidium (Rb) in its ground state.

5. Give the abbreviated electron configuration for arsenic (As) in its ground state.

6. Which of the following elements have similar chemical properties?

- a. K      b. Sc      c. Mg      d. Cl      e. Sr      f. Si

7. Draw the Lewis structures for the following elements:

a. potassium

b. carbon

c. arsenic (As)

d. iodine (I)

8. For the following elements, indicate the ions they will form to attain an ideal electron configuration. You don't have to make the Lewis structure. Just use the symbol and indicate the name of the ion.
- a. bromine (Br)
  - b. oxygen
  - c. calcium
  - d. cesium (Cs)
9. Metals are very good at losing electrons, while nonmetals are very good at gaining them. How well do you think the metalloids do these tasks?
10. Give the name and chemical formula of the ionic compound formed between the following elements.
- a. lithium and sulfur
  - b. calcium and chlorine
  - c. aluminum and nitrogen



11. Which of the following pairs of elements can react to form an ionic compound?

- a. phosphorus and chlorine      b. strontium (Sr) and fluorine      c. carbon and iodine (I)

12. What is the chemical formula of aluminum oxide?

13. Manganese (Mn) can have more than one charge in ionic compounds. What is the name of  $\text{MnO}$ ?

**Chapter 4 Review Questions**

1. Define the following terms:

a. Wave-particle duality

b. Aufbau principle

c. Valence electrons

d. Ion

e. Ionic compound

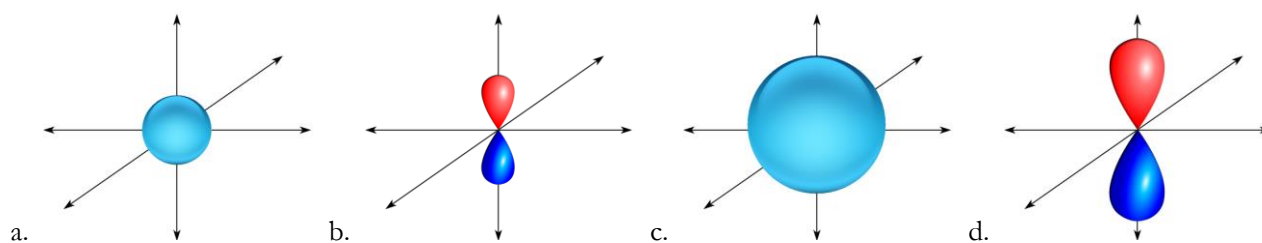
f. Electrolysis

g. Electrolytes

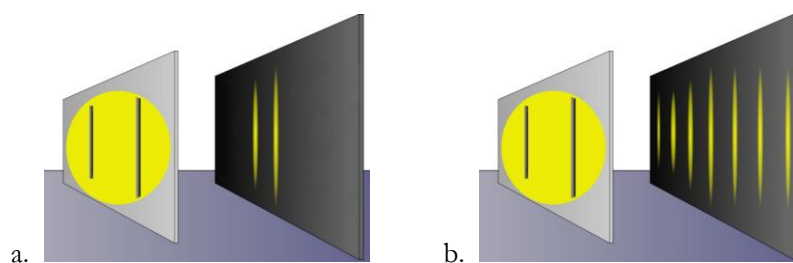
2. What do we call the current model of the atom?

3. When waves overlap so that their crests get higher and troughs get deeper, what is that called?
  
  
  
  
  
  
  
  
  
  
4. What is the opposite of your answer to #3.
  
  
  
  
  
  
  
  
  
  
5. Fill in the blank: In order to come up with his model of the atom, Erwin Schrödinger analyzed the electrons as if they were \_\_\_\_\_.
  
  
  
  
  
  
  
  
  
  
6. List all the orbitals available to electrons in the third energy level of an atom. Indicate how many of each orbital type exist.
  
  
  
  
  
  
  
  
  
  
7. How many electrons can a single orbital hold?

8. One of the following orbitals represents a 2p orbital, and the other represents a 3p orbital. Indicate which one represents each.



9. Sunlight from a pinhole is shined on two small slits that are very closely-spaced. Which of the following is the result?



10. If electrons were used in #9 instead of light, which of the two patterns shown in #9 would they produce?

11. Give the ground-state electron configuration for the following elements:

a. nitrogen

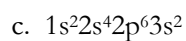
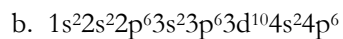
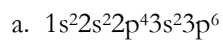
b. aluminum

c. magnesium

d. bromine (Br)

12. Give the abbreviated ground-state electron configuration for iodine (I).

13. What is wrong with the following electron configurations?



14. What element has the electron configuration  $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$ ?

15. What do we call the elements in Group 8A on the Periodic Table of Elements?

16. Which of the following elements should have very similar chemical behaviors?

- a. Ca      b. Se      c. Al      d. Pb      e. In

17. Draw the Lewis structure for oxygen.





21. What are the names and chemical formulas of the compounds formed between the following elements?
- Rb and oxygen
  - Sr and sulfur
  - calcium and nitrogen
  - magnesium and I
22. What is the chemical formula of calcium fluoride?
23. Nitrogen and hydrogen react to form a compound. Is it an ionic compound?
24. What is the chemical formula of iron (III) oxide? (Iron's chemical symbol is Fe.)
25. Chromium (Cr) can take on more than one charge in an ionic compound. What is the name of  $\text{CrCl}_3$ ?

26. Which of the following compounds, when dissolved in water, is likely to produce a solution that does not conduct electricity?

- a. NaCl      b. KHCO<sub>3</sub>      c. C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>      d. Rb<sub>2</sub>SO<sub>4</sub>

**Chapter 5 Comprehension Check Questions**

1. Draw the Lewis structure for bromine as it actually exists. Use a line to indicate the bond.

2. Draw the Lewis structures for the following compounds:

a. HCl

b. CH<sub>3</sub>Br

c. K<sub>2</sub>O

d. SF<sub>2</sub>

3. What are the Lewis structures for the following compounds?

a. CO

b. CO<sub>2</sub>

4. What is the name of the compound  $\text{CCl}_4$ ?

5. What is the chemical formula of silicon monoxide?

**It is best to use only the Periodic Table to answer the questions 6&7, because you won't be given the charts on pages 137 and 138 when you take the test!:**

6. Order the following elements in terms of their atomic radii, from lowest to highest:

Ga, K, Br, As, Sc

7. Order the following elements in terms of their electronegativity, from lowest to highest:

I, At, Cl, Br, F

8. Construct the Lewis structure for  $\text{NF}_3$ . Does it have polar covalent bonds? If so, where is the slight positive charge, and where is the slight negative charge?

9. The electronegativity of sulfur is 2.5, the electronegativity of chlorine is 3.0, and the electronegativity of iodine is 2.5. Which combinations of any two of these elements will produce a polar covalent bond?
10. What is the shape of  $\text{CH}_2\text{Cl}_2$ ? Draw a stick figure for it.

11. What is the shape of CO?

12. What is the shape of SF<sub>2</sub>? Draw a stick figure for it.

13. Is ammonia (NH<sub>3</sub>) polar or nonpolar?

14. Is  $\text{SiF}_4$  polar or nonpolar?

15. In problems 13 and 14, you determined whether or not  $\text{NH}_3$  and  $\text{SiF}_4$  were polar or nonpolar. Indicate which one has a good chance of dissolving in water, and which has a good chance of dissolving in oil.



**Chapter 5 Review Questions**

1. Define the following terms:

a. Covalent bond

b. Homonuclear diatomic

c. Covalent compound

d. Atomic radius

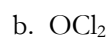
e. Electronegativity

f. Polar covalent bond

g. Polar covalent molecule

h. Purely covalent bond

- i. Purely covalent molecule
  
2. List the seven homonuclear diatomic elements you need to remember.
  
  
  
  
  
  
  
  
  
  
3. Draw the Lewis structures for the following molecules:



c.  $\text{NH}_2$

d.  $\text{S}_2$

e.  $\text{CS}_2$

f.  $\text{N}_2$

g.  $\text{HSiP}$

4. Name the following molecules:

a.  $\text{SO}_2$

b.  $\text{PH}_3$

c.  $\text{K}_2\text{O}$

d.  $\text{P}_2\text{O}_5$

e.  $\text{N}_2\text{O}_3$

5. Which of the following atoms is the largest? Cu, Cs, Ca, Co
  
  
  
  
  
  
  
  
  
  
6. Which of the following atoms is the most electronegative? Ga, Ba, As, S, O
  
  
  
  
  
  
  
  
  
  
7. Which of the following molecules will have a purely covalent bond in it? HI, CF<sub>4</sub>, I<sub>2</sub>
  
  
  
  
  
  
  
  
  
  
8. Is it possible for a molecule to be composed of polar covalent bonds and still be a purely covalent molecule?
  
  
  
  
  
  
  
  
  
  
9. Which kind of molecule (ionic, polar covalent, or purely covalent) has charges that are free to move around?
  
  
  
  
  
  
  
  
  
  
10. Which kind of molecule (ionic, polar covalent, or purely covalent) has charges that are weaker than the charges on a proton and electron?
  
  
  
  
  
  
  
  
  
  
11. Determine the shape and draw a stick figure for the following molecules:
  - a. SiCl<sub>4</sub>

b.  $\text{H}_2\text{S}$

c.  $\text{GeSe}_2$

d.  $\text{PF}_3$

e.  $\text{H}_2\text{SiO}$

f.  $\text{HBr}$

12. For each of the molecules above, indicate whether it is polar covalent or purely covalent.

13. For each of the molecules in #11, indicate whether or not it is likely to dissolve in water.

14. Fill in the blank: The active ingredient in soap is a long molecule with one end that is ionic, and one end that is \_\_\_\_\_.



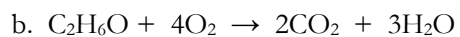
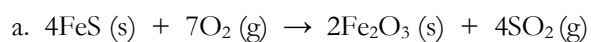
**Chapter 6 Comprehension Check Questions**

1. Identify each of the following changes as physical or chemical:
  - a. Crushing an aluminum can
  - b. Digesting food
  - c. Cutting paper
  - d. Melting ice
  - e. Making a hard-boiled egg
  - f. Exploding a bomb
  
2. At the highest parts of the interior of Antarctica, the average temperature is about  $-76^{\circ}\text{F}$ . What is that in Celsius?

3. Suppose you put one finger in a glass of water at room temperature and another finger in a glass of alcohol at the same temperature. Will one finger feel cooler than the other? If so, which one? Suppose you then pull both fingers out of the glass at the same time. Will one finger feel cooler than the other? If so, which one?
4. The heating curve of an unknown substance shows that the temperature rises steadily, then stays constant at 15 °C for a while. The temperature then starts increasing again for a while. However, at 234 °C, the temperature stays constant for quite some time. At what temperature does this substance melt? At what temperature does it vaporize?

5. The utility companies that supply natural gas for burning in stoves, furnaces, etc., store it in its liquid phase. It is then converted into its gas phase to be delivered to the customer. Why do they store it in its liquid phase, when the customers need it in its gas phase?

6. Determine whether or not the following chemical equations are balanced:



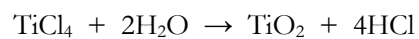
7. Nitrogen gas and hydrogen gas react to make nitrogen trihydride (also called “ammonia”) gas. What is the balanced chemical equation?

8. Many outdoor gas grills burn propane ( $C_3H_8$ ) gas. The propane reacts with oxygen in the air, making gaseous carbon dioxide and water vapor. What is the balanced chemical equation for this reaction?

9. Balance the following unbalanced equation:  $H_2 + CO \rightarrow C_8H_{18} + H_2O$

10. Solid iron (II) sulfide reacts with oxygen in the air to make solid iron (III) oxide and gaseous sulfur dioxide. What is the balanced chemical equation, including phase symbols? (Iron's symbol is "Fe.")

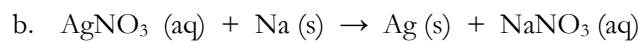
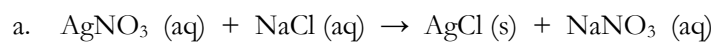
11. In the reaction given by the following balanced equation:



Indicate how many of each molecule are used up and how many of each molecule are made.

