Basic Skills in Chemistry

by
Lou J. Ambelon

J. Weston Walch, Publisher

A JWW REPRODUCIBLE BOOK

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DEDICATION

For Jeff Williams,
whose guidance, counsel,
and friendship have been
a source of strength for
many years.

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Introduction

How can we know if students understand ideas in chemistry? Few questions in the pedagogy of chemical education are as fundamental as this one. We want students to develop a command of basic concepts in chemistry. But how can we check to be sure that these ideas have been "caught" by students, and not just "taught" by instructors?

One answer to that question is **problem solving**. For many principles in chemistry, the test of understanding is the ability to solve problems related to that principle. For example, can students write the electronic configuration for various atoms? Can they calculate the amount of solute in a given solution? Can they determine the freezing point of a certain solution? Can they properly represent certain nuclear changes? If the answer is "yes" to these questions, we may have some level of confidence that students have mastered the related concepts.

But the ability to solve problems—and the ability to understand chemical concepts—does not come easily. Only through continual practice with many problems do most of us develop skills in formula writing, equation balancing, and other types of problem solving.

The purpose of this publication, then, is to provide students with a large number of opportunities to solve problems and develop basic skills in chemistry.

This reproducible book is divided into twenty-eight sections, each dealing with a fundamental skill in chemistry. Each section is introduced with a brief one-page description of the skill. At least two solved examples illustrate the application of the skill. In such a short space, no detailed explanation is possible. The introductory sheet should be regarded as a review or a "back-up" for your own explanation or that of your standard classroom text.

Also, please note that the sequence of exercises in this book may not follow that of your standard classroom text or your own course. Therefore, you may have to use the practice exercises in a somewhat different order from that shown here.

Following each introductory sheet are three pages of practice problems. The problems are divided into three levels: simple (Level 1), moderate (Level 2), and advanced (Level 3). In each case, the exact meaning of these terms is provided in the

Teacher Notes that accompany each topic. Answers for all problems at all levels are also provided in the Teacher Notes.

Basic Skills in Chemistry contains about 1000 problems. As carefully as we have checked, some errors are certainly present. If you locate errors or have suggestions about this publication, we would be delighted to hear from you. Please contact the author in care of the publisher.

The author wishes to express his very special appreciation to Constance C. Holden, Bangor Community College, Bangor, Maine, for her very careful review of the manuscript and her many useful suggestions for corrections and improvements. Any errors which remain are completely the responsibility of the author.

—Lou J. Ambelon

The Metric System

All measurements in scientific work are recorded in the metric system. The metric system consists of certain agreed-upon names for the basic units of measure (**liter**, **meter**, **gram**, **degree**, etc.) and certain prefixes which designate various powers of ten (**milli-**, **deci-**, **centi-**, **kilo-**, etc.). The most recent set of rules for the use of the metric system is called the Systême Internationale (SI). Appendix 1 (page 189) lists the most common units and prefixes encountered in the SI system.

Solved Examples

Example 1: Convert a measurement of 32.5 mm to its equivalent value in cm.

Solution: Conversions of this kind can be made in many ways. One of the most convenient is by means of dimensional analysis. See section 2 if you need further information on this system.

Step 1: Select the unit factor needed to convert from the given unit of measurement to the desired unit.

$$32.5\text{ mm} \times (?\text{ unit factor}) = \_\_\_\_\_\text{ cm}$$

The unit factor needed here is $\dfrac{1\text{ cm}}{10\text{ mm}}$.

Notice that when this unit factor is used, the labels "mm" divide out, and the proper label, "cm," is obtained.

$$32.5\,\cancel{\text{mm}} \times \frac{1\text{ cm}}{10\,\cancel{\text{mm}}} = \_\_\_\_\_\text{ cm}$$

Step 2: Multiply and divide in order to obtain the correct numerical answer.

$$32.5\text{ mm} \times \frac{1\text{ cm}}{10\text{ mm}} = 3.25\text{ cm}$$

Example 2: Convert a measurement of 438.3 mm to km.

Solution: Sometimes two or more unit factors may be required to make the necessary conversion.

Step 1:

$$438.3\ \cancel{\text{mm}} \times \frac{1\ \cancel{\text{m}}}{1000\ \cancel{\text{mm}}} \times \frac{1\text{ km}}{1000\ \cancel{\text{m}}} = \underline{\qquad\qquad}\text{ km}$$

Step 2:

$$438.3\ \cancel{\text{mm}} \times \frac{1\ \cancel{\text{m}}}{1000\ \cancel{\text{mm}}} \times \frac{1\text{ km}}{1000\ \cancel{\text{m}}} = 0.0004383\text{ km}$$

NAME______________________________ DATE____________________

The Metric System

Practice Problems (Level 1)

1. Convert a measurement of 100 cm to its equivalent in m.
2. Convert a measurement of 0.001 m to its equivalent in mm.
3. Convert a measurement of 10 L to its equivalent in mL.
4. Convert a measurement of 1000 g to its equivalent in kg.
5. Convert a measurement of 0.001 s to its equivalent in μs.
6. Convert a measurement of 18 cm to its equivalent in m.
7. Convert a measurement of 75 mL to its equivalent in L.
8. Convert a measurement of 342 mg to its equivalent in g.
9. Convert a measurement of 0.055 s to its equivalent in μs.
10. Convert a measurement of 32.9 g to its equivalent in cg.
11. Convert a measurement of 553.2 cm to its equivalent in m.
12. Convert a measurement of 396.77 μs to its equivalent in s.
13. Convert a measurement of 0.049 L to its equivalent in mL.
14. Convert a measurement of 8.026 m to its equivalent in μm.
15. Convert a measurement of 0.00805 kg to its equivalent in g.

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The Metric System

Practice Problems (Level 2)

1. Convert a measurement of 50 cm to its equivalent in mm.
2. Convert a measurement of 25 mg to its equivalent in cg.
3. Convert a measurement of 36 mL to its equivalent in dL.
4. Convert a measurement of 89 cg to its equivalent in mg.
5. Convert a measurement of 35.8 hg to its equivalent in kg.
6. Convert a measurement of 22.9 mL to its equivalent in μL.
7. Convert a measurement of 60.3 dg to its equivalent in cg.
8. Convert a measurement of 0.866 km to its equivalent in hm.
9. Convert a measurement of 82.5 cL to its equivalent in mL.
10. Convert a measurement of 0.8209 cg to its equivalent in mg.
11. Convert a measurement of 264.77 nm to its equivalent in cm.
12. Convert a measurement of 0.0489 μL to its equivalent in mL.
13. Convert a measurement of 1.1194 dm to its equivalent in mm.
14. Convert a measurement of 0.00804 cm to its equivalent in mm.
15. Convert a measurement of 37.84 mg to its equivalent in cg.
16. Convert a measurement of 1.0×10^{-4} cg to its equivalent in mg.
17. Convert a measurement of 5.00×10^{3} mm to its equivalent in cm.
18. Convert a measurement of 2.35×10^{2} km to its equivalent in hm.
19. Convert a measurement of 0.588×10^{3} mg to its equivalent in dg.
20. Convert a measurement of 7.755×10^{3} μg to its equivalent in mg.

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The Metric System

Practice Problems (Level 3)

1. Convert a measurement of 0.0030 km to its equivalent in cm.
2. Convert a measurement of 350.0 mg to its equivalent in kg.
3. Convert a measurement of 0.0895 kL to its equivalent in cL.
4. Convert a measurement of 266.7 pg to its equivalent in kg.
5. Convert a measurement of 4.82 hm to its equivalent in dm.
6. Convert a measurement of 5.002 hL to its equivalent in mL.
7. Convert a measurement of 44.82 Å to its equivalent in cm.
8. Convert a measurement of 9.802 mL to its equivalent in nL.
9. Convert a measurement of 35.87 dag to its equivalent in cg.
10. Convert a measurement of 0.00883 km to its equivalent in cm.
11. Convert a measurement of 0.0000202 cm to its equivalent in μm.
12. Convert a measurement of 355.8 hL to its equivalent in cL.
13. Convert a measurement of 0.0022882 kg to its equivalent in cg.
14. Convert a measurement of 0.620 dag to its equivalent in ng.
15. Convert a measurement of 1.357 mm to its equivalent in km.
16. Convert a measurement of 3.88×10^{4} mg to its equivalent in kg.
17. Convert a measurement of 1.995×10^{-6} km to its equivalent in cm.
18. Convert a measurement of 4.5×10^{12} pg to its equivalent in cg.
19. Convert a measurement of 3.44×10^{-16} kL to its equivalent in μL.
20. Convert a measurement of 7.095×10^{7} cg to its equivalent in kg.

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The Metric System

Teacher Notes

1. Explanation of Levels

Level 1: Conversions to and from basic unit only.

Level 2: Conversions between two units, both of which are less than **or** greater than the basic unit; simple problems involving exponential notation.

Level 3: Conversions of all kinds, from less than basic unit to greater than basic unit, and vice versa; more extensive use of exponential notation problems.

2. Answers

Level 1: 1. 1m; 2. 1mm; 3. 10,000 mL; 4. 1 kg; 5. 1000 μs; 6. 0.18 m; 7. 0.075 mL; 8. 0.342 g; 9. 55,000 μs; 10. 3290 cg; 11. 5.532 m; 12. 0.00039677 s; 13. 49 mL; 14. ~~0.~~8,026,000. μm; 15. 8.05 g

Level 2: 1. 500 mm; 2. 2.5 cg; 3. 0.36 dL; 4. 890 mg; 5. 3.58 kg; 6. 22,900 μL; 7. 603 cg; 8. 8.66 hm; 9. 825 mL; 10. 8.209 mg; 11. 2.6477×10^{-5} cm; 12. 0.0000489 mL; 13. 111.94 mm; 14. 000.0804 mm; 15. 3.784 cg; 16. 1.0×10^{-3} mg; 17. 5.00×10^{2} cm; 18. 2.35×10^{3} hm; 19. 5.88 dg; 20. 7.755 mg

Level 3: 1. 300 cm; 2. 0.000350 kg; 3. 8950 cL; 4. 2.667×10^{-13} kg; 5. 4820 dm; 6. 500,200 mL; 7. 44.82×10^{-8} cm; 8. 9.802×10^{6} nL; 9. 35,870 cg; 10. 883 cm; 11. 0.202 μm; 12. 3,558,000 cL; 13. 228.82 cg; 14. 6.2×10^{9} ng; 15. 1.357×10^{-6} km; 16. 3.88×10^{-2} kg; 17. 1.995×10^{-1} cm; 18. 450 cg; 19. 3.44×10^{-7} μL; 20. 709.5 kg

Dimensional Analysis
(Factor-Label Method)

Most calculations in chemistry involve measured quantities. In such calculations, the units in which quantities are measured must be treated mathematically just as the numerical parts of the quantities are. For example, in multiplying 1.2 cm by 2.0 cm, there are two separate calculations to be carried out. First, it is necessary to multiply the two numbers: $1.2 \times 2.0 = 2.4$. Second, it is necessary to multiply the two units: $cm \times cm = cm^2$. The complete answer then, is $1.2\ cm \times 2.0\ cm = 2.4\ cm^2$.

This concept can be applied in the solution of many problems. The application depends on the use of a "unit factor." A **unit factor** is a fraction in which the numerator and denominator both represent the same measurement. For example, the fraction $\frac{100\ cm}{1\ m}$ is a unit factor since both numerator and denominator represent the same length (one meter). The solved examples illustrate the use of unit factors in solving problems by dimensional analysis.

Solved Examples

Example 1: Convert 45.3 cm to its equivalent measurement in mm.

Solution: Select a unit factor which will convert the unit "cm" to the unit "mm." The appropriate unit factor is: $\frac{10\ mm}{1\ cm}$. Arrange the problem so that the given measurement, when multiplied by the correct unit factor, will yield an answer with the proper label:

$$45.3\ \cancel{cm} \times \frac{10\ mm}{1\ \cancel{cm}} = 453\ mm$$

Example 2: Change a speed of 72.4 miles per hour to its equivalent in feet per second.

Solution: In this example, two unit factors are needed: one to change miles to feet and the other to change hours to seconds.

$$72.4\ \frac{\cancel{mi}}{\cancel{hr}} \times \frac{5280\ \text{ft}}{1\ \cancel{mi}} \times \frac{1\ \cancel{hr}}{3600\ \text{s}} = 106\ \frac{\text{ft}}{\text{s}}$$

Example 3: The density of mercury is 13.6 g/mL. What is the mass of 43 mL of mercury?

Solution: Set up the problem so that the calculation will yield a result with the unit of mass (grams).

$$13.6\ \frac{\text{g}}{\cancel{mL}} \times 43\ \cancel{mL} = 580\ \text{g}$$

NAME____________________________ DATE________________

Dimensional Analysis
(Factor-Label Method)
Practice Problems (Level 1)

Use dimensional analysis in solving each of the following problems.

1. Convert 14 mm to its equivalent measurement in m.
2. Convert 35 kg to its equivalent measurement in g.
3. Convert 57 mL to its equivalent measurement in L.
4. Convert a speed of 88 m/s to its equivalent in cm/s.
5. Convert a density of 9.45 g/L to its equivalent in g/mL.
6. The density of mercury metal is 13.6 g/mL. What is the mass of 3.55 mL of the metal?
7. The density of lead is 11.3 g/mL. What is the mass of 45 mL of the metal?
8. The density of salt (sodium chloride) is 2.16 g/mL. What is the mass of 100.0 mL of this solid?
9. A particle moves through a gas at a speed of 15 km/s. How far will it move in 5.5 s?
10. A mole of copper contains 6.02×10^{23} atoms. How many atoms are there in 0.525 moles?
11. A solution of barium nitrate contains 61.2 g per liter of solution. How many grams of barium nitrate is contained in 2.75 L of this solution?
12. A sample of seawater contains 0.00245 g of sodium chloride per mL of solution. How much sodium chloride is contained in 50.0 mL of this solution?

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Dimensional Analysis
(Factor-Label Method)
Practice Problems (Level 2)

Use dimensional analysis in solving each of the following problems.

1. Convert 15.9 mm to its equivalent measurement in km.
2. Convert 0.0982 hg to its equivalent measurement in cg.
3. Convert 13,455 g to its equivalent measurement in kg.
4. Convert a speed of 73.5 km/hr to its equivalent in m/s.
5. Convert a density of 4.52 g/mL to its equivalent in kg/L.
6. The density of iron is 7.86 g/mL. What volume of iron will have a mass of 50.00 g?
7. The density of helium gas is 0.178 g/L. What would be the mass of 375.0 mL of this gas?
8. A particle moving through a gas at a speed of 45.8 m/s strikes one wall of the container, bounces off and hits the other wall 25.0 cm away. How long did it take to go from one wall to the other?
9. A mole of sodium atoms contains 6.02×10^{23} atoms. How many moles would be needed in order to have 25.0×10^{23} atoms?
10. A mole of hydrogen atoms contains 6.02×10^{23} atoms. A section of outer space contains 25 atoms. How many moles of hydrogen is this?
11. The speed of light is 3.0×10^{10} cm/s. Express this speed in km/hr.
12. A sample of seawater contains 6.277 g of sodium chloride per liter of solution. How many mg of sodium chloride would be contained in 15.0 mL of this solution?

NAME________________________DATE______________

Dimensional Analysis
(Factor-Label Method)
Practice Problems (Level 3)

Use dimensional analysis in solving each of the following problems.

1. Convert 32.5 oz to its equivalent measurement in cg.
2. Convert 3.55 yd to its equivalent measurement in cm.
3. Convert 143.55 mL to its equivalent in pints.
4. Convert a speed of 35.8 mi/hr to its equivalent in m/s.
5. Convert a density of 13.6 g/mL to its equivalent in lb/ft^3.
6. A mole of hydrogen atoms contains 6.02×10^{23} atoms and occupies 22.4 L. How many hydrogen atoms are contained in 25.00 mL of this gas?
7. What volume of hydrogen would contain 4.5×10^{18} hydrogen atoms? How many moles of hydrogen would this be?
8. A molecule of hydrogen moves at a speed of 115 cm/s. How long will it take to travel the length of a football field (100 yd long)?
9. The speed of light is 3.0×10^{10} cm/s. Express this speed in mi/hr.
10. A sample of seawater contains 0.075 g of sodium chloride per mL of solution. How many moles of sodium chloride are there per L of this solution? A mole of sodium chloride is equivalent to 58.5 g of sodium chloride.
11. A doctor orders that a patient receive 1.5×10^{-3} mole of sodium chloride. The only solution available contains 1.00 g per 100 mL of solution. How much of this solution should the nurse give the patient?
12. A sample of air contains 2.33×10^{-4} mg of lead per mL of gas. This air passes through an office, the volume of which is 3.25×10^4 L. Seven people normally work in this office. How many μg of lead will each person in the office receive from this sample of air?

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Dimensional Analysis
(Factor-Label Method)
Teacher Notes

1. Explanation of Levels

Level 1: Problems involve the use of a single unit factor only; simple multiplication skills only are required.

Level 2: Some problems involve the use of a single unit factor only, while others require two or more unit factors.

Level 3: Problems require two or more unit factors; exponential notation is involved in some problems.

2. Answers

Level 1: 1. 0.014 m; 2. 35,000 g; 3. 0.057 kL; 4. 8800 cm/s; 5. 0.00945 g/mL; 6. 48.3 g; 7. 510 (508.5) g; 8. 216 g; 9. 83 km; 10. 3.16×10^{23} atoms; 11. 168 g; 12. 0.123 g

Level 2: 1. 1.59×10^{-5} km; 2. 982 cg; 3. 13.455 kg; 4. 20.4 m/s; 5. 4.52 kg/L; 6. 6.36 mL; 7. 0.0668 g; 8. .00546 s; 9. 4.15 mol; 10. 4.2×10^{-23} mol; 11. 1.1×10^{9} km/hr; 12. 94.2 mg

Level 3: 1. 92,100 cg; 2. 325 cm; 3. 0.30346 pt; 4. 16.0 m/s; 5. 849 lbs/ft^3; 6. 6.72×10^{20} atoms; 7. 1.7×10^{-4} L; 7.5×10^{-6} mol; 8. 79.5 sec; 9. 6.7×10^{8} mi/hr; 10. 1.3 mol/L; 11. 8.8 mL; 12. 1.08×10^{6} μg/person

Significant Figures

Calculations involving measured figures must always take into consideration the precision of those measurements. All digits which are actually part of a measurement are known as **significant figures**. Those which serve some other function, as in acting as a place-holder, are not significant. It follows that all non-zero digits recorded in a measurement are significant. Zeroes may or may not be significant, depending on their function. Zeroes which have been measured are significant. Those which act as place-holders are not. The solved examples below and on the next page illustrate the ways in which one can determine if a zero is significant or not.

Significant digits must be taken into consideration in determining the correct answer to a calculation involving measured figures. Two rules apply. For addition and subtraction, the last significant figure to the right in the answer is in the same decimal place as the least precise number being added or subtracted. For multiplication and division, the answer may contain no more significant digits than the measurement with the fewest significant figures used in the calculation.

Solved Examples

Example 1: How many significant figures are there in the number 307 cm?

Solution: Both the first and last digits (3, 7) are non-zero, so they are both significant. Since the zero lies between two significant figures, it must also be significant. The answer, therefore, is **three**.

Example 2: How many significant digits are there in 0.00450 cm?

Solution: The first three zeroes are not part of the measurement. They are place-holders here. The two non-zero digits (4, 5) are both significant. The final zero does NOT serve as a placeholder. It must be significant also. The answer: three significant figures.

Example 3: What is the sum of 4.52 g, 13.8 g, and 7.9483 g?

Solution: Adding these three numbers gives an answer of 26.2683 g. The least accurate number among the addends is 13.8 g. Therefore, the answer may have no figure to the right of the tenths place. The correct answer is **26.3 g**.

Example 4: What is the product of 48.4398 m and 1.52 m?

Solution: The numerical product of these numbers is 73.628496 m^2. The number of significant figures in the factors is six (for 48.4398 m) and three (for 1.52 m). So there may be no more than three significant figures in the answer: 73.6 m^2.