12. Give the balanced chemical equation for the formation of  $\text{HNO}_3$ .

13. Identify the following equations as single displacement or double displacement:



14. Give the balanced chemical equation for the complete combustion of hexane ( $\text{C}_6\text{H}_{14}$ ).

**Chapter 6 Review Questions**

1. Define the following terms:
  - a. Physical Change
  - b. Chemical change
  - c. Kinetic Theory of Matter
  - d. Formation reaction
  - e. Decomposition Reaction
  - f. Single displacement reaction
  - g. Double displacement reaction
  - h. Combustion reaction

- i. Complete combustion
  
  
  
  
  
  
  
  
  
  
  - j. Incomplete combustion
- 
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2. Identify the following as chemical or physical change
    - a. A chunk of sodium is sliced in half.
  
  
  
  
  
  
  
  
  
  
    - b. A chunk of sodium is thrown in water, making sodium hydroxide and hydrogen
  
  
  
  
  
  
  
  
  
  
    - c. Water evaporates from a glass.
  
  
  
  
  
  
  
  
  
  
    - d. Soap is added to oil and water to make the oil spread out in the water.

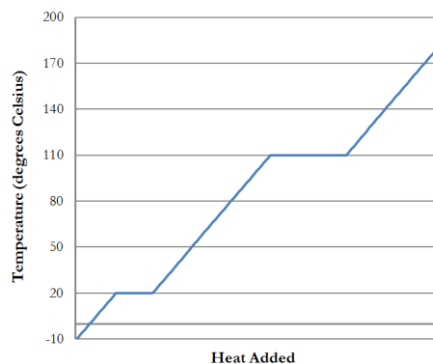


- e. Pancakes cook on a griddle.
  
  - f. Solid baking soda is added to vinegar, forming lots of bubbles.
3. Table salt (sodium chloride) melts at a temperature of 1,474 °F. What is that temperature in °C?
4. Summer daytime temperatures in parts of Australia can reach 40.0 °C. What is that in °F?
5. A comet develops a tail when it gets close to the sun. The tail is the result of frozen chemicals on the comet turning into gases. What word do we use to refer to that process?
6. A gas condenses on a surface. Does the substance gain energy or lose energy when this happens? Does the surface gain energy or lose energy?

7. How does sweating cool the skin?
8. Suppose you had a thermometer in an environment where a gas condenses on the thermometer's bulb. When that happens, will the temperature of the thermometer decrease, increase, or stay the same?
9. The temperature of a liquid is 115 °F, and the temperature of another sample of the same liquid is 70.0 °C. Which liquid's molecules are moving faster?

10. A chemist adds 12 mL of alcohol to 11 mL of a liquid called acetone. Assuming no chemical reaction occurs, which of the following is the most likely final volume: 30 mL, 23 mL, or 22 mL?

11. The graph on the right depicts the heating curve of the dark and sinister substance known as Wileium.



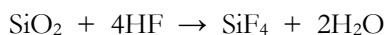
a. At what temperature does Wileium melt?

b. At what temperature does Wileium boil?

c. What is the phase of Wileium at 80 °C?

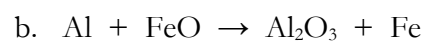
12. Solid iron sinks in liquid iron. Solid silicon floats in liquid silicon. Which of those elements behaves like water when it freezes?

13. For the following chemical reaction, identify the reactants and the products:

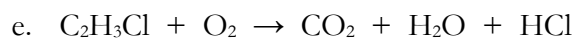
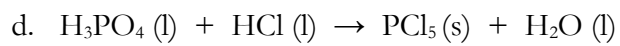


14. Balance the following chemical equations:

a. Solid zinc reacts with liquid hydrogen monochloride to make solid zinc (II) chloride and gaseous hydrogen. (The symbol for zinc is “Zn.”)



c. Calcium reacts with aluminum chloride to make calcium chloride and aluminum



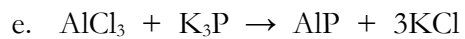
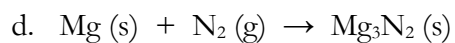
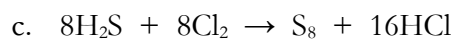
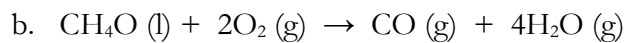
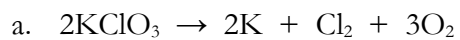
15. Give balanced chemical equations for the following:

a. The formation of  $\text{CaCO}_3$ .

b. The decomposition of  $\text{K}_2\text{CrO}_4$ .

c. The complete combustion of gaseous  $C_2H_2$ .

16. Identify the following chemical reactions as formation, decomposition, single displacement, double displacement, complete combustion, or incomplete combustion.



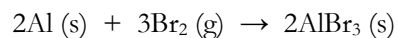
**Chapter 7 Comprehension Check Questions**

1. Suppose you need to collect a mole of magnesium atoms. How many grams of magnesium would you need?
2. If I have 0.500 moles of sulfur, how many atoms of sulfur do I have? How many grams of sulfur would that be?
3. How many grams must you gather to have 8.10 moles of  $\text{H}_2\text{O}$ ? How many molecules would you have?
4. If you have 309 grams of ammonia ( $\text{NH}_3$ ), how many moles is that?

5. Sodium carbonate ( $\text{Na}_2\text{CO}_3$ ) has a hydrated and anhydrous form. If 14.2 grams of anhydrous sodium carbonate can absorb 24.0 grams of water, how many water molecules can be absorbed by each  $\text{Na}_2\text{CO}_3$ ? How would that be noted in the chemical formula of the compound?



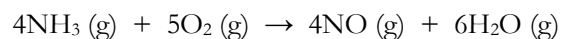
6. The balanced chemical equation for the formation of aluminum bromide is:



How many moles of Al and Br<sub>2</sub> would you need to make 0.407 moles of AlBr<sub>3</sub>?

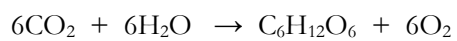
7. In your experiment, which reactant (acetic acid or baking soda) was in excess for the bottle that had the balloon with about 2 grams of baking soda in it? Which reactant was in excess for the bottle that had the balloon with about 6 grams of baking soda in it?

8. Ammonia (NH<sub>3</sub>) can be burned according to the following reaction:



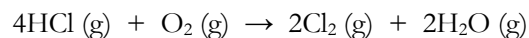
If 115.9 moles of ammonia are burned in excess oxygen, how many moles of H<sub>2</sub>O are made?

9. Plants make their own food with photosynthesis, which boils down to the following equation:



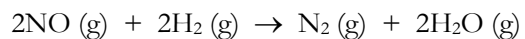
If a plant needs 67.1 moles of  $\text{C}_6\text{H}_{12}\text{O}_6$  for food, how many moles of  $\text{CO}_2$  will it need to absorb?

10. Hydrogen monochloride can be burned in the following reaction:



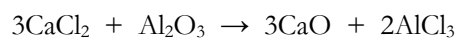
If 15.0 moles of HCl are burned in excess oxygen, how many grams of chlorine gas are made?

11. Nitrogen monoxide can be converted into elemental nitrogen with the addition of hydrogen:



How many moles of nitrogen monoxide are needed to produce 855 grams of nitrogen?

12. Consider the following double displacement reaction:



If you have 100.0 grams of  $\text{CaCl}_2$ , how many grams of  $\text{Al}_2\text{O}_3$  should you add to make sure there are no reactants left at the end of the reaction?

**Chapter 7 Review Questions**

1. Define the following terms:
  - a. Hygroscopic
  - b. Limiting reactant
2. How many molecules are in a mole?
3. How many atoms are in 58.93 grams of cobalt (Co)?
4. Indicate how many moles are contained in each sample below:

a. A 100.0-g sample of iron (Fe)

b. 250.0 grams of methane (CH<sub>4</sub>)

c. 500.0 grams of potassium iodate (KIO<sub>3</sub>)

5. Determine the mass of each sample below:
- 14.5 moles of strontium (Sr)
  - 0.045 moles of calcium sulfide (CaS)
  - 0.876 moles of magnesium sulfate (MgSO<sub>4</sub>)
6. How many atoms exist in 100.0 grams of lithium?

7. How many molecules exist in 50.0 grams of phosphorus trihydride ( $\text{PH}_3$ )?



8. A chemist has  $1.0 \times 10^{22}$  molecules of water. How many moles is that? How many grams is that?
9. The hydrate of magnesium sulfate ( $\text{MgSO}_4$ ) has seven water molecules for every magnesium sulfate molecule. Write the chemical formula of hydrated magnesium sulfate, then write the chemical formula of the anhydrous form.
10. Suppose a student has 100.0 grams of the anhydrous magnesium sulfate discussed in problem #9. How many grams of water could it absorb?

11. If the anhydrous form of cobalt chloride absorbs water to become the hydrated form, is that a chemical or physical change?

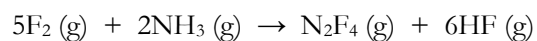
12. The formation reaction for methanol ( $\text{CH}_4\text{O}$ ) is given below:



If you want to make 100.0 moles of  $\text{CH}_4\text{O}$ , how many moles of hydrogen would you need? How many moles of oxygen would you need?

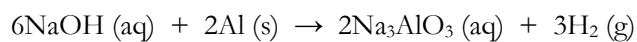
13. A chemist makes methanol according to the equation given in problem #12. She starts with 24.0 grams of carbon, 10.0 grams of hydrogen, and 40.0 grams of oxygen. In the end, she makes methanol, but there are 2.0 grams of hydrogen and 8.0 grams of oxygen left over. What was the limiting reactant?

14. A chemist is making  $\text{N}_2\text{F}_4$  according to the following chemical equation:



If he adds 150 moles of  $\text{F}_2(\text{g})$  to an excess of  $\text{NH}_3(\text{g})$ , how many moles of  $\text{N}_2\text{F}_4(\text{g})$  will he make?

15. Consider the following reaction:



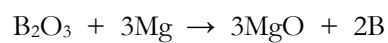
If 150.0 moles of  $\text{NaOH}$  are added to an excess of  $\text{Al}$ , how many grams of  $\text{Na}_3\text{AlO}_3$  will be made?

16. How many moles of Na would be needed to make 500.0 grams of  $\text{Na}_2\text{S}$  according to the following equation?



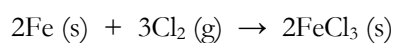
17. Suppose a chemist wants to make  $\text{Na}_2\text{S}$  according to the equation in problem #16. She adds 15 moles of Na to 1 mole of  $\text{S}_8$ . Which is the limiting reactant?

18. A chemist is using the following reaction



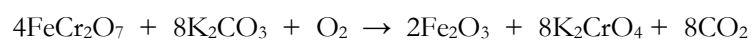
to make boron. If he wants 125 grams of boron, how many grams of magnesium will he need?

19. A student is making iron (III) chloride according to the reaction:



If he adds 50.0 grams of  $\text{Cl}_2$  to an excess of Fe, how many grams of  $\text{FeCl}_3$  will he make? How many grams of Fe actually react with the  $\text{Cl}_2$ ?

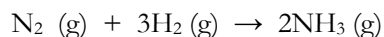
20. A chemist is running the following chemical reaction:



He has plenty of oxygen and  $\text{K}_2\text{CO}_3$ , but he has only 115 grams of  $\text{FeCr}_2\text{O}_7$ . How much  $\text{K}_2\text{CrO}_4$  can he make?

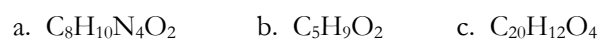
**Chapter 8 Comprehension Check Questions**

1. A chemist uses stoichiometry to calculate that 10.0 grams of product should be made. Then he runs the experiment three times. The first time, he gets 9.4 grams of product. The second time, he gets 10.9 grams of product. The third time, he gets 8.5 grams of product. Which of his experiments had the most experimental error? Which had the least?
2. Suppose you forgot to put the medicine dropper on the watch glass for the final measurement of mass in step 19 of your experiment. Would that cause the measured mass of carbon dioxide produced to be too high or too low?
3. A chemist wants to make ammonia ( $\text{NH}_3$ ) using 500.0 grams of nitrogen and an excess of hydrogen.