NAME________________________ DATE______________

Significant Figures

Practice Problems (Level 1)

1. How many significant figures are there in each of the following measurements?

a. 23 cm
b. 1.498 g
c. 248.3 s
d. 9.855 mL
e. 76.414 kg
f. 32.8 m
g. 107 mm
h. 0.238 km
i. 8.0335 cm
j. 0.05587 m

2. Express the answer to each of the following calculations with the correct number of significant figures.

a. 3.42 cm + 8.13 cm
b. 4.939 g + 3.822 g
c. 17.8 cm + 12.11 cm
d. 4.552 kg + 3.14 kg
e. 1.966 s + 3.4422 s
f. 3.882 g − 2.114 g
g. 4.894 cm − 2.33 cm
h. 15.6674 m − 12.838 m
i. 11.22 g − 8.8 g
j. 133 L − 6.45 L

3. Express the answer to each of the following calculations with the correct number of significant figures.

a. 1.2 cm × 1.3 cm
b. 2.1 m × 1.8 m
c. 1.45 m × 2.2 m
d. 2.5 mm × 1.33 mm
e. 4.3324 km × 1.2 km
f. 32.88 m^2 ÷ 4.388 m
g. 16.5 km^2 ÷ 1.8 km
h. 84.99 m^2 ÷ 2.63 m
i. 9.9 mm^2 ÷ 3.4484 mm
j. 3.085 cm^2 ÷ 2.77448 cm

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Significant Figures

Practice Problems (Level 2)

1. How many significant figures are there in each of the following measurements?

   a. 3.58 g
   b. 14.809 cm
   c. 107.334 km
   d. 0.0004898 mm
   e. 3000. cm
   f. 804.58 kg
   g. 0.007832 cg
   h. 130,004.5 mm
   i. 250.00 km
   j. 14.380 s

2. Express the answer to each of the following calculations with the correct number of significant figures.

   a. 82.5 cm + 13.56 cm
   b. 16.8892 km + 3.5 km
   c. 45.456 g + 3.56 g
   d. 106.22 mm + 80.0 mm
   e. 30.44 kg + 3.9422 kg
   f. 13.80 cm − 6.0741 cm
   g. 8.472 cg − 1.440 cg
   h. 30. s − 1.442 s
   i. 54.00 g − 30.2020 g
   j. 1.45050 kg − 0.00667 kg

3. Express the answer to each of the following calculations with the correct number of significant figures.

   a. 8.4 cm × 3.58 cm
   b. 1.075 m × 2.0 m
   c. 3.0899 mm × 22.4 mm
   d. 0.00457 cm × 0.18 cm
   e. 10.00 m × 84.767 m
   f. 35.068 km^3 ÷ 5.7 km
   g. 85.0869 m^2 ÷ 9.0049 m
   h. 0.00826 cm^2 ÷ 0.00033 cm
   i. 0.005600 mm^2 ÷ 0.200 mm
   j. 3.4500 cm^2 ÷ 450 cm

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Significant Figures

Practice Problems (Level 3)

1. How many significant figures are there in each of the following measurements?

a. 307 g
b. 1.40082 cm
c. 0.00058900 g
d. 0.00300900870 mm
e. 4500 km
f. $350{,}\bar{0}00$ cm
g. 180.00 s
h. 3.50×10^{3} cm
i. 1.604×10^{-4} m
j. 0.0459×10^{3} g

2. Express the answer to each of the following calculations with the correct number of significant figures.

a. 80 cm + 13.0 cm
b. 72.60 m + 0.0950 m
c. 13.89 cm + 6.8932 cm
d. 1.30×10^{-2} cm + 2.4×10^{-4} cm
e. 8.99×10^{3} m + 1.400×10^{4} m
f. 750. cm − 677.4 cm
g. 10,000 m − 940 m
h. 0.0890 cm − 0.0666 cm
i. 0.340×10^{-1} g − 1.20×10^{-2} g
j. 4.5×10^{5} km − 3.00×10^{3} km

3. Express the answer to each of the following calculations with the correct number of significant figures.

a. 3.0 cm × 4.000 cm
b. 2.005 cm × 5.0 cm
c. 400 m × 87,488 m
d. 2.3×10^{-6} m × 1.45×10^{-2} m
e. 8.70×10^{-2} mm × $40. \times 10^{-1}$ mm
f. 0.0045 mm^2 ÷ 0.90 mm
g. 120 km^2 ÷ 8.56 km
h. 0.7600 mm^3 ÷ 1.50 mm
i. 4.80×10^{5} m^2 ÷ 8.5×10^{3} m
j. 0.630×10^{-1} m^3 ÷ 0.0804×10^{2} m

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Significant Figures

Teacher Notes

1. Explanation of Levels

Level 1: A minimal amount of problems involve the use of zero as a digit. Since all non-zero digits are significant, decisions should be very simple.

Level 2: A greater number of zeroes are included in given numbers.

Level 3: The most difficult decisions to be made regarding the significance of zeroes are involved; problems involving exponential notation are introduced.

2. Answers

Level 1:

1. a. 2; b. 4; c. 4; d. 4; e. 5; f. 3; g. 3; h. 3; i. 5; j. 4
2. a. 11.55 cm; b. 8.761 g; c. 29.9 cm; d. 7.69 kg; e. 5.408 s; f. 1.768 g; g. 2.56 cm; h. 2.829 m; i. 2.4 g; j. 127 L
3. a. 1.6 cm^2; b. 3.8 m^2; c. 3.2 m^2; d. 3.3 mm^2; e. 5.2 km^2; f. 7.493 m; g. 9.2 km; h. 32.3 m; i. 2.9 mm; j. 1.112 cm

Level 2:

1. a. 3; b. 5; c. 6; d. 4; e. 4; f. 5; g. 4; h. 7; i. 5; j. 5
2. a. 96.1 cm; b. 20.4 km; c. 49.02 g; d. 186.2 mm; e. 34.38 kg; f. 7.73 cm; g. 7.032 cg; h. 29 s; i. 23.80 g; j. 1.44383 kg
3. a. 30. cm^2; b. 2.2 m^2; c. 69.2 mm^2; d. 0.00082 cm^2; e. 847.7 m^2; f. 6.2 km^2; g. 9.4490 m; h. 25 cm; i. 0.0280 mm; j. 0.0077 cm

Level 3:

1. a. 3; b. 6; c. 5; d. 9; e. 2; f. 4; g. 5; h. 3; i. 4; j. 3
2. a. 90 cm; b. 72.70 m; c. 20.78 cm; d. 1.32×10^{-2} cm; e. 2.30×10^4 m; f. 73 cm; g. 9100 m; h. 0.0224 cm; i. 2.2×10^{-2} g; j. 4.5×10^5 km;
3. a. 12 cm^2; b. 10 cm^2; c. 3×10^7 m^2; d. 3.3×10^{-8} m^2; e. 0.35 mm^2; f. 0.0050 mm; g. 14 km; h. 0.507 mm^2; i. 56 m; j. 7.84×10^{-3} m^2

Exponential Notation

Exponential notation is used to express very large or very small numbers. For example, we could express the number of molecules of hydrogen in a mole as 602,000,000,000,000,000,000,000 or as 6.02×10^{23}. The latter is clearly a simpler number with which to work.

One form of exponential notation commonly used in chemistry is **scientific notation**. In scientific notation a number is expressed as a product of a number between 1 and 10 and a power of ten. The steps by which very large or very small numbers are converted to scientific notation are these:

1. Separate the very large or very small number into two factors, one of which is the largest factor of ten contained within the number. For example, the number 50,000 would be separated into two factors, 5 and 10,000. We select 10,000 since that is the largest power of ten contained within 50,000.

2. The two factors are then written as a product, with the second factor expressed exponentially as a power of ten. In the preceding example, 50,000 would be expressed as $5 \times 10{,}000$, or as 5×10^4. Note that 10,000 is the same as 10^4.

Solved Examples

Example 1: Express 350,000 in exponential notation.

Solution: Step 1: $350{,}000 = 3.5 \times 100{,}000$
Step 2: $3.5 \times 100{,}000 = 3.5 \times 10^5$

Example 2: Express 0.000000587 in exponential notation.

Solution: Step 1: $0.000000587 = 5.87 \times 0.0000001$
Step 2: $5.87 \times 0.0000001 = 5.87 \times 10^{-7}$

Example 3: Write out the number 5.82×10^6 in long form.

Solution: In this case, just reverse the steps given above.

Step 2: $5.82 \times 10^6 = 5.82 \times 1{,}000{,}000$
Step 1: $5.82 \times 1{,}000{,}000 = 5{,}820{,}000$

Example 4: Write out the number 4.92×10^{-5} in long form.

Solution: Step 2: $4.92 \times 10^{-5} = 4.92 \times 0.00001$
Step 1: $4.92 \times 0.00001 = 0.0000492$

Example 5: Find the product of 3.2×10^4 and 6.1×10^2.

Solution: When multiplying exponential numbers, (1) multiply the first factors times each other, as usual, and (2) multiply the exponential factors by adding exponents. Thus,

$$3.2 \times 10^4 \times 6.1 \times 10^2 = 3.2 \times 6.1 \times 10^{4+2}$$
$$= 19.52 \times 10^6$$

Note that this would usually be expressed as 1.952×10^7.

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Exponential Notation

Practice Problems (Level 1)

1. Express each of the following numbers in exponential notation.

a. 10,000
b. 0.0001
c. 10,000,000,000
d. 0.000 000 000 000 001
e. 400,000
f. 500,000,000,000,000
g. 0.000,000,08
h. 0.000,000,000,000,9
i. 7,000
j. 0.004

2. Write out each of the following numbers in the long form.

a. 1×10^{3}
b. 1×10^{-5}
c. 1×10^{14}
d. 1×10^{-2}
e. 7×10^{4}
f. 5×10^{-6}
g. 9×10^{-9}
h. 3×10^{5}
i. 2×10^{-4}
j. 8×10^{8}

3. Use exponential notation to find the product of each of the following:

a. 30,000 × 7,000
b. 0.0005 × 0.003
c. 400,000 × 50,000
d. 0.000,000,06 × 0.0004
e. 0.000,007 × 80,000
f. 10,000,000 × 0.000005
g. 3,000 × 0.0004
h. 0.000,000,000,05 × 0.003
i. 0.000,000,03 × 7,000,000
j. 900,000 × 0.000,000,009

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Exponential Notation

Practice Problems (Level 2)

1. Express each of the following numbers in exponential notation.

a. 67,000
b. 890,000,000
c. 0.000,004,5
d. 0.000,000,000,000,009,8
e. 805,000
f. 0.001,000,045
g. 0.0038809083
h. 45,887,950
i. 36,000,000,000,000
j. 0.000,000,000,000,015,05

2. Express each of the following numbers in the long form.

a. 8.2×10^{3}
b. 1.5×10^{-6}
c. 1.775×10^{7}
d. 7.065×10^{-5}
e. 8.900×10^{-8}
f. 3.86×10^{6}
g. 6.892×10^{-3}
h. 9.035×10^{-10}
i. 5.425×10^{2}
j. 2.662×10^{-1}

3. Solve each of the following problems, using exponential notation only.

a. 25,000 × 16,000
b. 57,500,000 × 8,000
c. 43,500 × 9,000,000
d. 0.00454 × 0.00898
e. 0.000,000,158 × 0.000,67
f. 389,000,000 × 0.00075
g. 0.000,000,069,05 × 57,500
h. 98,800,000 ÷ 35,600,000,000
i. 0.000,686,8 ÷ 87,000
j. 257,000,000 ÷ 0.00489

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Exponential Notation

Practice Problems (Level 3)

1. Express each of the following numbers in exponential notation.

a. 75,890

b. 1

c. 0.000,189

d. 0

e. −45,000,000

f. 8.75

g. −0.000,006,895

h. 100,500,500

i. 0.000,000,609,866

j. 12,500,000,000,000,000,000

2. Express each of the following numbers in the long form.

a. 3.907×10^{-15}

b. 18.89×10^{5}

c. 577.8×10^{-6}

d. 125.5×10^{-2}

e. 0.00854×10^{-5}

f. 0.00387×10^{5}

g. 0.0000552×10^{2}

h. 555.5×10^{-8}

i. 35.882×10^{-6}

j. 477.8×10^{7}

3. Carry out each of the indicated operations, using exponential notation only.

a. $235{,}000 \times 0.00456 \times 20{,}000$

b. $137{,}500{,}000{,}000 \times 0.000{,}450 \times 0.002{,}000$

c. $0.000{,}005{,}80 \times 18{,}000{,}000{,}000 \times 38{,}000$

d. $54{,}000{,}000 \times 0.000{,}000{,}450 \times 167{,}000$

e. $809{,}000{,}000{,}000{,}000 \times 13{,}500{,}000{,}000{,}000 \times 0.000{,}000{,}000{,}067$

f. $\dfrac{75{,}000 \times 800{,}000 \times 0.000{,}000{,}045}{66{,}500 \times 85{,}000}$

g. $\dfrac{0.0045 \times 78{,}000 \times 132{,}000{,}000{,}000}{0.000{,}555 \times 0.00155 \times 82{,}500{,}000}$

h. $\dfrac{115{,}000{,}000{,}000 \times 0.00688 \times 0.009{,}000{,}000}{25{,}600{,}000 \times 75{,}000 \times 24{,}500 \times 0.00092}$

Exponential Notation

Teacher Notes

1. Explanation of Levels

Level 1: All problems involve exact powers of ten or numbers having only one non-zero digit; multiplication problems only included.

Level 2: Problems involve more than one non-zero digit; both multiplication and division problems included.

Level 3: Problems of all levels of complexity, including some unusual numbers to be expressed in exponential notation (e.g., 1); multiplication, division, and combination problems; problems involving non-standard form of exponential notation.

2. Answers

Level 1:

1. a. 10^4; b. 10^{-4}; c. 10^{10}; d. 10^{-15}; e. 4×10^5; f. 5×10^{14}; g. 8×10^{-8}; h. 9×10^{-13}; i. 7×10^3; j. 4×10^{-3}

2. a. 1000; b. 0.000,01; c. 100,000,000,000,000; d. 0.01; e. 70,000; f. 0.000,005; g. 0.000,000,009; h. 300,000; i. 0.0002; j. 800,000,000

3. a. 2.1×10^8; b. 1.5×10^{-6}; c. 2×10^{10}; d. 2.4×10^{-11}; e. 0.56; f. 50; g. 1.2; h. 1.5×10^{-13}; i. .21; j. 8.1×10^{-3}

Level 2:

1. a. 6.7×10^4; b. 8.9×10^8; c. 4.5×10^{-6}; d. 9.8×10^{-15}; e. 8.05×10^5; f. 1.000045×10^{-3}; g. 3.8809083×10^{-3}; h. 4.5887950×10^7; i. 3.6×10^{13}; j. 1.505×10^{-14}

2. a. 8200; b. 0.000,0015; c. 17,750,000; d. 0.000,07065; e. 0.000,000,0890; f. 3,860,000; g. 0.006892; h. 0.000,000,000,9035; i. 542.5; j. 0.2662

3. a. 4×10^{8}; b. 4.6×10^{11}; c. 3.915×10^{11}; d. 4.07692×10^{-5}; e. 1.0586×10^{-10}; f. $2.9175 \times 10^{+5}$; g. 3.970375×10^{-3}; h. 2.77528×10^{-3}; i. 7.89425×10^{-9}; j. 5.25562×10^{10}

Level 3:

1. a. 7.589×10^{4}; b. 10^{0}; c. 1.89×10^{-4}; d. (no representation); e. -4.5×10^{7}; f. 8.75×10^{0}; g. -6.895×10^{-6}; h. 1.005005×10^{8}; i. 6.09866×10^{-7}; j. 1.25×10^{19}

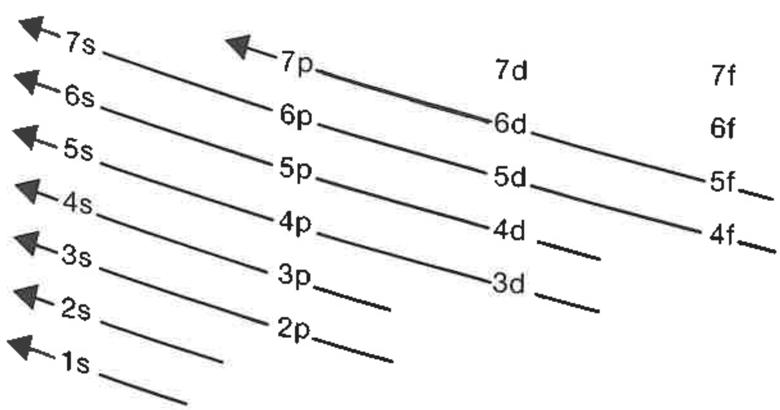
2. a. 0.000,000,000,000,003907; b. 1,889,000; c. 0.0005778; d. 1.255; e. 0.000,000,0854; f. 387; g. 0.00552; h. 0.000,005555; i. 0.000,035882; j. 4,778,000,000

3. a. 2.1432×10^{7}; b. 1.2375×10^{5}; c. 3.9672×10^{9}; d. 4.0581×10^{6}; e. 7.317405×10^{17}; f. 4.777×10^{-7}; g. 6.528×10^{11}; h. 1.6454×10^{-7}

Atomic Structure

The structure of an atom can be predicted quite accurately knowing only two characteristics of that atom: its **atomic number** and its **mass number**. The atomic number of an element is the same as the number of protons in an atom of that element. In a neutral atom, this is the same as the number of electrons in the atom. The mass number is equal to the total number of protons plus neutrons in the nucleus of the atom. It follows, then, that the number of neutrons in the atom is equal to its mass number minus its atomic number.

The electronic configuration of an atom may also be predicted with some accuracy by following the chart below. Beyond the third level, orbitals are often of such similar energies that precise predictions are not possible. For the first three dozen elements or so, however, the chart gives a close approximation to the correct electronic structure. Beyond that point, it gives good, but not precise, predictions.



Solved Examples

Example 1: Write the predicted structure for the sodium atom.

Solution: For sodium, the atomic number is 11 and the mass number 23. Then,

$$\text{number of protons} = \text{atomic number} = 11$$

$$\text{number of neutrons} = \text{mass number} - \text{atomic number}$$
$$= 23 - 11 = 12$$

Nucleus: 11 protons and 12 neutrons

Electron configuration: 11 electrons: $1s^2 2s^2 2p^6 3s^1$

Example 2: Write the predicted structure for the silicon atom.

Solution: For silicon, the atomic and mass numbers are 14 and 28, respectively.

$$\text{number of protons} = \text{atomic number} = 14$$

$$\text{number of neutrons} = \text{mass number} - \text{atomic number}$$
$$= 28 - 14 = 14$$

Nucleus: 14 protons and 14 neutrons

Electron configuration: 14 electrons: $1s^2 2s^2 2p^6 3s^4$

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Atomic Structure

Practice Problems (Level 1)

1. Write the complete atomic structure for the carbon atom. The atomic and mass numbers for carbon are 6 and 12 respectively.
2. Write the complete atomic structure for the sulfur atom. The atomic and mass numbers for sulfur are 16 and 32 respectively.
3. Write the complete atomic structure for the calcium atom. The atomic and mass numbers for calcium are 20 and 40 respectively.
4. Write the complete atomic structure for the helium atom. The atomic and mass numbers for helium are 2 and 4 respectively.
5. Write the complete atomic structure for the beryllium atom. The atomic and mass numbers for beryllium are 4 and 9 respectively.
6. Write the complete atomic structure for the aluminum atom. The atomic and mass numbers for aluminum are 13 and 27 respectively.
7. Write the complete atomic structure for the hydrogen atom. The atomic and mass numbers for hydrogen are 1 and 1 respectively.
8. Write the complete atomic structure for the argon atom. The atomic and mass numbers for argon are 18 and 40 respectively.
9. Write the complete atomic structure for the oxygen atom. The atomic and mass numbers for oxygen are 8 and 16 respectively.
10. Write the complete atomic structure for the lithium atom. The atomic and mass numbers for lithium are 3 and 7 respectively.

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Atomic Structure

Practice Problems (Level 2)

Write the predicted atomic structure for each of the atoms listed below. The symbol of the element is given immediately following its name.

1. boron (B)
2. nitrogen (N)
3. phosphorus (P)
4. neon (Ne)
5. scandium (Sc)
6. krypton (Kr)
7. strontium (Sr)
8. chromium (Cr)
9. nickel (Ni)
10. potassium (K)
11. fluorine (F)
12. iron (Fe)
13. rubidium (Rb)
14. bromine (Br)
15. antimony (Sb)
16. xenon (Xe)

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Atomic Structure

Practice Problems (Level 3)

1. Write the predicted atomic structure for each of the atoms listed below. Refer to the periodic table for information needed.

   a. magnesium
   b. cesium
   c. iodine
   d. titanium
   e. radon
   f. cadmium
   g. chlorine
   h. astatine
   i. arsenic
   j. francium
   k. selenium
   l. copper
   m. molybdenum
   n. europium
   o. thorium
   p. mercury

2. Write the predicted atomic structure for any two isotopes of the following atoms. Where necessary, refer to your answers in part 1 above for reference.

   a. magnesium
   b. iodine
   c. radon
   d. chlorine
   e. francium
   f. copper
   g. europium
   h. mercury

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Atomic Structure

Teacher Notes

1. Explanation of Levels

Level 1: Atomic and mass numbers given for all elements. Only elements of atomic numbers less than 21 are used.

Level 2: Symbols of elements, but not their atomic and mass numbers, are given. Students must refer to periodic table for atomic and mass numbers. More complex elements included.

Level 3: Names only of elements given. Students must know or find symbols and then locate atomic and mass numbers on periodic table. Ambiguous atomic weights make determination of mass numbers more difficult. Elements of all complexities used. Isotope questions included.