If the chemist has a percent yield of 85.1%, how much ammonia was actually made?

4. For each formula below, identify it as molecular or empirical. If it is not an empirical formula, give the empirical formula.



5. A compound has an empirical formula of CH. If its molar mass is 78.12 g, what is its molecular formula?



6. What is the molecular formula of aluminum oxide?
7. A 50.0-gram sample of titanium (Ti) is burned in excess oxygen. It produces one product with a mass of 83.3 g. What is the empirical formula of the product?
8. 100.0 grams of an unknown substance containing only carbon and hydrogen were burned. If the complete combustion produced 313.7 grams of carbon dioxide and 128.4 grams of water, what is the empirical formula of the unknown substance?

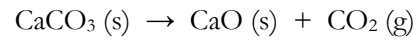
- 100.0 grams of an unknown substance are burned. If the complete combustion produced 149.1 grams of carbon dioxide and 45.8 grams of water, what is the empirical formula of the unknown substance?

- What is the percent composition for each element in the compound  $\text{H}_3\text{PO}_4$ ?

11. An unknown substance is 62.2% iron (Fe), 35.6% oxygen, and 2.2% hydrogen. What is its empirical formula?
12. What is the name of  $K_2SO_4$ ?
13. What is the chemical formula of Aluminum carbonate? How many oxygen atoms are in that empirical formula?
14. What is the name of  $FePO_4$ ? Please note that iron can have different charges, so you need to include Roman numerals in the name. You should be able to figure out the charge of the iron atom so you can do that.

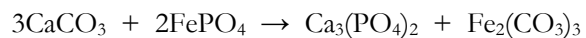
**Chapter 8 Review Questions**

1. Define the following terms:
  - a. Experimental error
  - b. Empirical formula
  - c. Molecular formula
  - d. Polyatomic ion
2. Three students produce sodium nitrate using exactly the same process. The first has a percent yield of 76%, the second's percent yield is 104%, and the third's percent yield is 87%. Which student has the lowest experimental error?
3. What two strategies discussed in the text can be used to reduce experimental error?
4. A student heats calcium carbonate to produce calcium oxide, according to the following equation:



If the student starts with 100.0 grams of  $\text{CaCO}_3$  and makes 45.0 g of  $\text{CaO}$ , what is his percent yield?

5. A student performs the following chemical reaction:



She starts with 500.0 g of  $\text{CaCO}_3$  and an excess of  $\text{FePO}_4$ . If she ends up with a 91.1% yield, how many grams of  $\text{Ca}_3(\text{PO}_4)_2$  did she end up making?

6. Which of the following is the chemical formula for a carbohydrate?

a.  $\text{C}_5\text{H}_{10}\text{O}$       b.  $\text{CHO}$       c.  $\text{C}_7\text{H}_{14}\text{O}_7$       d.  $\text{C}_2\text{H}_4\text{O}$

7. Which of the following is an empirical formula?

a.  $\text{K}_2\text{SO}_4$       b.  $\text{P}_2\text{O}_6$       c.  $\text{C}_5\text{H}_{10}\text{O}$       d.  $\text{H}_2\text{O}_2$       e.  $\text{C}_6\text{H}_{12}\text{N}_3\text{O}_6$

8. For any chemical formula in #7 that is not an empirical formula, what is the empirical formula for that compound?
  
  
  
  
  
  
  
  
  
  
9. What kind of compound can have both empirical and molecular formulas? What kind of compound cannot have a molecular formula?
  
  
  
  
  
  
  
  
  
  
10. A compound has an empirical formula of  $C_3H_7NO$  and a molar mass of 146.22 g. What is its molecular formula?

11. A 15.0-g sample of nitrogen is burned in excess oxygen. The result is a single gas with a mass of 49.3 grams. What is the empirical formula of the gas?
12. A liquid hydrocarbon (a molecule containing only carbon and hydrogen) is burned in a combustion analysis experiment. If 450.0 grams of liquid produce 1,362.5 grams of carbon dioxide and 697.4 grams of water, what is the empirical formula?
13. A 200.0-g sample of an unknown substance was burned in excess oxygen to produce 439.3 grams of carbon dioxide and 239.9 grams of water. There was no other product. What is the empirical formula of the unknown substance? In a separate experiment, a single molecule of the substance was found to have a mass of 180.33 amu. What is the molecular formula of the substance?



14. What is the percent composition for each element in  $\text{Na}_3\text{PO}_4$ ?

15. A sample of an unknown substance is decomposed and found to be 24.7% potassium, 34.7% manganese (Mn), and 40.6% oxygen. What is the empirical formula?

16. A 500.0-gram sample of an unknown substance is decomposed into 144.2 g of magnesium, 71.2 g of carbon, and some oxygen that escaped and could not be measured. What is the empirical formula?

17. Give the chemical formulas for the following substances:

a. Potassium sulfate

b. strontium carbonate

c. ammonium phosphate

d. magnesium nitrate

18. How many phosphorus and oxygen atoms are in the empirical formula of magnesium phosphate?

19. Manganese (Mn) can have different charges in an ionic compound. What is the name of  $\text{MnCO}_3$ ?

20. What is wrong with the following statement?

Ionic compounds do not have any physical bonds.

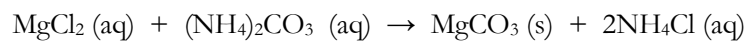
**Chapter 9 Comprehension Check Questions**

1. The following solids are dissolved in water. For each solute, identify what specific substances are in the solution and whether or not the solution will conduct electricity.
  - a.  $\text{K}_2\text{CO}_3$
  
  
  
  
  
  
  
  
  
  
  - b.  $\text{SCl}_2$
  
  
  
  
  
  
  
  
  
  
  - c.  $\text{C}_2\text{H}_4\text{O}_2$
  
  
  
  
  
  
  
  
  
  
  - d.  $\text{NH}_4\text{NO}_3$
  
  
  
  
  
  
  
  
  
  
2. You have a saturated solution of ammonia, a gaseous solute.
  - a. If temperature is decreased, what happens? Is the resulting solution still saturated?
  
  
  
  
  
  
  
  
  
  
  - b. If pressure is decreased, what happens? Is the resulting solution still saturated?

3. You have a solution of sodium bromide, a solid solute. It is not saturated, but you want to make it saturated. Unfortunately, you have no more sodium bromide. What can you do to the temperature in order to make the solution saturated?
4. Calcium phosphate is insoluble in water. If a solution of potassium phosphate is mixed with a solution of calcium chloride, what is the chemical equation for the reaction that occurs?
5. A student measures out 15.67 grams of ammonium carbonate and dissolves it in enough water to make 200.0 mL of solution. What is the molarity of the solution?

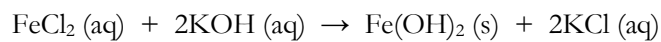
6. If the student takes the solution made in the previous problem and uses 34.2 mL in an experiment, how many moles of ammonium carbonate were used?

7. A chemist needs to make 450.0 grams of magnesium carbonate according to the following reaction:



If she has a 1.5 M solution of magnesium chloride, what volume will she need to use?

8. A chemist performs the following reaction:



If he uses 156.7 mL of a 1.34 M solution of  $\text{FeCl}_2$  and an excess of  $\text{KOH}$ , what mass of  $\text{Fe}(\text{OH})_2$  should be produced?

9. If you dissolve 125.0 grams of KBr in 750.0 g of water, what is the molality?

10. A chemist needs a solution of  $\text{Sr}(\text{NO}_3)_2$  that has a concentration of 0.75 m. If she uses 350.0 grams of water, how many grams of  $\text{Sr}(\text{NO}_3)_2$  must she use?
11. The freezing point of acetic acid is 16.6 °C. If you dissolve 50.0 grams of glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) into 500.0 grams of acetic acid, what is the freezing point of the solution? ( $K_f$  for acetic acid is 3.90 °C/m.)
12. A chemist wants to lower the freezing point of some water to -5.6 °C. If he has 568.0 grams of water, how many grams of  $\text{Na}_2\text{CO}_3$  should he add? ( $K_f$  for water is 1.86 °C/m.)



13. A saltwater solution is made with 35.0 grams of NaCl and an unknown mass of water. If the boiling point of the solution is 102.0 °C, how many kg of water were used in the making of the solution? ( $K_b$  for water is 0.512 °C/m)

## Chapter 9 Review Questions

1. Define the following terms:

a. Solution

b. Solute

c. Solvent

d. Solubility

e. Saturated solution

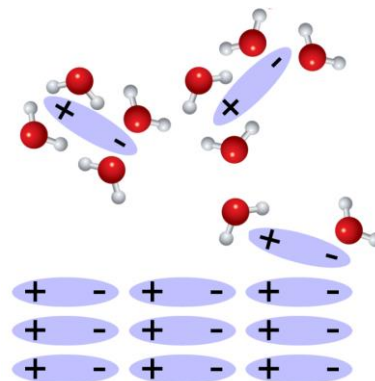
f. Precipitate

g. Concentration

2. Bronze is a homogeneous mixture of mostly copper with small amounts of another metal, such as tin. Identify the solute and solvent. What do we call solutions made from a metal and other solids?

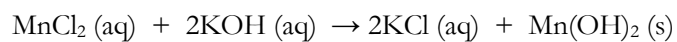
3. Which will conduct electricity: A solution of water and  $(\text{NH}_4)_2\text{S}$  or a solution of water and  $\text{C}_6\text{H}_{12}\text{O}_6$ ?

4. If you were able to look at a solution of  $\text{K}_3\text{PO}_4$  (aq) and see all the molecules and ions, what specific molecules and ions would you see? Of those, which make up most of the solution? Which make up the least amount of the solution?



5. The illustration on the right attempts to show how a polar covalent molecule dissolves in water. What is wrong with it?

6. An aqueous solution contains three sulfate ions for every two aluminum ions. What is the solute?
7. You are making a solution by dissolving solid  $\text{AgNO}_3$  in water. What can you do with temperature to increase the solubility of the  $\text{AgNO}_3$ ?
8. You are trying to dissolve as much of a gas into water as you can possibly dissolve. How should you adjust the pressure and temperature?
9. A solution of a solid solute is saturated. If the temperature is lowered, what will happen to the solution?
10. A student performs the following reaction by mixing a solution of  $\text{MnCl}_2$  with a solution of  $\text{KOH}$ :

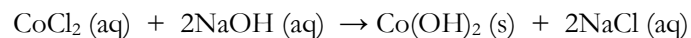


Did a precipitate form when he mixed the solutions? If so, what is the precipitate?

11. Calcium carbonate is insoluble in water. If a solution of calcium nitrate and water is mixed with a solution of potassium carbonate and water, what reaction will occur?

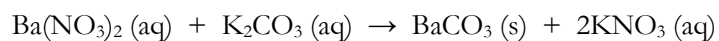
12. A solution is made by dissolving 100.0 g of  $\text{Li}_2\text{SO}_4$  in enough water to make 325.0 mL of solution. What is the concentration in moles/liter?
13. You want to make 225.0 mL of a 4.5 M solution of NaOH. How many grams of NaOH will you need?

14. A chemist wants to make  $\text{Co}(\text{OH})_2$  according to the following reaction:



If the student uses 167.4 mL of a 1.39 M solution of NaOH and an excess of  $\text{CoCl}_2$ , how many grams of  $\text{Co}(\text{OH})_2$  will be formed?

15. A chemist wants to make 150.0 grams of  $\text{BaCO}_3$  using the following reaction:



If she has an excess of  $\text{K}_2\text{CO}_3$ , what volume of a 2.56 M solution of  $\text{Ba}(\text{NO}_3)_2$  will she need to use?

16. Of the two concentration units you learned in this chapter, which is more like a ratio of the amount of solute to the amount of solvent?
17. You make a solution by mixing 67.8 grams of KCl with 450.0 grams of water. What is the molality of the solution?
18. A chemist makes a 2.34 m solution by mixing 23.1 g of  $\text{MgCl}_2$  with water. How many kilograms of water were used?