2. Answers

Level 1:

1. 6 +, 6 ±; $1s^2 2s^2 2p^2$
2. 16 +, 16 ±; $1s^2 2s^2 2p^6 3s^2 3p^4$
3. 20 +, 20 ±; $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
4. 2 +, 2 ±; $1s^2$
5. 4 +, 5 ±; $1s^2 2s^2$
6. 13 +, 14 ±; $1s^2 2s^2 2p^6 3s^2 3p^1$
7. 1 +, 0 ±; $1s^1$
8. 18 +, 22 ±; $1s^2 2s^2 2p^6 3s^2 3p^6$
9. 8 +, 8 ±; $1s^2 2s^2 2p^4$
10. 3 +, 4 ±; $1s^2 2s^1$

Level 2:

1. 5 +, 6 ±; $1s^2 2s^2 2p^1$
2. 7 +, 7 ±; $1s^2 2s^2 2p^3$
3. 15 +, 16 ±; $1s^2 2s^2 2p^6 3s^2 3p^3$
4. 10 +, 10 ±; $1s^2 2s^2 2p^6$
5. 21 +, 24 ±; $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$

6. 36 +, 48 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^6$
7. 38 +, 50 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^65s^2$
8. 24 +, 28 ±; $1s^22s^22p^63s^23p^63d^44s^2$
9. 28 +, 31 ±; $1s^22s^22p^63s^23p^63d^84s^2$
10. 19 +, 20 ±; $1s^22s^22p^63s^23p^64s^1$
11. 9 +, 10 ±; $1s^22s^22p^5$
12. 26 +, 30 ±; $1s^22s^22p^63s^23p^63d^64s^2$
13. 37 +, 48 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^65s^1$
14. 35 +, 45 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^5$
15. 51 +, 71 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^3$
16. 54 +, 77 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^6$

Level 3: 1. a. 12 +, 12 ±; $1s^22s^22p^63s^2$
b. 55 +, 78 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^66s^1$
c. 53 +, 74 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^5$
d. 22 +, 26 ±; $1s^22s^22p^63s^23p^63d^24s^2$
e. 86 +, 136 ±;
$1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^{14}5s^25p^65d^{10}6s^26p^6$
f. 48 +, 64 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^2$
g. 17 +, 18 ±; $1s^22s^22p^63s^23p^5$
h. 85 +, 125 ±;
$1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^{14}5s^25p^65d^{10}6s^26p^5$
i. 33 +, 42 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^3$
j. 87 +, 136 ±;
$1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^{14}5s^25p^65d^{10}6s^26p^67s^1$
k. 34 +, 45 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^4$
l. 29 +, 35 ±; $1s^22s^22p^63s^23p^63d^94s^2$
m. 42 +, 54 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^45s^2$
n. 63 +, 89 ±; $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^75s^25p^66s^2$
o. 90 +, 142 ±;
$1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^{14}5s^25p^65d^{10}6s^26p^66d^27s^2$
p. 80 +, 121 ±;
$1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^{14}5s^25p^65d^{10}6s^2$

2. Any structure which contains the same number of protons, the same electronic configuration, but a different number of neutrons, when compared to the comparable structure from a, is satisfactory.

Bonding Reactions

Atoms combine with each other by exchanging electrons in various ways. When one atom completely loses electrons and another completely takes control of them, the bond formed is called **ionic**. When two atoms share pairs of electrons with each other, more or less equally, the bond is a **covalent** one. In the vast majority of cases, the bond formed between two atoms is neither purely ionic nor purely non-polar (exactly equal sharing [covalent]). The precise nature of the bond formed between the two can be described by the difference in the ability of the two atoms to attract a pair of electrons shared between them. This ability is given the name **electronegativity**. Thus, the difference in electronegativities between two atoms is an indication of the type of bond that will exist between the two atoms.

Various methods can be used to express the bonding reactions that take place between two atoms. Perhaps the most convenient involves the use of Lewis electron dot symbols. In an electron dot symbol, the chemical symbol of the element in question represents not the complete atom, as is usually the case. Instead, it represents the kernel of the atom (the nucleus plus all inner orbitals of electrons). The electron dot symbol is completed by adding the valence electrons of the atom.

Solved Examples

Example 1: Show the bonding reaction between sodium and chlorine.

Solution: The electron dot symbols for the two elements are:

$$\text{Na}^{\bullet} \quad \text{and} \quad {}^{\times}\overset{\times\times}{\underset{\times\times}{\text{Cl}}}{}^{\times}_{\times}$$

The electronegativities of the two elements (from Appendix 4) are:

$$\text{Na} = 0.9;\ \ \text{Cl} = 3.0;\ \ \Delta E_n = 3.0 - 0.9 = 2.1$$

This is a fairly large difference and suggests that an ionic bond will form between the two atoms.

$$\mathrm{Na}^{\bullet} \quad + \quad \overset{\times\times}{\underset{\times\times}{{}^{\times}\mathrm{Cl}{}^{\times}_{\times}}} \qquad \mathrm{Na}^{+} \quad + \quad \overset{\times\times}{\underset{\times\times}{{}^{\times}_{\bullet}\mathrm{Cl}^{-}{}^{\times}_{\times}}}$$

Example 2: Show the bonding reaction to be expected between beryllium and tellurium.

Solution: The electron dot symbols for the two elements are:

$$\bullet\mathrm{Be}^{\bullet} \qquad \overset{\times\times}{\underset{\times\times}{\mathrm{Te}{}^{\times}_{\times}}}$$

The difference in electronegativities is:

$$E_n = 2.1 \text{ (for Te)} - 1.5 \text{ (for Be)} = 0.6$$

In this case, the difference is much smaller, and a (polar) covalent bond is to be expected:

$$\bullet\mathrm{Be}^{\bullet} \quad + \quad \mathrm{Te} \qquad \mathrm{Be} {}^{\bullet}_{\bullet}{}^{\times}_{\times} \overset{\times\times}{\mathrm{Te}}{}^{\times}_{\times}$$

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Bonding Reactions

Practice Problems (Level 1)

1. Write the electron dot symbol for each of the following elements.

a. aluminum
b. fluorine
c. helium
d. potassium
e. nitrogen
f. sulfur
g. magnesium
h. hydrogen

2. Use electron dot symbols to show the electron change involved in each of the following bonding reactions.

a. potassium (K) + fluorine (F) (ionic bond)

b. magnesium (Mg) + chlorine (Cl) (ionic bond)

c. hydrogen (H) + sulfur (S) (covalent bond)

d. lithium (Li) + arsenic (As) (covalent bond)

e. rubidium (Rb) + oxygen (O) (ionic bond)

f. aluminum (Al) + iodine (I) (covalent bond)

g. hydrogen (H) + phosphorus (P) (covalent bond)

h. calcium (Ca) + tellurium (Te) (covalent bond)

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Bonding Reactions

Practice Problems (Level 2)

1. Write the electron dot symbol for each of the following elements.

a. lithium
b. bromine
c. calcium
d. arsenic
e. scandium
f. boron
g. zinc
h. copper
i. antimony
j. xenon

2. For each of the following bonding reactions, predict the type of bond to be expected. Then use electron dot symbols to show the electron change that takes place in the reaction.

a. sodium + bromine
b. lithium + oxygen
c. calcium + fluorine
d. hydrogen + selenium
e. zinc + tellurium
f. beryllium + sulfur
g. lanthanum + chlorine

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Bonding Reactions

Practice Problems (Level 3)

1. Write the electron dot symbol for each of the following elements.

a. silicon

b. astatine

c. cesium

d. mercury

e. titanium

f. krypton

g. iron

h. niobium

i. platinum

j. samarium

2. For each of the following bonding reactions, predict the type of bond to be expected. Then, use electron dot symbols to show the electron change that takes place during the reaction.

a. iron + fluorine

b. chromium + iodine

c. hydrogen + antimony

d. potassium + sulfur

e. boron + nitrogen

f. magnesium + silicon

g. gold + bromine

h. scandium + oxygen

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Bonding Reactions

Teacher Notes

Explanation of Levels

Level 1: Simple electron dot symbols to write; bond types given to students.

Level 2: Somewhat more difficult electron dot symbols to write; students asked to predict bond types from electronegativity differences between reacting elements.

Level 3: Still more difficult electron dot symbols to write; students asked to predict bond types from electronegativity differences between reacting elements.

Answers

N.B.: In some of these problems, students should not be expected to predict exactly the type of bonding which occurs because, in some cases, this involves bonding with which they are familiar (e.g., metallic bonding). Instead, they should be expected to produce **reasonable** predictions for possible bonding reactions.

Level 1:

1. a. ·Al· (3 dots) b. :F: (7 dots) c. He: d. K· e. ·N· (5 dots) f. ·S· (6 dots) g. Mg: h. H·

2. a. $K^{\cdot} + \cdot F: \rightarrow K^{+} + :F:^{-}$ b. $Mg: + 2\cdot Cl: \rightarrow Mg^{2+} + 2:Cl:^{-}$

   c. $2\,H^{\cdot} + \cdot S\cdot \rightarrow H:S:H$ d. $3\,Li^{\cdot} + \cdot As\cdot \rightarrow Li:As:Li$ (with Li below As)

   e. $2\,Rb^{\cdot} + \cdot O\cdot \rightarrow 2\,Rb^{+} + :O:^{2-}$ f. $\cdot Al\cdot + 3\cdot I: \rightarrow Al:I:Al$ (with Al below I)

g. 3 H + P → H P H
H

h. Ca + Te → Ca Te or Ca Te

Level 2: 1. a. Li b. Br c. Ca d. As e. Sc f. B g. Zn

h. Cu i. Sb j. Xe

2. a. ΔE_n = 1.9; ionic; Na + Br → Na^+ + Br^-

b. ΔE_n = 2.5; ionic; 2 Li + O → 2 Li^+ + O^{2-}

c. ΔE_n = 3.0; ionic; Ca + 2F → Ca^{2+} + 2F

d. ΔE_n = 0.3; covalent; 2 H + Se → H Se H

e. ΔE_n = 0.5; covalent; Zn + Te → Zn Te or Zn Te

f. ΔE_n = 1.0; covalent; Be + S → Be Se or Be Se

g. ΔE_n = 1.9; ionic; La + 3 Cl → La^{3+} + 3 Cl

Level 3: 1. a. Si b. At c. Cs d. Hg e. Ti f. Kr g. Fe

h. Nb i. Pt j. Sm (Fe and Pt from Group VIII)

2. a. ΔE_n = 2.2; ionic; Fe + 3F → Fe^{3+} + 3F

b. ΔE_n = 0.9; covalent; Cr + 2I → I Cr I

c. ΔE_n = 0.2; covalent; 3 H + Sb → H Sb H
H

d. ΔE_n = 1.7; covalent; 2 K + S → K S K

e. ΔE_n = 1.0; covalent; B + N → B N (many possibilities)

f. $\Delta E_n = 0.6$; covalent; $2\,\mathrm{Mg}{:} + \cdot\underset{\cdot}{\dot{\mathrm{Si}}}\cdot \rightarrow \mathrm{Mg}{:}\underset{\cdot\cdot}{\ddot{\mathrm{Si}}}{:}\mathrm{Mg}$

g. $\Delta E_n = 0.4$; covalent; $\mathrm{Au}^{\cdot} + \cdot\underset{\cdot\cdot}{\ddot{\mathrm{Br}}}{:} \rightarrow \mathrm{Au}{:}\underset{\cdot\cdot}{\ddot{\mathrm{Br}}}{:}$

h. $\Delta E_n = 2.2$; ionic; $2\,\underset{\cdot}{\dot{\mathrm{Sc}}}\cdot + 3\cdot\underset{\cdot\cdot}{\ddot{\mathrm{O}}}\cdot \rightarrow 2\,\mathrm{Sc}^{3+} + 3{:}\underset{\cdot\cdot}{\ddot{\mathrm{O}}}{:}^{2-}$

Writing Chemical Formulas

A chemical formula is a shorthand method for representing a chemical compound. A formula consists of a collection of chemical symbols, telling the kinds and numbers of atoms present in the compound. In many cases, the chemical formula for a compound can be predicted by knowing the combining tendencies (**valences**) of the elements involved. In writing formulas, it is helpful to know that, in some situations, certain groups of atoms tend to stay together and behave as if they were individual atoms. Such groups of atoms are known as **radicals** or **polyatomic ions**. A list of common polyatomic ions is found in Appendix 5 (page 193). In addition, the student should be aware that some elements have more than one valence. A list of these multivalenced elements is found in Appendix 6 (page 194).