19. Solutions are made from water and the following solutes:  $\text{Mg}(\text{NO}_3)_2$ ,  $\text{CaCl}_2$ ,  $\text{NaCl}$ ,  $\text{NH}_3$ ,  $(\text{NH}_4)_2\text{CO}_3$ , and  $\text{AlCl}_3$ . If the molalities of the solutions are all the same, which will have the lowest freezing point?
20. What is the freezing point of a solution made by mixing 32.0 g of  $\text{NaCl}$  with 250.0 grams of water? ( $K_f$  is  $1.86\text{ }^\circ\text{C}/\text{m}$  for water.)
21. A solution made from water and  $\text{N}_2\text{H}_4$  has a freezing point of  $-1.5\text{ }^\circ\text{C}$ . What is the molality of the solution? ( $K_f$  is  $1.86\text{ }^\circ\text{C}/\text{m}$  for water.)
22. What is the boiling point of the solution in #20? ( $K_b$  is  $0.512\text{ }^\circ\text{C}/\text{m}$  for water.)



23. The boiling point of ethanol is  $78.37\text{ }^{\circ}\text{C}$ . If the boiling point of a solution of  $\text{SrCl}_2$  and ethanol is  $80.56\text{ }^{\circ}\text{C}$ , what is the molality of the solution? ( $K_b$  for ethanol is  $1.07\text{ }^{\circ}\text{C}/m$ )

**Chapter 10 Comprehension Check Questions**

1. A barometer reads a pressure of 750.1 torr. How many atmospheres (atm) is that? Is it above or below the average atmospheric pressure found at sea level?
2. A chemical reaction is run at a temperature of 156.7 K. What is the temperature in degrees Celsius and degrees Fahrenheit?
3. A gas is compressed under constant temperature. It starts out with a volume of 150.1 mL and ends up with a volume of 75.6 mL. If the initial pressure was 750 torr, what is the final pressure?
4. A balloon is floating on the surface of a lake. Its volume is 3.91 liters, its pressure is 1.09 atm, and its temperature is 28.4 °C. It is then pulled to the bottom of the lake, where the pressure is 3.70 atm. If its new volume is 1.09 liters, what is the temperature at the bottom of the lake in °C?

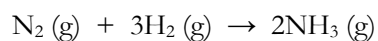
5. In each of the following cases, choose the situation that would result in the gas behaving most ideally.

- a. (1) one mole of gas at 275 K or (2) one mole of gas at 350 K
  
- b. (1) one mole of gas at 740 torr or (2) one mole of gas at 1,500 torr
  
- c. (1) one mole of gas in a very small container or (2) one mole of gas in a much larger container
  
- d. (1) one mole of  $\text{H}_2$  gas or (2) one mole of  $\text{C}_{12}\text{H}_4\text{Cl}_4\text{O}_2$  gas under the same conditions

6. A sample of gas contains 18.45 moles at STP. What volume does it occupy?

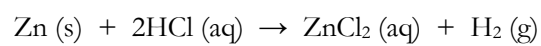
7. A sample of gas that contains 12.5 moles is at a pressure of 735.6 torr. If it occupies a volume of 0.45 kiloliters, what is its temperature?

8. A chemist is making ammonia gas according to the following equation:



If she starts with 150.0 liters of  $\text{N}_2$  and an excess of  $\text{H}_2$  at STP, what volume of  $\text{NH}_3$  will be made at STP?

9. A chemist is making hydrogen gas by reacting zinc and HCl:



If he has an excess of zinc and 850.0 mL of 5.0 M HCl, what volume of H<sub>2</sub> will be made at STP?

10. In a mixture of two gases, the partial pressure of oxygen is 341 torr, while the partial pressure of nitrogen is 761 torr. If there are a total of 6.78 moles of gas, how many moles of oxygen and nitrogen are present?
11. Suppose you are camping at Leavitt Lake in California, which is at an altitude of about 3,050 meters. The atmospheric pressure there is 522 torr. Will water boil at a temperature above 100 °C, between 90 and 100 °C, between 60 °C and 90 °C, or below 60 °C? (Use the table on page 311 to answer this question.)

**Chapter 10 Review Questions**

1. Define the following terms:

- a. Boyle's Law
- b. Charles's Law
- c. Extrapolation
- d. Avogadro's Law
- e. Boiling point

2. What is a Pascal?

3. What does a barometer measure?

4. Why are inches of mercury and millimeters of mercury used as units for pressure?

5. If you put an inflated balloon into a container and then add lots more pressure to the container without changing the temperature, what will happen to the size of the balloon?
  
  
  
  
  
  
  
  
  
  
6. If you inflate a balloon at room temperature and then put it in a freezer that is at the same pressure, what will happen to the size of the balloon?
  
  
  
  
  
  
  
  
  
  
7. What does it mean for a temperature scale to be an absolute temperature scale?
  
  
  
  
  
  
  
  
  
  
8. What is the boiling point of water at 1 atm in Kelvin?



9. Nitrogen condenses at 77 K. What is that temperature in degrees Celsius?
10. Fill in the blank: When extrapolating, a good scientist keeps the range of extrapolation \_\_\_\_\_ compared to the range over which the data have been measured.
11. Why must you use an absolute temperature scale like the Kelvin scale when working with the gas laws?
12. A piston contains 45.6 mL of gas at a pressure of 814 torr. If the temperature doesn't change, what is the volume of the gas at a pressure of 760.0 torr?
13. A flexible container is held at constant pressure. It starts with a volume of 1.2 liters at a temperature of 25.0 °C. What is the volume if the temperature rises to 50.0 °C?

14. A helium-filled balloon has a volume of 35.4 liters at an altitude where the temperature is 0.00 °C and the pressure is 0.619 atm. It is brought back to the ground, where the temperature is 22.90 °C, and its volume is 21.1 liters. What is the pressure on the ground?

15. In each case, indicate which of the two situations will be most likely to produce ideal behavior in a gas:

a. (1)  $T = 124 \text{ K}$ ,  $P = 1 \text{ atm}$  or (2)  $T = 280 \text{ K}$ ,  $P = 1 \text{ atm}$

b. (1)  $T = 275 \text{ K}$ ,  $P = 700 \text{ torr}$  or (2)  $T = 275 \text{ K}$ ,  $P = 4,567 \text{ torr}$

c. (1) STP or (2)  $T = 100 \text{ K}$ ,  $P = 15 \text{ atm}$

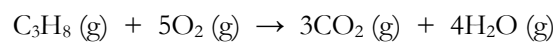
16. How many moles of nitrogen are present in a 891-mL container at 23.0 °C and 345 torr.

17. What is the volume of 10.0 grams of water vapor at 134.5 °C and a pressure of 567 torr?

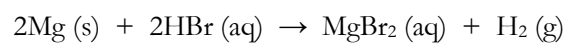
18. In the following decomposition of 150.0 grams of potassium chlorate ( $\text{KClO}_3$ ), what volume of oxygen is made at STP?



19. A chemist burns 567 liters of propane ( $\text{C}_3\text{H}_8$ ). How many liters of carbon dioxide are made at 1.04 atm and  $1,995\text{ }^\circ\text{C}$ ?



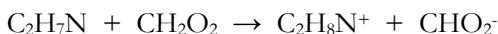
20. A balloon is filled with 345 torr of oxygen gas and 416 torr of nitrogen gas. What is the mole fraction of each gas?
21. A container holds 0.34 moles of chlorine gas, 0.22 moles of xenon gas, and 0.56 moles of krypton gas. If the total pressure of the gases is 1.1 atm, what is the partial pressure of each gas?
22. A camper boils water to cook dinner. He cooks the same dinner two nights in a row, but it takes longer for the same amount of the same dinner to cook on the second night. Compare the atmospheric pressure he experiences on the two nights.
23. A chemist makes hydrogen gas according to the equation:



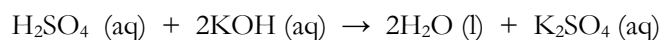
She collects the hydrogen over the water in which the reaction is taking place. If she collects 3.41 liters of hydrogen at 0.981 atm and 60.0 °C, how many moles of hydrogen gas did she make? (You can use the table on page 311 for this.)

**Chapter 11 Comprehension Check Questions**

1. Arm and Hammer makes a product called “washing soda” that is supposed to aid your laundry detergent so that your clothes get even cleaner. Based on what you learned in the experiment, do you expect washing soda to be an acid or a base?
2. Milk tends to “go sour” if left for too long. Compare the acid content in fresh milk and sour milk.
3. In the following chemical reaction, which reactant is the acid, and which reactant is the base?



4. For the following reaction, determine the acid and base:



5. For each of the following compounds, indicate whether it is an acid, base, or neither. If it is an acid, indicate how many  $\text{H}^+$  ions it can donate:  $\text{CsOH}$ ,  $\text{HI}$ ,  $\text{H}_2\text{CrO}_4$ ,  $\text{NaBr}$ ,  $\text{Sr}(\text{OH})_2$ , and  $\text{HC}_3\text{H}_5\text{O}_2$ .

6. What is the equation for the chemical reaction between  $\text{HClO}_3$  and  $\text{NH}_3$ ?

7. What is the equation for the chemical reaction between  $\text{H}_2\text{CO}_3$  and  $\text{Al}(\text{OH})_3$ ?

8. Five solutions are labelled with the following pH's: 1, 3, 7, 9, 14.

a. Which is the most acidic?

b. Which is neutral?

c. Which is the most basic?

d. Which one might be a baking soda solution?

e. Which one would turn anthocyanins pink?

f. Which one could be distilled water?



9. A student needs to neutralize 750.0 mL of a 5.00-M solution of KOH. How many mL of a 3.41-M solution of sulfuric acid ( $\text{H}_2\text{SO}_4$ ) will be needed to get the job done?
10. A solution of  $\text{H}_3\text{PO}_4$  has an unknown concentration. 115.0 mL of the solution are titrated against a 1.12-M solution of LiOH. If 97.1 mL of the base are required to reach the endpoint of the titration, what is the concentration of the acid?
11. A chemist has a 9.12-M solution of  $\text{MgCl}_2$ . To do an experiment, he needs 500.0 mL of a 4.50-M solution of  $\text{MgCl}_2$ . How can he make that out of the solution he already has?

**Chapter 11 Review Questions**

1. Define the following terms:

a. Indicator

b. Acid

c. Base

d. Amphoteric

e. Polyprotic acid

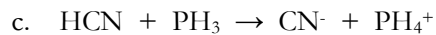
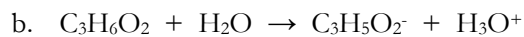
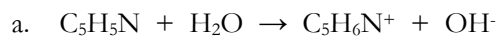
f. Acid/Base Titration

2. What are the three characteristics of acids that you learned in this chapter?

3. What are the three characteristics of bases that you learned in this chapter?

4. Name the two acid/base indicators you learned about in this chapter.
5. Is lye an acid, base, or neither?
6. Fill in the blanks with either acids or bases: \_\_\_\_\_ are found in many of the things we eat and drink (like fruits, milk, and soda pop), while \_\_\_\_\_ are found in many household cleaning products (like soap, laundry detergent, and bleach).

7. Identify the acid and base in the following chemical reactions:



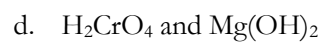
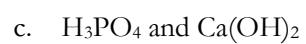
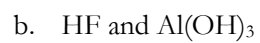
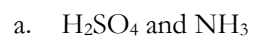
8. Give an example of an amphoteric chemical.

9. Identify the following as acids or bases:  $\text{NH}_3$ ,  $\text{HNO}_3$ ,  $\text{HC}_3\text{H}_5\text{O}_2$ ,  $\text{LiOH}$ ,  $\text{Zn}(\text{OH})_2$

10. Suppose you are told by a chemist that citric acid ( $\text{C}_6\text{H}_8\text{O}_7$ ) is an acid that can donate three  $\text{H}^+$  ions. How could you rewrite the chemical formula to make citric acid more recognizable as a polyprotic acid?

11.  $\text{HBr}$  is a strong enough acid that it will make  $\text{H}_2\text{O}$  act as a base. What is the chemical reaction between  $\text{HBr}$  and  $\text{H}_2\text{O}$ ?

12. Determine the reaction that occurs between the following chemicals:



13. You see five clear solutions labeled with the following pH's: 0.1, 4.5, 7.0, 9.0, 14.0.

- a. One of them is a solution of NaCl and water. Which one?
  