Solved Examples

Example 1: Write the correct chemical formula for sodium chloride.

Solution: Both sodium and chlorine have a single valence: +1 in the case of sodium, and −1 in the case of chlorine. Thus, a single atom of sodium will combine with a single atom of chlorine, and the correct chemical formula for the compound is NaCl.

Example 2: Write the correct chemical formula for potassium oxide.

Solution: The valences of potassium and oxygen are +1 and −2 respectively. Thus, it will take two potassium atoms to satisfy a single oxygen atom, and the correct formula for the compound is K_2O.

A shortcut method for determining formulas exists. The valence for each element or polyatomic ion becomes the subscript for the other element or ion in the compound:

$K^{+1}\ O^{-2}$ becomes K_2O

Example 3: Write the correct chemical formula for sulfuric acid.

Solution: In an acid, the positive element is always hydrogen. The name "sulfur**ic**" is related to the polyatomic ion named "sul**fate.**" So, the two parts of the compound are hydrogen (H^{+1}) and sulfate (SO_4^{-2}). And the correct formula is:

$$H^{+1}\ SO_4^{-2} \quad \text{or} \quad H_2SO_4$$

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Writing Chemical Formulas

Practice Problems (Level 1)

Write the correct chemical formula for each of the following compounds.

1. potassium bromide
2. lithium iodide
3. magnesium chloride
4. hydrogen sulfide
5. sodium oxide
6. calcium fluoride
7. beryllium chloride
8. aluminum bromide
9. hydrogen iodide
10. magnesium sulfide
11. calcium hydroxide
12. barium sulfide
13. zinc bromide
14. scandium chloride
15. magnesium oxide
16. sodium hydroxide
17. gallium iodide
18. hydrogen phosphide
19. nitric acid
20. potassium hydroxide

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Writing Chemical Formulas

Practice Problems (Level 2)

Write the correct chemical formula for each of the following compounds.

1. potassium sulfate
2. magnesium selenide
3. aluminum oxide
4. beryllium hydroxide
5. sodium nitrate
6. phosphoric acid
7. tin (IV) chloride
8. ferric hydroxide
9. copper (I) sulfate
10. scandium sulfide
11. ammonium fluoride
12. calcium sulfate
13. magnesium phosphate
14. acetic acid
15. sodium bisulfate
16. cupric nitrate
17. iron (III) sulfite
18. chromium phosphide
19. strontium bicarbonate
20. ammonium phosphate

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Writing Chemical Formulas

Practice Problems (Level 3)

Write the correct chemical formula for each of the following compounds.

1. lead (II) bicarbonate
2. manganous acetate
3. disodium hydrogen phosphate
4. perbromic acid
5. cupric sulfate pentahydrate
6. diphosphorus pentoxide
7. ammonium dichromate
8. sodium cyanate
9. hypochlorous acid
10. gold (III) sulfate
11. ferric ferrocyanide
12. potassium uranate
13. potassium aluminum sulfate
14. nitrous oxide
15. tin (II) permanganate
16. chromic acid
17. sodium dihydrogen phosphate
18. cadmium acetate
19. ammonium silicate
20. iodous acid

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Writing Chemical Formulas

Teacher Notes

1. Explanation of Levels

Level 1: Problems include elements that demonstrate only a single valence; no polyatomic ions included; primarily uni- and bivalenced elements; no unusual names.

Level 2: Problems include elements with any valence; multivalenced elements also included; a few common polyatomic ions included; no unusual names.

Level 3: Problems include all elements and all polyatomic ions of any valence; some less familiar names are included.

2. Answers

Level 1: 1. KBr; 2. LiI; 3. $MgCl_2$; 4. H_2S; 5. Na_2O; 6. CaF_2; 7. $BeCl_2$; 8. $AlBr_3$; 9. HI; 10. MgS; 11. $Ca(OH)_2$; 12. BaS; 13. $ZnBr_2$; 14. $ScCl_3$; 15. MgO; 16. $NaOH$; 17. GaI_3; 18. H_3P; 19. HNO_3; 20. KOH

Level 2: 1. K_2SO_4; 2. $MgSe$; 3. Al_2O_3; 4. $Be(OH)_2$; 5. $NaNO_3$; 6. H_3PO_4; 7. $SnCl_4$; 8. $Fe(OH)_3$; 9. Cu_2SO_4; 10. $Sc_2(SO_4)_3$; 11. NH_4F; 12. $CaSO_4$; 13. $Mg_3(PO_4)_2$; 14. $HC_2H_3O_2$; 15. $NaHSO_4$; 16. $Cu(NO_3)_2$; 17. $Fe_2(SO_3)_3$; 18. CrP; 19. $Sr(HCO_3)_2$; 20. $(NH_4)_3PO_4$

Level 3: 1. $Pb(HCO_3)_2$; 2. $Mn(C_2H_3O_2)_2$; 3. Na_2HPO_4; 4. $HBrO_4$; 5. $CuSO_4 \cdot 5H_2O$; 6. P_2O_5; 7. $(NH_4)_2Cr_2O_7$; 8. $NaCNO$; 9. $HClO$; 10. $Au_2(SO_4)_3$; 11. $Fe_4[Fe(CN)_6]_3$; 12. KUO_4; 13. $KAl\,(SO_4)_2$; 14. N_2O; 15. $Sn(MnO_4)_2$; 16. H_2CrO_4; 17. NaH_2PO_4; 18. $Cd(C_2H_3O_2)_2$; 19. $(NH_4)_2SiO_4$; 20. HIO_2

Chemical Nomenclature

The term **nomenclature** refers to a system for naming objects. Chemical nomenclature is the system used for naming chemical compounds. Precise rules for naming compounds are fairly detailed and extensive. Space does not permit us to describe all of these rules here. The student should refer to any standard chemistry text or handbook in chemistry. The solved examples illustrate the application of these rules to specific compounds.

Solved Examples

Example 1: Name the compound whose formula is KI.

Solution: The compound KI is a salt. In naming a salt, we give the name of the positive part of the compound (**cation**) first, followed by the name of the negative part of the compound (**anion**). This makes the name of this compound

potassium (K) + iodide (I) = potassium iodide

Example 2: Name the compound whose formula is $Ca(NO_3)_2$.

Solution: This is also a salt, and it is named as described above.

calcium (Ca) + nitrate (NO_3) = calcium nitrate

Example 3: Name the compound whose formula is $Fe(OH)_3$.

Solution: The compound $Fe(OH)_3$ is a base. Bases are named by giving the name of the positive element (cation) and adding the word "hydroxide." Thus, this compound is named:

iron (III) + hydroxide = iron (III) hydroxide

Note that iron is a multivalenced element. So, we must indicate the oxidation state of iron in this compound. We have done this by inserting the valence of iron, expressed in Roman numerals, following the name of the element. It would also be possible to use the "-ic"/"-ous" suffix system. In that case, the name of the compound would be

ferric (+3 is higher valence of iron) + hydroxide

= ferric hydroxide

Example 4: Name the compound whose formula is HNO_3.

Solution: The compound HNO_3 is a ternary (three-element) acid. Ternary acids are named by modifying the name of the polyatomic ion ($-NO_3$ = nitrate) and adding the word "acid." So, the compound is named as

nitric ("-ate" ions become "-ic" acids) + acid

= nitric acid

Example 5: Name the compound whose formula is SO_2.

Solution: The compound SO_2 is a non-metallic oxide. Non-metallic oxides are named by adding the name of the non-metal (sulfur) to the word "oxide," to which has been added a prefix which tells the number of oxygen atoms in the oxide (in this case, "di-" for two). The name of this compound, then, is

sulfur (for S) + dioxide (for "2 oxygens")

= sulfur dioxide

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Chemical Nomenclature

Practice Problems (Level 1)

Give the correct names for each of the compounds listed below.

1. NaCl

2. KF

3. LiI

4. $CaSO_4$

5. $AlBr_3$

6. HCl

7. KOH

8. CO_2

9. HNO_3

10. $NaC_2H_3O_2$

11. K_2O

12. NH_4Cl

13. $Al_2(SO_4)_3$

14. SO_3

15. NH_4NO_3

16. H_2SO_4

17. $Al(OH)_3$

18. ZnO

19. $NaHCO_3$

20. $(NH_4)_3PO_4$

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Chemical Nomenclature

Practice Problems (Level 2)

Give the correct names for each of the compounds listed below.

1. $Ba(OH)_2$
2. H_3PO_4
3. $FeCl_3$
4. N_2O
5. HF
6. NH_4OH
7. $PbSO_4$
8. Cu_2S
9. CO
10. $Sn(HCO_3)_4$
11. P_2O_5
12. $HC_2H_3O_2$
13. Hg_2SO_4
14. $FePO_4$
15. Al_2O_3
16. N_2O_3
17. $CuCO_3$
18. $Co_2(SO_3)_3$
19. $Ba(BrO_3)_2$
20. $SnCrO_4$

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Chemical Nomenclature

Practice Problems (Level 3)

Give the correct names for each of the compounds listed below.

1. Hg_2Cl_2
2. N_2O_5
3. $CoSO_4 \cdot 6H_2O$
4. $FeSeO_4$
5. Er_2O_3
6. K_2Te
7. LiH
8. HFO_2
9. SnI_4
10. As_2O_5
11. $OsCl_4$
12. NaH_2PO_4
13. $Pb(C_2H_3O_2)_2$
14. $Cr_2(SO_3)_3$
15. $Al(MnO_4)_3$
16. Hg_3N_2
17. $KAl(SO_4)_2$
18. SiF_4
19. Au_3P_2
20. $Fe[Fe(CN)_6]$

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Chemical Nomenclature

Teacher Notes

1. Explanation of Levels

Level 1: Problems include simplest possible examples; primarily common salts and bases and most familiar oxides and acids; no multivalenced cations.