  
  
  
  
  
  
  
  
  
- b. One of them is a very concentrated solution of a strong acid. Which one?
  
  
  
  
  
  
  
  
  
  
- c. One of them is a sour but pleasant drink. Which one?
  
  
  
  
  
  
  
  
  
  
- d. One of them is a glass cleaner you can buy at the supermarket. Which one?
  
  
  
  
  
  
  
  
  
  
- e. One of them is a very caustic base. Which one?

14. What volume of a 3.17-M solution of  $\text{H}_2\text{SO}_4$  would be required to neutralize 750.0 mL of a 2.12-M solution of NaOH?

15. A student needs to neutralize 1.5 liters of a 5.6-M HCl solution. How many grams of  $\text{Al}(\text{OH})_3$  should be used?

16. What is wrong with the following statement: “The endpoint of a titration occurs when the number of moles of acid equals the number of moles of base.”

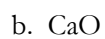
17. Bromthymol blue is an indicator that is yellow from a pH of 0 to about 6.8. It is green from a pH of 6.9 to a pH of 8.2, and it is blue from a pH of 8.3 to a pH of 14. You are told to use it as an indicator for an acid/base titration. The titration starts with an acid in a flask, and you are adding base to it during the titration. What will be the initial color of the solution in the flask? What color change will you be looking for to tell you when you have reached the endpoint?
18. 15.0 mL of a solution of  $\text{H}_2\text{CrO}_4$  of unknown concentration is titrated with a 1.16-M solution of KOH. If it takes 13.1 mL of KOH to reach the endpoint, what is the concentration of the  $\text{H}_2\text{CrO}_4$ ?
19. About 10% of the mass of Lime-A-Way toilet bowl cleaner is its main active ingredient, HCl. Muriatic acid is sold in hardware stores. About 30% of its mass is HCl. Which one (Lime-A-Way or Muriatic acid) is better at getting rid of lime stains? Which one is safer to use?
20. Concentrated acetic acid is 17.4 M. How would you make 250.0 mL of a 3.5-M acetic acid solution from concentrated acetic acid?



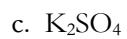
21. A chemist makes 50.0 mL of an NaCl solution with a concentration of 2.50 M. A student wants to use that to make a 1.75-M solution. What volume of 1.75-M solution can he make if he uses all of the original solution?

**Chapter 12 Comprehension Check Questions**

1. Give the oxidation state for each atom or ion in the following compounds:



2. Give the oxidation state for each atom in the following compounds:



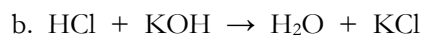
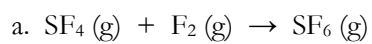
d.  $\text{Cl}_2$

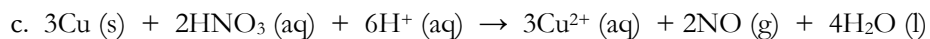
e.  $\text{NF}_3$

f.  $\text{CH}_4$

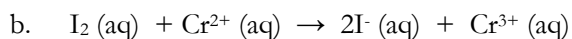
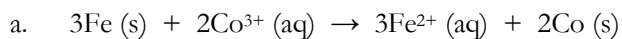
g.  $\text{PCl}_3$

3. For each of the following reactions, indicate whether or not it is a redox reaction. If it is, indicate what was oxidized and what was reduced. Also, indicate the oxidizing agent and the reducing agent.

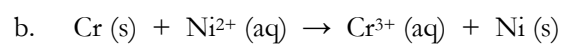
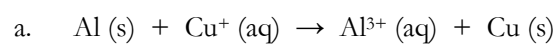




4. One side of a Galvanic cell has magnesium metal (Mg) becoming magnesium ions ( $\text{Mg}^{2+}$ ). Is it the anode or the cathode?
5. A student makes a Galvanic cell that runs really well for a while. After a while, however, the cell stops working. There are plenty of reactants in each side of the cell, and the electrical connection between both electrodes is fine. What, most likely, is the problem?
6. Galvanic cells are powered by the following reactions. Draw each cell, labeling the solutions on each side of the cell, the anode, and the cathode. Also, indicate the direction in which the electrons will flow in the cell.



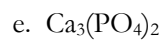
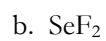
7. Balance the following redox reactions.



8. Identify the reactant found on the positive side of a lead-acid battery. Identify the substance found at the negative side.
9. The inside of a dry cell has a pH of 13.9. Is it a zinc-carbon battery, an alkaline battery, or a lead-acid battery?
10. Suppose you have a piece of gold jewelry, and you want to determine if it is all gold or just electroplated with gold. How could you tell without destroying the jewelry? (HINT: It has to do with something you learned about in Chapter 1.)

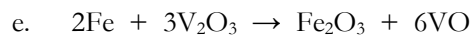
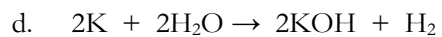
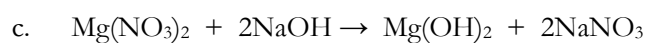
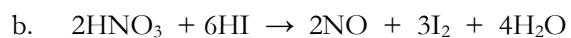
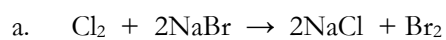
**Chapter 12 Review Questions**

1. Define the following terms:
  - a. Oxidation state
  - b. Galvanic Cell
  - c. Anode
  - d. Cathode
  - e. Electroplating
  - f. Corrosion
  
2. Using the definition of oxidation state (not the rules we developed), what is the oxidation state of each atom in IBr?
  
  
  
  
  
  
  
  
  
  
3. Using the rules you learned in this chapter, give the oxidation states for every atom in the following substances:



4. For each of the following reactions, indicate whether or not it is a redox reaction. If it is, indicate the atom being oxidized and the atom being reduced. Also, indicate the oxidizing agent and the reducing agent.

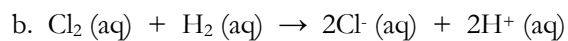
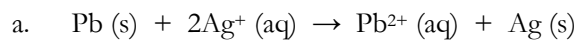




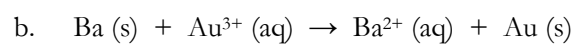
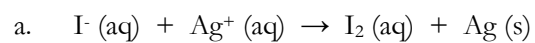
5. Fill in the blank: Batteries are based on \_\_\_\_\_ cells.

6. What does the salt bridge do in a Galvanic cell?

7. One side of a Galvanic cell has  $\text{Fe}^{3+}$  becoming  $\text{Fe (s)}$ .
- Is this the anode or the cathode of the cell?
  - Are electrons flowing towards the electrode or away from it on this side of the cell?
  - In the salt bridge, are positive or negative ions flowing towards this side of the cell?
8. The reactions that run two Galvanic cells are given below. Draw the cell, identifying the anode and cathode, the electron flow, and the substances on each side of the cell.



9. Balance the following redox reactions.



10. A student says that since the battery he is using is a dry cell, it contains no water. Is the student correct? Why or why not?

11. In a zinc-carbon battery, is carbon the anode or the cathode?

12. Which battery was the first to be developed, the zinc-carbon battery or the alkaline battery?



**Chapter 13 Comprehension Check Questions**

1. Two solid blocks of equal temperature are placed in contact with one another. Is heat produced?
2. In a mixture of ice and water, the ice is melting. The temperature remains the same the entire time, however. Is there heat in the system?
3. When I go to McDonald's, I usually get a Quarter Pounder with cheese, which contains about 500 Calories. If I used the energy in my sandwich to heat up 10 liters of water that is initially at 25 °C, what would its final temperature be?
4. Equal amounts of energy are removed from equal masses of the following elements: silver, lead, zinc, copper, aluminum. Which element gets the coolest? Which stays the warmest?
5. If a 1.5-kg object cools down from 22.5 °C to 8.1 °C when it loses 6.3 kJ of energy, what is its specific heat capacity in J/g·°C?

- An object made of nickel is heated. Its temperature rises from  $23.4\text{ }^{\circ}\text{C}$  to  $44.3\text{ }^{\circ}\text{C}$  when  $987\text{ J}$  of energy are added to it. What is the mass of the object? ( $c = 0.440\text{ J/g}\cdot^{\circ}\text{C}$ )

7. A 20.0-gram sample of an unknown metal is used in a calorimetry experiment. The calorimeter holds 115.0 grams of water that starts off at a temperature of 24.1 °C. The metal is heated to 110.0 °C and dropped in the calorimeter. The experiment ends when the water reaches a temperature of 26.3 °C. What is the specific heat capacity of the metal? Ignore the effects of the calorimeter.
8. A calorimeter is filled with 150.0 grams of an unknown liquid that has a temperature of 23.7 °C. A 25.0-gram sample of copper ( $c = 0.386 \text{ J/g}\cdot\text{°C}$ ) is heated to 70.0 °C and dropped into the calorimeter. If the final temperature of the experiment is 25.3 °C, what is the specific heat capacity of the liquid? The calorimeter's effects can be ignored.



9. A 300.0-gram sample of an unknown metal is used in a calorimetry experiment. The calorimeter holds 1,200.0 grams of water that starts off at a temperature of 25.0 °C. The metal is heated to 136.3 °C and dropped in the calorimeter. The experiment ends when the water reaches a temperature of 27.3 °C. If the heat capacity of the calorimeter is 1,100 J/°C, what is the specific heat capacity of the metal?

10. A 300.0-gram sample of ice at 0.0 °C is converted into water vapor at 100.0 °C. How much heat is absorbed by the sample? ( $L_f = 334 \text{ J/g}$ ,  $L_v = 2,260 \text{ J/g}$ )
11. When it comes to how kinetic and potential energy change, which kind of reaction (endothermic or exothermic) is similar to a skier climbing up a hill?

**Chapter 13 Review Questions**

1. Define the following terms:

- a. Energy
- b. Work
- c. calorie
- d. Specific heat capacity
- e. Latent heat
- f. Exothermic reaction
- g. Endothermic reaction
- h. Kinetic energy
- i. Potential energy

2. Fill in the blank: \_\_\_\_\_ is the study of heat that is absorbed or released by chemical processes.

3. What is wrong with the following statement: If there is no temperature change, there is no heat.
4. Which of the following units would be acceptable for measuring heat? Calories, Pascals, Joules
5. A Hostess Twinkie contains 270 Calories. How many grams of water could you warm  $1\text{ }^{\circ}\text{C}$  with that energy?

6. Glass has a specific heat capacity of  $0.840 \text{ J/g}\cdot^{\circ}\text{C}$ , solid magnesium has a specific heat capacity of  $1.020 \text{ J/g}\cdot^{\circ}\text{C}$ , and solid silver has a specific heat capacity of  $0.240 \text{ J/g}\cdot^{\circ}\text{C}$ . If I add equal amounts of heat to equal masses of these substances at the same initial temperature, which gets the warmest? Which stays the coolest?
7. Which of the following is a unit for specific heat capacity, and which is a unit for heat capacity?

Joules,  $\frac{\text{J}}{^{\circ}\text{C}}$ , grams,  $\frac{\text{J}}{\text{g}\cdot^{\circ}\text{C}}$ ,  $^{\circ}\text{C}$ , Calories

8. A 2.5-kg sample of copper ( $c = 0.386 \text{ J/g}\cdot^{\circ}\text{C}$ ) at  $24.1^{\circ}\text{C}$  experiences a change in temperature to  $19.7^{\circ}\text{C}$ . How much energy is involved? Did the copper gain or lose the energy?
9. 500.0 g of water in its liquid phase loses 5,892 J of heat. If it started at a temperature of  $25.0^{\circ}\text{C}$ , what is its final temperature?

10. 175.0 grams of an unknown metal gain 5.6 kJ of heat. The metal's temperature rises from 24.3 °C to 110.5 °C. What is the specific heat capacity of the metal in J/g·°C?
11. Why does a calorimeter need to be insulated?
12. Why does a calorimeter need to be stirred?
13. In a calorimetry experiment, what two things start off at the same temperature? What three things end up at the same temperature?
14. A calorimetry experiment is performed on a 112.1-g sample of metal. It is heated to a temperature of 175.0 °C and put in a calorimeter that contains 250.0 g of water. If the water's temperature rises from 27.1 °C to 31.7 °C, what is the specific heat capacity of the metal? Ignore the calorimeter in your calculation.

15. A 145.1-g sample of tin ( $c = 0.210 \text{ J/g}\cdot^\circ\text{C}$ ) is heated to  $150.0^\circ\text{C}$  and dropped in a calorimeter that contains an unknown mass of water. If the water starts out at  $22.6^\circ\text{C}$  and ends up at  $30.0^\circ\text{C}$ , what mass of water was in the calorimeter? Ignore the calorimeter in your calculations.
16. An object has a heat capacity of  $16.7 \text{ J}/^\circ\text{C}$ . How much heat must it absorb to increase its temperature by  $5.0^\circ\text{C}$ ?
17. A metal object that has a mass of 315.0 grams is heated to  $317.0^\circ\text{C}$  and dropped into a calorimeter that has a heat capacity of  $1,300 \text{ J}/^\circ\text{C}$ . The calorimeter contains 500.0 grams of water at an initial temperature of  $25.6^\circ\text{C}$ . If the water's temperature rises to  $31.5^\circ\text{C}$  by the end of the experiment, what is the specific heat capacity of the metal?

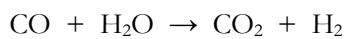
18. An iron ( $c = 0.444 \text{ J/g}\cdot^\circ\text{C}$ ) object that has a mass of 275.0 grams is heated to  $165.0^\circ\text{C}$  and dropped into a calorimeter that has a heat capacity of  $1,300 \text{ J}/^\circ\text{C}$ . The calorimeter contains 500.0 grams of an unknown liquid at a temperature of  $45.3^\circ\text{C}$ . If the liquid's temperature rises to  $51.5^\circ\text{C}$  by the end of the experiment, what is the specific heat capacity of the liquid?
19. How much energy is required to vaporize 4.51 kg of ethanol ( $L_f = 109 \text{ J/g}$ ,  $L_v = 838 \text{ J/g}$ )?
20. A 136.0-g ice cube at  $-15.0^\circ\text{C}$  is heated until it is liquid with a temperature of  $50.0^\circ\text{C}$ . How much energy did it absorb? ( $c_{\text{ice}} = 2.093 \text{ J/g}\cdot^\circ\text{C}$ ,  $L_f = 334 \text{ J/g}$ ,  $L_v = 2,260 \text{ J/g}$ )



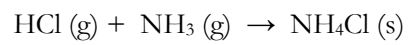
21. A chemical reaction occurs in a beaker, and the beaker gets cold. Is this an endothermic or exothermic reaction?
22. When it comes to chemicals, where is the potential energy stored?
23. Is combustion exothermic or endothermic? Is the heat you feel from the fire kinetic or potential energy?

**Chapter 14 Comprehension Check Questions**

- The  $\Delta H$ 's for three chemical reactions are measured. For reaction A,  $\Delta H = -1.73$  kJ/mole. For reaction B,  $\Delta H = 5.14$  kJ/mole. For reaction C,  $\Delta H = -16.8$  kJ/mole.
  - Identify each reaction as exothermic or endothermic.
  - Assuming equal moles of the reactants in each, which reaction would feel the hottest?
- Use bond energies to determine the  $\Delta H$  for the following reaction. Is it exothermic or endothermic?

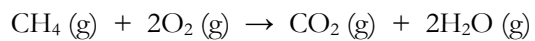


- Use Hess's Law to determine the  $\Delta H$  for the following reaction at 25°C:

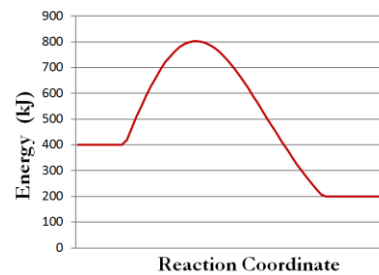
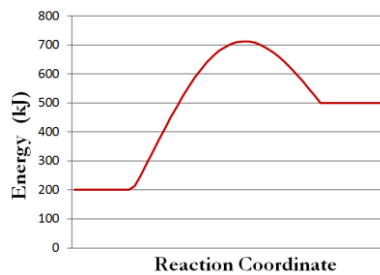
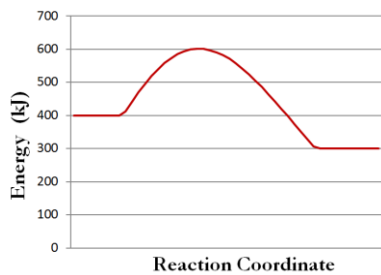


4. Use Hess's Law to determine the  $\Delta H$  for the complete combustion of  $\text{C}_2\text{H}_6$  (g) at 25 °C.

5. How much energy is produced when 250.0 g of methane ( $\text{CH}_4$ ) burns in excess oxygen?

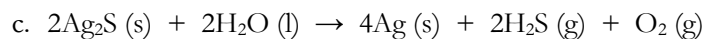
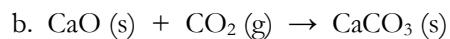
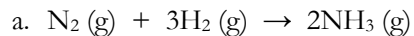


6. Consider the three reaction coordinate diagrams below:



- Which represents the endothermic reaction?
- Which represents the reaction that releases the most energy?
- Which represents the reaction that is easiest to start?

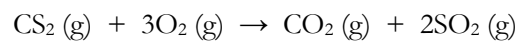
7. When you did the calorimetry experiment in this chapter, you assumed that all energy released by the chemical reaction had to go into the calorimeter or the water. What law of thermodynamics were you using to make that assumption?
8. For each letter below, choose the situation that represents the higher entropy.
- a. 100 molecules at 100 °C or 200 molecules at 200 °C
  - b. A mole of water vapor at 100 °C or a mole of liquid water at 100 °C
  - c. 500 molecules of gas in a 1-liter container or 500 molecules of gas in a 2-liter container at the same temperature.
9. Some creationists say that the Second Law of Thermodynamics forbids evolution, because evolution requires an increase in the complexity of an organism, which would require a decrease in the organism's entropy. Based on the material you just learned, what's wrong with that argument?
10. Determine the sign of  $\Delta S$  for each of the following reactions:



11. Determine the  $\Delta S$  for the formation of ammonia:  $\text{N}_2 (\text{g}) + 3\text{H}_2 (\text{g}) \rightarrow 2\text{NH}_3 (\text{g})$

12. A chemical reaction has a  $\Delta H$  of  $-131.6 \text{ kJ}$  and a  $\Delta S$  of  $-112 \text{ J/K}$  at  $25^\circ\text{C}$ . Is it spontaneous at  $25^\circ\text{C}$ ? What is the temperature range for which it is spontaneous?

13. We know that the following chemical reaction is spontaneous at  $25^\circ\text{C}$ . Give a range for the possible values of the  $\Delta G_f^\circ$  for  $\text{SO}_2 (\text{g})$ .



**Chapter 14 Review Questions**

1. Define the following terms:

a. Change in enthalpy

b. Bond energy

c. State function

d. Activation energy

e. Thermodynamics

f. First Law of Thermodynamics

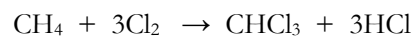
g. Entropy

h. Second Law of Thermodynamics

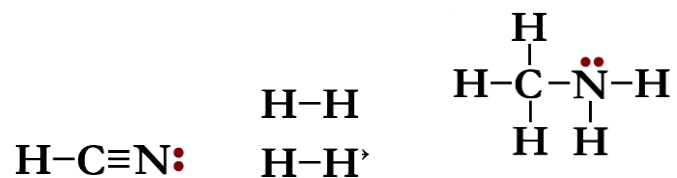


2. A reaction is determined to be spontaneous at 25 °C. Does that mean it will automatically happen when you mix the reactants at 25 °C?
3. A reaction has a positive  $\Delta H$ . Is it exothermic or endothermic?
4. Two reactions occur, and each causes the beaker in which it is done to get hot. The first one gets much hotter than the second, however. The same number of moles of reactant are used in each. If the  $\Delta H$ 's of the reactions are -355 kJ/mole and -214 kJ/mole, which belongs to the first reaction?

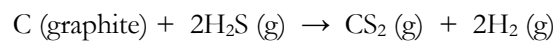
5. Using the bond energies table on page 420, determine the  $\Delta H$  of the following reaction:



6. The  $\Delta H$  of the following reaction is  $-137 \text{ kJ/mole}$ . What is the energy of the  $\text{C}\equiv\text{N}$  bond?



7. What is the standard enthalpy of formation of Fe (s)?
8. What is the  $\Delta H$  for the complete combustion of  $C_6H_6$  (l) at  $25^\circ C$ ? Use Hess's Law.
9. The  $\Delta H$  for the following reaction is 157.9 kJ at  $25^\circ C$ . What is the  $\Delta H_f^\circ$  of  $H_2S$  (l)? (NOTE: graphite is the natural phase for carbon at  $25^\circ C$ .)



10. Given the information in problem 9, how much energy would it take to make 100.0 grams of  $\text{CS}_2$  (g) from the reaction between carbon and  $\text{H}_2\text{S}$ ?

11. Draw a reaction coordinate diagram for each of the following situations. Don't put numbers on the energy axis. Leave the energy axis unnumbered like the reaction coordinate axis.

a. An exothermic reaction that has a high activation energy and releases a small amount of energy.

b. An endothermic reaction with an activation energy that is only a bit larger than its  $\Delta H$ .

c. A reaction that is much easier to start than the reaction in (a) but releases more energy.

12. For each letter below, choose the situation that has more entropy:

a. A vase sitting on a shelf or that same vase broken into many pieces

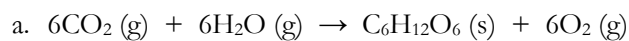
b. A sample of water as a liquid at 0°C or the same mass of water as an ice cube at 0°C

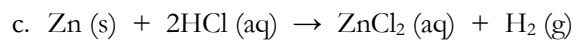
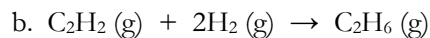
c. A copper pipe that is cold or the same copper pipe that is hot

d. A bunch of confetti scattered evenly on the floor of a 3 meter  $\times$  3 meter room or the same amount of confetti scattered evenly on the floor of a 2 meter  $\times$  2 meter room

13. A student says that because a chemical reaction has a negative  $\Delta S$ , it cannot be spontaneous. Is he correct? Why or why not?

14. Determine the sign of  $\Delta S$  for the following reactions:





15. In problem 8, you calculated the  $\Delta\text{H}$  for the complete combustion of  $\text{C}_6\text{H}_6(\text{l})$  at  $25^\circ\text{C}$ . Determine the  $\Delta\text{S}$  for that same reaction.

16. Use the Hess's Law technique to calculate the  $\Delta\text{G}$  of the complete combustion of  $\text{C}_6\text{H}_6(\text{l})$  at  $25^\circ\text{C}$ . Is the reaction spontaneous?

17. The  $\Delta\text{H}$  and  $\Delta\text{S}$  of several reactions are given below. For each reaction, determine whether it is spontaneous at all temperatures, never spontaneous no matter what the temperature, or spontaneous for a range of temperatures. If it is spontaneous for a range of temperatures, give the range.

a.  $\Delta H = 146.1 \text{ kJ}, \Delta S = 86.1 \text{ J/K}$

b.  $\Delta H = -115.2 \text{ kJ}, \Delta S = 97.3 \text{ J/K}$

c.  $\Delta H = 219.5 \text{ kJ}, \Delta S = -99.9 \text{ J/K}$

d.  $\Delta H = -419.4 \text{ kJ}, \Delta S = -114.1 \text{ J/K}$

18. Which of the following is not zero for an element in its natural phase:  $\Delta H_f^\circ$ ,  $S^\circ$ ,  $\Delta G_f^\circ$ ?

19. You and a friend are going to a store. You each leave at the same time and take a different route to get there. You arrive before your friend. Which of the following is a state function: the path you traveled, the time it took you to get there, or your destination?

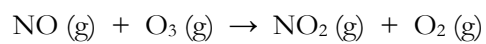
20. A physical process in a system cools its surroundings and is spontaneous. Does it result in an increase or decrease of entropy for the system?