Level 2: Problems include more advanced compounds of all kinds; multivalenced cations included.

Level 3: Problems include all possible compounds, including some less familiar examples.

2. Answers

Level 1:

1. sodium chloride
2. potassium fluoride
3. lithium iodide
4. calcium sulfate
5. aluminum bromide
6. hydrochloric acid
7. potassium hydroxide
8. carbon dioxide
9. nitric acid
10. sodium acetate
11. potassium oxide
12. ammonium chloride
13. aluminum sulfate
14. sulfur trioxide
15. ammonium nitrate
16. sulfuric acid
17. aluminum hydroxide
18. zinc oxide
19. sodium bicarbonate (hydrogen carbonate)
20. ammonium phosphate

Level 2:

1. barium hydroxide
2. phosphoric acid
3. ferric or iron (III) chloride
4. nitrous oxide
5. hydrofluoric acid
6. ammonium hydroxide
7. lead (II) or plumbous sulfate
8. copper (I) or cuprous sulfide

9. carbon monoxide
10. tin (IV) or stannic bicarbonate or hydrogen carbonate
11. diphosphorus pentoxide
12. acetic acid
13. mercury (I) or mercurous sulfate
14. iron (III) or ferric phosphate
15. aluminum oxide
16. dinitrogen trioxide
17. copper (II) or cupric carbonate
18. cobalt (III) or cobaltic sulfate
19. barium bromate
20. tin (II) or stannous chromate

Level 3:

1. mercury (I) or mercurous chloride
2. dinitrogen pentoxide
3. cobalt (II) or cobaltous sulfate hexahydrate
4. iron (II) or ferrous selenate
5. erbium oxide
6. potassium telluride
7. lithium hydride
8. fluorous acid
9. tin (IV) or stannic iodide or tetraiodide
10. arsenic (V) or arsenic oxide or pentoxide
11. osmium tetrachloride
12. sodium dihydrogen phosphate
13. lead (II) or plumbous acetate
14. chromium (III) or chromous sulfite
15. aluminum permanganate
16. mercury (II) or mercuric nitride
17. potassium aluminum sulfate
18. silicon tetrafluoride
19. gold (II) or auric phosphide
20. iron (III) or ferric ferricyanide

Reaction Prediction

In many situations, one can predict the type of reaction which will occur if two substances are combined with each other. Or, it may be possible to say whether no reaction at all will take place. For each of the four most common types of reactions—**synthesis**, **decomposition**, **single** and **double replacement**—certain rules allow one to make these predictions. Space does not permit us to list these rules in detail here. However, the solved examples that follow illustrate the **application** of these rules in specific situations.

Solved Examples

Example 1: Predict whether a reaction between zinc metal and chlorine gas will take place.

Solution: The reaction between zinc and chlorine is potentially a synthesis reaction. A synthesis reaction will occur provided the two elements involved are capable of demonstrating opposite valences. In this case, zinc has a valence of +2 and chlorine, of −1. Thus,

$$Zn + Cl_2 \rightarrow ZnCl_2$$

is the expected reaction.

Example 2: Predict whether a reaction will occur when mercury (II) oxide is heated.

Solution: This is potentially a decomposition reaction. Almost any decomposition reaction will occur provided the proper conditions are supplied. The expected reaction, then, is

$$2\,HgO \xrightarrow{\Delta} 2\,Hg + O_2\uparrow$$

Example 3: Predict whether a reaction between silver metal and copper (II) nitrate solution will occur.

Solution: In a single replacement reaction like this one, the criterion is whether the element or the cation with which it is in "competition" is more active. According to the Activity Series (Appendix 7, page 195), copper is the more active element. Therefore, no reaction will occur.

$$Ag + Cu(NO_3)_2 \rightarrow NR$$

Example 4: Predict the nature of the reaction, if any, between solutions of barium nitrate and sodium sulfate.

Solution: The criterion for a double replacement reaction like this one is whether both reactants are soluble in water (Appendix 8, page 196). Since they are, the reaction will occur, as shown here.

$$Ba(NO_3)_2 + Na_2SO_4 \rightarrow BaSO_4\downarrow + 2\ NaNO_3$$

Also, since $BaSO_4$ is insoluble, the reaction goes to completion.

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Reaction Prediction

Practice Problems (Level 1)

For each of the following problems, tell

1. what type of reaction might be expected
2. whether the reaction will occur or not
3. if not, why it will not occur; then write the symbols and formulas for the reactants
4. if so, what the balanced equation for the reaction is
5. in the case of double replacement reactions, whether the reaction goes to completion or not

1. potassium and iodine

2. water $\xrightarrow{\text{electrolyzed}}$

3. zinc and lead (II) chloride

4. sodium nitrate and ammonium chloride

5. mercury and cadmium nitrate

6. manganese and sodium

7. silver nitrate and hydrogen sulfide

8. potassium bromide $\xrightarrow{\text{electrolyzed}}$

9. mercury (I) sulfide and ammonium nitrate

10. hydrogen and oxygen

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Reaction Prediction

Practice Problems (Level 2)

For each of the following problems, tell

1. what type of reaction might be expected
2. whether the reaction will occur or not
3. if not, why it will not occur; then write the symbols and formulas for the reactants
4. if so, what the balanced equation for the reaction is
5. in the case of double replacement reactions, whether the reaction goes to completion or not

1. tin and copper (II) sulfate

2. iron (III) nitrate and sodium chromate

3. calcium and iodine

4. magnesium and hydrochloric acid

5. calcium oxide $\xrightarrow{\text{electrolyzed}}$

6. carbon and oxygen

7. sodium carbonate and sulfuric acid

8. iron (II) sulfide $\xrightarrow{\text{electrolyzed}}$

9. platinum and lead (II) nitrate

10. lithium oxide and water

NAME________________________ DATE______________

Reaction Prediction

Practice Problems (Level 3)

For each of the following problems, tell

1. what type of reaction might be expected
2. whether the reaction will occur or not
3. if not, why it will not occur; then write the symbols and formulas for the reactants
4. if so, what the balanced equation for the reaction is
5. in the case of double replacement reactions, whether the reaction goes to completion or not

1. aluminum and sulfuric acid

2. ammonium phosphate and lithium hydroxide

3. chlorine and fluorine

4. sodium carbonate $\xrightarrow{\text{heated}}$

5. potassium chlorate $\xrightarrow{\text{heated}}$

6. hydrogen and sodium

7. calcium oxide and water

8. potassium sulfate and zinc phosphate

9. ammonium sulfite and phosphoric acid

10. aluminum hydroxide $\xrightarrow{\text{heated}}$

11. bromine and sodium iodide

12. rubidium fluoride and ammonium nitrate

Reaction Prediction

Teacher Notes

1. Explanation of Levels

Level 1: Predictions are all relatively simply made; equations are all easily balanced.

Level 2: Predictions are somewhat more difficult than in Level 1; equations are somewhat more difficult to balance.

Level 3: Predictions involve less familiar instances, such as synthesis of oxides with water and decomposition of ternary salts; equations are somewhat more difficult to balance.

2. Answers

Level 1: (S = synthesis; D = decomposition; SR = single replacement; DR = double replacement)

1. S; yes; $2\,K + I_2 \rightarrow 2\,KI$
2. D; yes; $2\,H_2O \xrightarrow{\sim} 2\,H_2\uparrow + O_2\uparrow$
3. SR; yes; $Zn + PbCl_2 \rightarrow Pb + ZnCl_2$
4. DR; yes; $NaNO_3 + NH_4Cl \rightleftharpoons NaCl + NH_4NO_3$
5. SR; no; $Hg + Cd(NO_3)_2$
6. S; no; $Mn + Na$
7. DR; yes; $2\,AgNO_3 + H_2S \rightarrow Ag_2S\downarrow + 2\,HNO_3$
8. D; yes; $2\,KBr \xrightarrow{\sim} 2\,K + Br_2$
9. DR; no; $Hg_2SO_4 + NH_4NO_3$
10. S; yes; $2\,H_2 + O_2 \rightarrow 2\,H_2O$

Level 2:

1. SR; yes; $Sn + CuSO_4 \rightarrow Cu + SnSO_4$
2. DR; yes; $2\,Fe(NO_3)_3 + 3\,Na_2CrO_4 \rightarrow Fe_2(CrO_4)_3\downarrow$ $+ 6NaNO_3$
3. S; yes; $Ca + I_2 \rightarrow CaI_2$
4. SR; yes; $Mg + 2\,HCl \rightarrow MgCl_2 + H_2\uparrow$
5. D; yes; $2\,CaO \xrightarrow{\Delta} 2\,Ca + O_2$
6. S; yes; $C + O_2 \xrightarrow{\Delta} CO_2\uparrow$
7. DR; yes; $Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 +$ $H_2CO_3(\rightarrow H_2O + CO_2\uparrow)$
8. D; yes; $FeS \xrightarrow{\sim\sim} Fe + S$
9. SR; no; $Pt + Pb(NO_3)_2$
10. S; yes; $Li_2O + H_2O \rightarrow 2\,LiOH$

Level 3:

1. SR; yes; $2\,Al + 3\,H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3\,H_2$
2. DR; yes; $(NH_4)_3PO_4 + 3\,LiOH \rightarrow Li_3PO_4 +$ $3\,NH_4OH(\rightarrow 3\,NH_3\uparrow + 3\,H_2O)$
3. S; yes; $Cl_2 + F_2 \rightarrow 2\,ClF$
4. D; yes; $Na_2CO_3 \xrightarrow{\Delta} Na_2O + CO_2\uparrow$
5. D; yes; $2\,KClO_3 \rightarrow 2\,KCl + 3\,O_2\uparrow$
6. S; yes; $H_2 + 2\,Na \rightarrow 2\,NaH$
7. S; yes; $CaO + H_2O \rightarrow Ca(OH)_2$
8. DR; no; $K_2SO_4 + Zn_3(PO_4)_2$
9. DR; yes; $3\,(NH_4)_2SO_3 + 2\,H_3PO_4 \rightarrow 2(NH_4)_3PO_4 +$ $3\,H_2SO_3(\rightarrow 3\,H_2O + 3\,SO_2\uparrow)$
10. D; yes; $2\,Al(OH)_3 \xrightarrow{\Delta} Al_2O_3 + 3\,H_2O$
11. SR; yes; $Br_2 + 2\,NaI \rightarrow I_2 + 2\,NaBr$
12. DR; yes; $RbF + NH_4NO_3 \rightleftharpoons RbNO_3 + NH_4F$

Equation Writing

A **chemical equation** is a way of using chemical symbols and formulas to express the changes that take place in a chemical reaction. All chemical equations must conform to the law of **conservation of mass**. This means that no atoms may be created in or lost from any chemical change. The process by which a chemical equation is made to conform to this law is known as **balancing the equation**.