**Chapter 15 Comprehension Check Questions**

1. In a study of kinetics, the concentration of one substance is measured to be 1.15 M at  $t = 0.0$  seconds. At  $t = 15.0$  seconds, the concentration is measured to be 0.98 M. Is the substance a reactant or product? What is the rate of the reaction over this time interval?
2. In each of the situations given below, indicate which will result in the lower reaction rate:
  - a. Using large chunks of a solid in a reaction or using the same mass of solid in powdered form
  - b. Performing an experiment at 15 °C or performing the same experiment at 50 °C.
  - c. Doing an acid/base reaction with 5.0-M base or doing the same reaction with 1.0-M base.
3. A chemical reaction has only two reactants. The order of the reaction with respect to the first reactant is 1, and the overall order is 1.
  - a. What is the order of the reaction with respect to the second reactant?
  - b. What are the units on the rate constant?



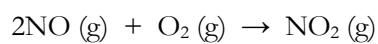
4. A chemist is experimenting with the following reaction:



The following data are collected. Determine the rate equation, including the rate constant.

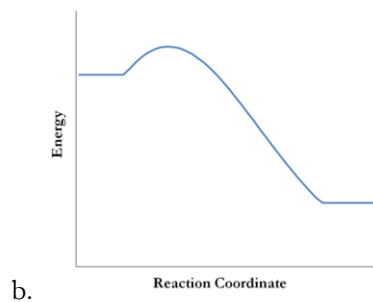
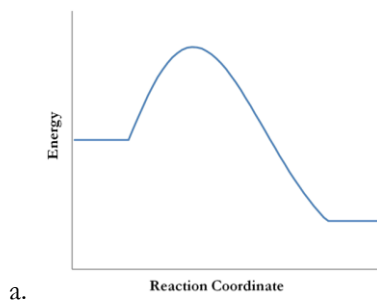
<b>Trial</b>	<b>[NO]</b>	<b>[O<sub>3</sub>]</b>	<b>Instantaneous Rate</b>
<b>1</b>	0.0010 M	0.0010 M	22.0 M/s
<b>2</b>	0.0010 M	0.0020 M	44.0 M/s
<b>3</b>	0.0020 M	0.0010 M	44.0 M/s

5. A series of trials are run on the reaction below. Given the data in the table, determine the rate equation, including the rate constant.



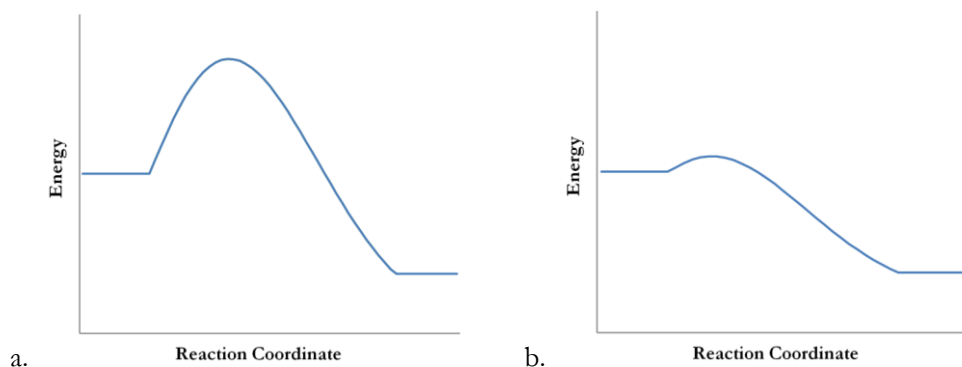
Trial	[NO]	[O <sub>2</sub> ]	Instantaneous Rate
1	0.00500 M	0.00100 M	0.000178 M/s
2	0.0100 M	0.00100 M	0.000712 M/s
3	0.0100 M	0.00200 M	0.001424 M/s

6. The reaction coordinate diagrams of two reactions are given below. Which represents the reaction with the higher rate constant?



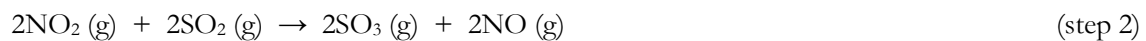


9. The two reaction coordinate diagrams below are for the same reaction. Which one represents the reaction with a catalyst added?



10. Besides catalase, there are two other well-known catalysts for the decomposition of  $\text{H}_2\text{O}_2$  (aq):  $\text{MnO}_2$  (s) and  $\text{KI}$  (aq). Which acts by changing the reaction mechanism of the decomposition?

11. For the following reaction mechanism, identify the catalyst and the overall reaction:



**Chapter 15 Review Questions**

1. Define the following terms:

- a. Kinetics
- b. Catalyst
- c. Heterogeneous catalyst
- d. Homogeneous catalyst

2. What are the units for reaction rate?

3. The reaction rate for a particular reaction can be defined in two different ways:

$$\text{Rate} = \frac{\Delta[\text{SO}_2]}{\Delta t} \quad \text{Rate} = \frac{-\Delta[\text{SO}_3]}{\Delta t}$$

Which is the reactant:  $\text{SO}_3$  or  $\text{SO}_2$ ?



6. The rate of a reaction is measured at 50.0 °C. If the experiment is later done at 25 °C with the same reactants at the same concentration, will its rate be faster or slower than the original experiment?
  
  
  
  
  
  
  
  
  
  
7. A reaction is allowed to proceed for a while, and then its rate is measured. More of one reactant is put into the vessel, increasing the concentration of that reactant. The rate is measured again right away. How will the two rates compare?
  
  
  
  
  
  
  
  
  
  
8. Fill in the blank: In order for reactants to produce products, they must participate in an effective \_\_\_\_\_.
  
  
  
  
  
  
  
  
  
  
9. How does collision theory explain why reaction rate increases with increasing temperature and increasing reactant concentration?
  
  
  
  
  
  
  
  
  
  
10. The rate equation for a reaction is determined to be  $\text{Rate} = k[\text{A}]^2[\text{B}]$ .
  - a. What is the order of the reaction with respect to reactant A?
  
  
  
  
  
  
  
  
  
  
  - b. What is the order with respect to reactant B?
  
  
  
  
  
  
  
  
  
  
  - c. What is the overall order of the reaction?
  
  
  
  
  
  
  
  
  
  
  - d. What is the unit of the rate constant?

11. The rate constant for a reaction is  $0.0519 \frac{1}{\text{M}^3 \cdot \text{s}}$ . What is the overall order of the reaction?
12. The rate equation for the reaction  $aA + bB \rightarrow cC + dD$  is  $R = k[B]^2$ . What is the order of the reaction with respect to A? If you double the concentration of A, what happens to the reaction rate?



13. If a reaction is left to run on its own, what happens to the rate of the reaction as time goes on?
14. A chemist is doing a rate study on the following reaction:

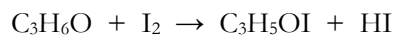


The following data are obtained:

Trial	[ClO <sub>2</sub> ]	[OH <sup>-</sup> ]	Instantaneous Rate
1	0.060 M	0.030 M	0.0248 M/s
2	0.020 M	0.030 M	0.00276 M/s
3	0.020 M	0.090 M	0.00828 M/s

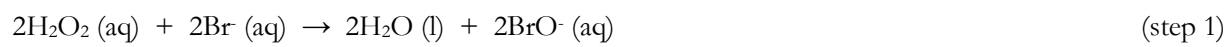
What is the rate equation for the reaction, including the rate constant?

15. Use the data in the table below to determine the rate equation (including the rate constant) for the following reaction:



<b>Trial</b>	<b>[C<sub>3</sub>H<sub>6</sub>O]</b>	<b>[I<sub>2</sub>]</b>	<b>Instantaneous Rate</b>
<b>1</b>	0.0010 M	0.0020 M	376 M/s
<b>2</b>	0.0010 M	0.0040 M	376 M/s
<b>3</b>	0.0020 M	0.0040 M	752 M/s

16. At 25 °C, the rate constant of reaction A is 824 1/M·s. The rate constant of reaction B is 4,260 1/M<sup>2</sup>·s at the same temperature. Finally, the rate constant of reaction C is 15 1/M·s. Which reaction has the highest activation energy? Which has the lowest?
17. Consider reaction A discussed in the previous problem. A series of experiments at a different temperature indicates that the rate constant is 0.501 1/M·s. Were the experiments run at a temperature above or below 25 °C?
18. There are two different catalysts that can be used for a particular reaction. The first catalyst increases the rate by a factor of 10, while the second increases the rate by a factor of 100. If the reaction is exothermic, draw reaction coordinate diagrams that represent the original reaction, the reaction with the first catalyst, and the reaction with the second catalyst.
19. What is the name of the chemical in yeast that catalyzes the decomposition of hydrogen peroxide? Where is it found besides in yeast?
20. Solid platinum can catalyze the decomposition of ethanol vapors. Does it act as a heterogeneous or homogeneous catalyst?
21. Given the following reaction mechanism:



What is the catalyst and the overall reaction?

22. A catalyst speeds up a reaction without changing the mechanism of the reaction. Is it a homogeneous or heterogeneous catalyst?

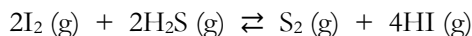
### Chapter 16 Comprehension Check Questions

1. A chemist is studying a reaction, measuring the rates of the forward reaction and the reverse reaction. Here are the data collected:

Time	Forward rate	Reverse rate		Time	Forward rate	Reverse rate
5 seconds	145 M/s	2 M/s		65 seconds	69 M/s	69 M/s
25 seconds	101 M/s	19 M/s		85 seconds	69 M/s	69 M/s
45 seconds	73 M/s	44 M/s		95 seconds	69 M/s	69 M/s

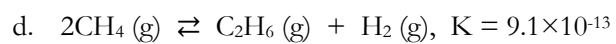
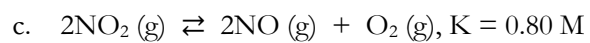
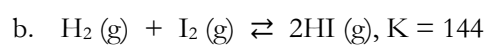
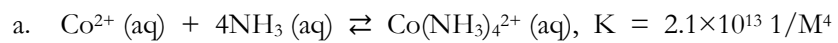
At what time did the concentration of reactants and products stop changing?

2. A chemist is studying the following reaction:

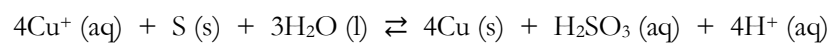


Initially, the concentration of  $\text{S}_2$  is 0.100, the concentration of HI is 0.200 M, and the concentrations of both  $\text{I}_2$  and  $\text{H}_2\text{S}$  are zero. At equilibrium, the concentration of  $\text{S}_2$  is .055 M, the concentration of HI is .065 M, and the concentrations of  $\text{I}_2$  and  $\text{H}_2\text{S}$  are both .045 M. What is the equilibrium constant?

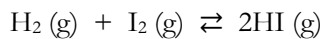
3. For the following reactions, write them as one-way reactions where possible. If they must be written as equilibrium reactions, indicate which side of the equation they are weighted toward.



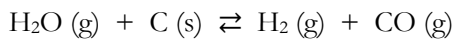
4. Write the equation for the equilibrium constant of the following reaction:



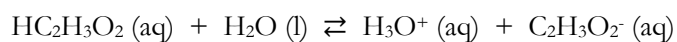
5. The following equilibrium has an equilibrium constant of 144. If  $[HI] = 0.540 \text{ M}$ ,  $[H_2] = .0460 \text{ M}$ , and  $[I_2] = 0.0440 \text{ M}$ , is the reaction at equilibrium? If not, which way will it shift to get to equilibrium?



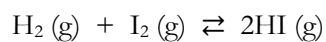
6. The reaction below has an equilibrium constant of 0.242 M at 900 K. At a specific time, there are 500.0 grams of carbon, the concentration of water vapor is 0.045 M, the concentration of hydrogen gas is 0.125 M, and the concentration of carbon monoxide is also 0.125 M. Is the reaction at equilibrium? If not, which way will it shift?



7. The following equilibrium is established:



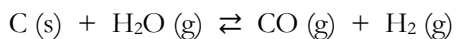
- a. What happens to the concentration of  $\text{HC}_2\text{H}_3\text{O}_2$  if more  $\text{H}_3\text{O}^+$  is added?
  - b. What happens to the concentration of  $\text{HC}_2\text{H}_3\text{O}_2$  if  $\text{H}_3\text{O}^+$  is removed?
  - c. What happens to the concentration of  $\text{HC}_2\text{H}_3\text{O}_2$  if more  $\text{H}_2\text{O}$  is added?
8. A sample of 1 mole of hydrogen gas and 2 moles of iodine gas are added to a reaction vessel. The following equilibrium is established:



A solid is added that reacts with  $\text{HI} (\text{g})$ , removing it from the reaction vessel. Assuming there is plenty of solid, how much hydrogen and iodine will be in the vessel after a long time has passed?



9. The following equilibrium has a  $\Delta H$  of 131.3 kJ.



What happens to the concentration of each gas in the reaction when the temperature is decreased?

10. What will happen to the concentration of  $\text{O}_2$  (g) when the pressure is increased on the following reaction?



11. Pyridine ( $\text{C}_5\text{H}_5\text{N}$ ) is a weak base.

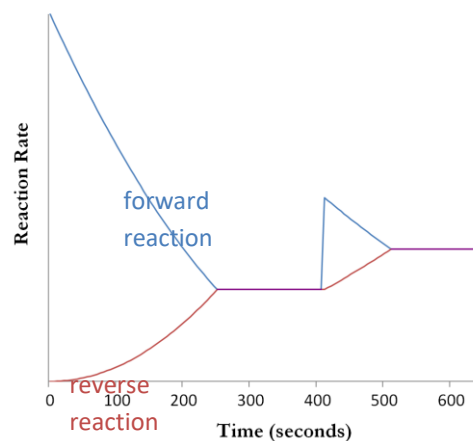
- Give the expression for its base ionization constant.
- The beaker labelled “pH 9” in the picture on page 504 contains a 0.1 M solution of pyridine. If ammonia’s base ionization constant is  $1.8 \times 10^{-5}$  M, which of the following is a possible base ionization constant for pyridine? 5.6M,  $1.4 \times 10^{-9}$  M, or 0.00094 M

### Chapter 16 Review Questions

1. Define the following terms:

- a. Chemical equilibrium
  - b. Le Chatelier's Principle
  - c. Acid ionization constant
  - d. Base ionization constant
2. When you first mix the reactants in an equilibrium reaction, which reaction (forward or reverse) is the faster one? As time goes on, which reaction slows down, and which reaction speeds up?
3. A student measures the concentration of products in an equilibrium reaction and sees that they are changing. Over time, however, the concentration levels off and stops changing. The student decides that the reaction must have stopped. Is the student correct? Why or why not?

4. The rates of the forward reaction (blue) and reverse reaction (red) for an equilibrium reaction are graphed on the right. At what time did the reaction first reach equilibrium? At what time was the equilibrium disturbed? At what time did the reaction reach equilibrium after being disturbed?

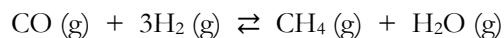


5. The equilibrium constant for a particular reaction is given by:

$$K = \frac{[\text{NH}_3]_{\text{eq}}^4 [\text{O}_2]_{\text{eq}}^5}{[\text{NO}]_{\text{eq}}^4 [\text{H}_2\text{O}]_{\text{eq}}^6}$$

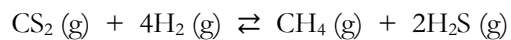
What is the balanced chemical equation for this equilibrium?

6. A chemist starts with CO at a concentration of 0.884 M and H<sub>2</sub> at a concentration of 0.445 M to start the following reaction:

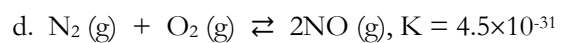
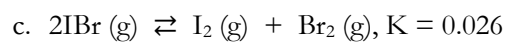
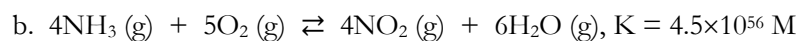
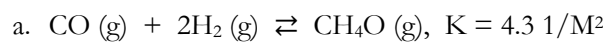


At 930 °C, the equilibrium concentrations are found to be as follows: [CO] = 0.784 M, [H<sub>2</sub>] = 0.145 M, [CH<sub>4</sub>] = 0.100 M, and [H<sub>2</sub>O] = 0.100 M. What is the equilibrium constant?

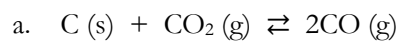
7. The equilibrium constant for the following reaction at 900 °C is 27.8 1/M<sup>2</sup>. At equilibrium, [H<sub>2</sub>S] = 0.045 M, [CH<sub>4</sub>] = 0.0610 M, and [H<sub>2</sub>] = 0.115 M. What is the concentration of CS<sub>2</sub> at equilibrium?



8. Given the information below, write each reaction as a one-way reaction if possible. If not, indicate which side the equilibrium is weighted toward:



9. Write the equation for the equilibrium constants of the following reactions:

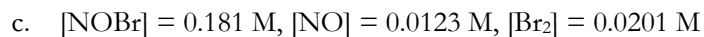
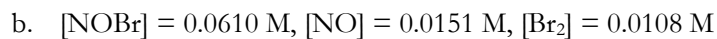
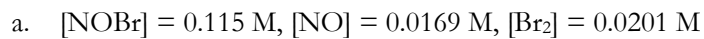




10. A chemist is studying the following chemical reaction at 25 °C:

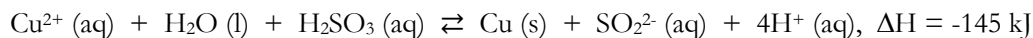


For each of the following situations, determine if the reaction is at equilibrium. If it isn't, indicate which way it must shift to reach equilibrium.



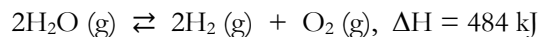
11. Fill in the blanks: Salt melts ice by disturbing the \_\_\_\_\_ that exists between the solid and \_\_\_\_\_ phases of water.

12. The following reaction reaches equilibrium:



- What will happen to the concentration of  $\text{SO}_2^{2-}$  if the temperature is increased?
- What will happen to the concentration of  $\text{SO}_2^{2-}$  if  $\text{H}_2\text{SO}_3$  is added?
- What will happen to the concentration of  $\text{SO}_2^{2-}$  if  $\text{H}^+$  is removed?
- What will happen to the concentration of  $\text{SO}_2^{2-}$  if Cu is added?

13. The following reaction reaches equilibrium:



- What will happen to the concentration of  $\text{H}_2$  if the pressure is increased?
- What will happen to the concentration of  $\text{H}_2\text{O}$  if the pressure is decreased?
- What will happen to the concentration of  $\text{O}_2$  if the temperature is decreased?

14. Three beakers contain three different acids, all at the same concentration. The pH of the solution in beaker A is 1.4, the pH of the solution in beaker B is 5.1, and the pH of the solution in beaker C is 3.2. The acid ionization constants of the three acids are  $9.4 \times 10^{-11}$  M,  $3.2 \times 10^{-2}$  M, and  $4.5 \times 10^{-6}$  M. Match the ionization constant to the beaker that contains the corresponding acid.

15. Two bases at the same concentration have different pH's. One is 14, and the other is 10. Which of them is a weak base?
16. What is the equation for the acid ionization constant of HCN?
17. What is the equation for the base ionization constant of  $\text{CH}_5\text{N}$ ?
18. The  $K_b$  for  $\text{NH}_3$  is  $1.8 \times 10^{-5}$  M. The  $K_b$  for  $\text{C}_2\text{H}_7\text{N}$  is  $5.1 \times 10^{-4}$  M. At the same concentration, which produces a solution with the higher pH?



## *Documenting Experiments*

The problem with learning things is that you can forget them pretty easily. As a result, it is important to document what you learn. That way, you can always go back later and review the material. When it comes to documenting what you learned by taking this course, you answer questions and take tests. How do you document what you learned in an experiment? By recording important aspects of the experiment in your laboratory notebook.

Start your record by writing the number and title of the experiment. Then write "Data:" underneath the title. That's how you should start to document every lab you do. You should write the experiment number and its title. Underneath, you should write "Data:" to indicate that what follows will be all the data you collect from the experiment.

What are the data you collect in an experiment? Those are your *observations*. Every experiment has data, because every experiment requires you to make observations as the experiment progresses. So under "Data:" you should list every observation you make. Each new observation should be written underneath the previously-made observation. Some experiments require more than just observations. Some require measurements. If the experiment instructions tell you to measure how long something is, that measurement is also considered data and should be written down along with your observations.

Each piece of data needs to have a short explanation regarding when you collected it in the experiment. That allows you to remember what you did right before you made the observation. It doesn't need to be a long explanation. It just needs to be a short note that will help remind you of what was done right before the observation was made.

The data section of each lab report, then, contains the quick notes you make while you are doing the experiment. They help remind you what you saw at each important step in the experiment. You write these things down while you are doing the lab so that they are fresh in your mind.

But that's not all you need to do to document your lab. So far, I have discussed things you write down *while you are doing the lab*. Once the lab is over, you need to finish documenting it. How do you do that? You add another section to your lab report that is labeled "Summary:". In that section of the lab, you write your own summary of what you did. It should not be a step-by-step listing of the instructions, and it *cannot* be a copy of the lab instructions that are in the book. Instead, it needs to be your "story" about what you did in the experiment.

Once you have finished your summary, there is one more section you need to add. Label this section “Conclusions:”, and it should contain a discussion of what you were supposed to learn by doing the experiment. This is actually easy, because I always explain that after the experiment. So all you have to do is give that same explanation, but in your own words. Once you’ve done that, you are finished documenting your lab.

While this might seem like a lot of work, it’s important for three reasons. First, it gives you something you can review later so that you can remember what you learned. Second, when you write something out in your own words, you think through it. As a result, you learn it better. Finally, there are times where you have to actually show evidence that you did experiments. Most universities, for example, require that students do experiments as a part of their high-school science courses. If you apply to a university, the people who decide whether or not you can come to the university might ask you to demonstrate that you did experiments. A lab notebook is exactly what they are looking for.

Now remember, the main goal for doing this is so that you can go back and review it later to recall what you did, what you saw, and what you learned. However, it is also possible that you will need to use this report to give evidence that you did laboratory work in your science course. Since that’s one of the goals, you need to write your report so that someone who has never seen the book can understand what you did and what you learned. Obviously, just reading the data section will be confusing to someone who doesn’t have access to the instructions, but that’s why you add a summary after the data. It helps someone who has not read the book to understand what you did, what you saw, and what you learned.

Now please understand that there is no standard among science courses regarding how you should document your labs. Some high school science courses require you to write at least a few of your experiments the way you would write about them in a scientific journal. This is usually called a “formal laboratory report.”

I don’t think students in middle school and high school should do those kinds of reports. First, most students who take science will never actually write such a report in real life. As a result, it seems like a waste of time for most students. More importantly, the way you write a formal lab report changes depending on the kind of science you are doing. As a nuclear chemist, for example, the papers that I have published in the scientific literature follow a completely different format than the papers my wife (a biophysicist) has published in the scientific literature. In my opinion, you should determine what kind of science you will be doing before you start worrying about writing a formal laboratory report.

Other science courses want you to follow the scientific method when you document your lab work. They want you to start your report with a hypothesis and end your report with a conclusion about whether or not your hypothesis was confirmed. I don’t see that as reasonable for most situations involving students.

After all, that's not what you are doing. You aren't making or testing a hypothesis. You are simply following my instructions. Also, when you make a hypothesis, you should design your experiment to address the hypothesis. Making a hypothesis for an already-designed experiment is backwards when it comes to the scientific method.

If you follow my method for documenting your labs, you will practice the most important aspect of laboratory work: making a record of what you did, what you saw, and what you learned. No matter what kind of science you end up doing, you will have to do that. Thus, by getting experience documenting labs in this way, you will be honing a skill that you will use if you pursue any kind of science.

This is important, since a scientist's laboratory notebook can become a legal document. If you discover something new and need to demonstrate that you were the one who discovered it, you can do that with your laboratory notebook. In addition, if someone disputes what you have concluded based on your experiments, your laboratory notebook can be used to resolve that dispute. In the end, then, getting used to properly documenting your experiments is an important part of science education.



*Laboratory Notebook*































































































































