Balancing an equation involves basically two steps.

1. Formulas for all substances involved in the reaction must be written correctly. If an improper formula is assigned to a reactant or a product, the equation cannot be properly balanced.

2. Coefficients must be assigned to all formulas such that the number of atoms of each kind on the left side of the equation is identical to the number of atoms of the same kind on the right side. Formulas of substances may NOT be rewritten in order to achieve this balance.

Solved Examples

Example 1: Balance this equation: $Ca + HCl \rightarrow CaCl_2 + H_2\uparrow$

Solution: In balancing an equation, it is usually best to leave hydrogens and oxygens to the end. Notice in this case that there is only one "Cl" on the left and two on the right. To balance the "Cl's," place a coefficient of 2 in front of the HCl:

$$Ca + 2\,HCl \rightarrow CaCl_2 + H_2\uparrow$$

Notice that the number of hydrogens also balances now. The full equation is balanced.

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Example 2: Write and balance the equation for the reaction in which sodium carbonate and aluminum hydroxide react to form sodium hydroxide and aluminum carbonate.

Solution: Begin by writing the correct formulas for all substances involved:

$$Na_2CO_3 + Al(OH)_3 \rightarrow NaOH + Al_2(CO_3)_3$$

In balancing this equation, think of all polyatomic ions as a single element. Balance carbonates (CO_3) first. Since there are 3 on the right in $Al_2(CO_3)_3$, add a coefficient of 3 in front of the Na_2CO_3:

$$3\ Na_2CO_3 + Al(OH)_3 \rightarrow NaOH + Al_2(CO_3)_3$$

Now balance sodiums (Na). Since there are 6 (3×2) on the left, add a coefficient of 6 in front of the NaOH on the right.

$$3\ Na_2CO_3 + Al(OH)_3 \rightarrow 6\ NaOH + Al_2(CO_3)_3$$

Now balance hydroxides (OH). There are 6 on the right in 6 NaOH. So add a coefficient of 2 ($2 \times 3 = 6$) in front of the $Al(OH)_3$. Notice that this balances the aluminums (Al) at the same time, and the equation is fully balanced.

$$3\ Na_2CO_3 + 2\ Al(OH)_3 \rightarrow 6\ NaOH + Al_2(CO_3)_3$$

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Equation Writing

Practice Problems (Level 1)

1. Balance each of the following equations.

a. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2\uparrow$

b. $Na + Br_2 \rightarrow NaBr$

c. $H_2O \rightleftharpoons H_2\uparrow + O_2\uparrow$

d. $Cl_2 + KI \rightarrow KCl + I_2$

e. $HNO_3 + LiOH \rightarrow H_2O + LiNO_3$

f. $N_2 + O_2 \rightarrow N_2O$

g. $NH_4Cl + NaOH \rightarrow NH_4OH + NaCl$

h. $HCl + K_2CO_3 \rightarrow KCl + H_2O + CO_2\uparrow$

2. Balance each of the following equations.

a. sodium + iodine → sodium iodide

b. zinc + hydrobromic acid → zinc bromide + hydrogen ↑

c. potassium hydroxide $\xrightarrow{\text{heated}}$ potassium oxide + water

d. magnesium + water → magnesium hydroxide + hydrogen ↑

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Equation Writing

Practice Problems (Level 2)

1. Balance each of the following equations.

   a. $K_3PO_4 + HCl \rightarrow KCl + H_3PO_4$

   b. $Na + HNO_3 \rightarrow NaNO_3 + H_2\uparrow$

   c. $S + O_2 \rightarrow SO_3\uparrow$

   d. $Ca(ClO_3)_2 \xrightarrow{\Delta} CaCl_2 + O_2\uparrow$

2. Balance each of the following equations.

   a. potassium iodide + lead (II) nitrate → potassium nitrate + lead (II) iodide

   b. iron (III) oxide + carbon → carbon monoxide ↑ + iron

   c. mercury (II) oxide $\xrightarrow{\text{heated}}$ mercury + oxygen ↑

   d. calcium + aluminum chloride → calcium chloride + aluminum

   e. mercury (I) nitrate + sodium carbonate → sodium nitrate + mercury (I) carbonate

   f. potassium bromide + aluminum nitrate → potassium nitrate + aluminum bromide

   g. calcium phosphate + aluminum sulfate → calcium sulfate + aluminum phosphate

   h. rubidium + acetic acid → rubidium acetate + hydrogen ↑

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Equation Writing

Practice Problems (Level 3)

Write and balance chemical equations for each of the following reactions.

1. nitrogen + hydrogen → ammonia ↑ (NH_3)
2. butane (C_4H_{10}) + oxygen $\xrightarrow{\Delta}$ carbon dioxide ↑ + water ↑
3. aluminum oxide → aluminum + oxygen ↑
4. ethyl alcohol (C_2H_5OH) + oxygen $\xrightarrow{\Delta}$ carbon monoxide ↑ + water ↑
5. nitrogen + oxygen → dinitrogen pentoxide ↑
6. octane (C_8H_{18}) + oxygen $\xrightarrow{\Delta}$ carbon dioxide ↑ + water ↑
7. aluminum sulfate + phosphoric acid → aluminum phosphate + sulfuric acid
8. diphosphorus pentoxide + water → phosphoric acid
9. ammonia + nitric oxide → nitrogen ↑ + water
10. iron (III) oxide + carbon monoxide → iron + carbon dioxide ↑
11. copper + nitric acid → copper (II) nitrate + nitric oxide ↑ + water
12. iron (II) sulfide + oxygen → iron (III) oxide + sulfur dioxide ↑

Equation Writing

Teacher Notes

1. Explanation of Levels

Level 1: Two thirds of equations given in formula form; coefficients of 1 or 2 required for balancing.

Level 2: One third of equations given in formula form; more difficult balancing problems.

Level 3: All problems stated in word form only; even more difficult problems of balancing.

2. Answers

(Numbers correspond to coefficients of each species in reaction.)

Level 1:

1. a. OK (means balanced as is); b. 2, 1, 2; c. 2, 2, 1; d. 1, 2, 2, 1; e. OK; f. 2, 1, 2; g. OK; h. 2, 1, 2, 1, 1
2. a. $2\ Na + I_2 \rightarrow 2\ NaI$
   b. $Zn + 2\ HBr \rightarrow ZnBr_2 + H_2\uparrow$
   c. $2\ KOH \xrightarrow{\Delta} K_2O + H_2O$
   d. $Mg + 2\ H_2O \rightarrow Mg(OH)_2 + H_2\uparrow$

Level 2:

1. a. 1, 3, 3, 1; b. 2, 2, 2, 1; c. 2, 3, 2; d. 1, 1, 3
2. a. $2\ KI + Pb(NO_3)_2 \rightarrow 2\ KNO_3 + PbI_2$
   b. $Fe_2O_3 + 3C \rightarrow 3\ CO\uparrow + 2\ Fe$
   c. $2\ HgO \xrightarrow{\Delta} 2\ Hg + O_2\uparrow$

d. $3\ Ca + 2\ AlCl_3 \rightarrow 3\ CaCl_2 + 2\ Al$

e. $2\ HgNO_3 + Na_2CO_3 \rightarrow 2\ NaNO_3 + Hg_2CO_3$

f. $3\ KBr + Al(NO_3)_3 \rightarrow 3\ KNO_3 + AlBr_3$

g. $Ca_3(PO_4)_2 + Al_2(SO_4)_3 \rightarrow 3\ CaSO_4 + 2\ AlPO_4$

h. $2\ Rb + 2\ HC_2H_3O_2 \rightarrow 2\ RbC_2H_3O_2 + H_2\uparrow$

Level 3:

1. $N_2 + 3\ H_2 \rightarrow 2\ NH_3\uparrow$
2. $2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2\uparrow + 10\ H_2O\uparrow$
3. $2\ Al_2O_3 \rightarrow 4\ Al + 3\ O_2\uparrow$
4. $C_2H_5OH + 2\ O_2 \rightarrow 2\ CO\uparrow + 3\ H_2O\uparrow$
5. $2\ N_2 + 5\ O_2 \rightarrow 2\ N_2O_5$
6. $2\ C_8H_{18} + 25\ O_2 \xrightarrow{\Delta} 16\ CO_2\uparrow + 18\ H_2O\uparrow$
7. $Al_2(SO_4)_3 + 2\ H_3PO_4 \rightarrow 2\ AlPO_4 + 3\ H_2SO_4$
8. $P_2O_5 + 3\ H_2O \rightarrow 2\ H_3PO_4$
9. $4\ NH_3 + 6\ NO \rightarrow 5\ N_2\uparrow + 6\ H_2O\uparrow$
10. $Fe_2O_3 + 3\ CO \rightarrow 3\ CO_2\uparrow + 2\ Fe$
11. $3\ Cu + 8\ HNO_3 \rightarrow 3\ Cu(NO_3)_2 + 2\ NO\uparrow + 4\ H_2O$
12. $4\ FeS + 7\ O_2 \rightarrow 2\ Fe_2O_3 + 4\ SO_2\uparrow$

Molecular Weight and Moles

The relative weights of elements are represented by their atomic weights. The relative weights of compounds are represented by their molecular or formula weights. (The terms **atomic mass, molecular mass**, and **formula mass** are more precise, but the use of the term **weight** still tends to be more common.)

The molecular or formula weight of a compound is equal to the sum of the weights of all atoms present in the molecule or formula unit. The molecular or formula weight of a substance is often expressed in units of **grams**. At one time, the quantity gram molecular weight or gram formula weight was given the designation **mole**. Today, a somewhat more technical definition is given for this term. A mole is defined as the amount of a substance containing $6.02 + 10^{23}$ particles (atoms, molecules, ions, etc.). This number is known as the **Avogadro number**.

Solved Examples

Example 1: Find the molecular weight of ethyl alcohol, C_2H_5OH.

Solution: The molecular weight is equal to the sum of the weights of all atoms in the molecule.

$$\begin{aligned}\text{Molecular weight} &= 2 \times C + 6 \times H + 1 \times O \\ &= 2 \times 12.01 + 6 \times 1.01 + 1 \times 16.00 \\ &= 46.08\end{aligned}$$

Example 2: Find the weight of one mole (1.00 mol) of sulfuric acid.

Solution:

$$\begin{aligned}1.00 \text{ mol of } H_2SO_4 &= 1.00 \text{ gram molecular weight (gmw) of } H_2SO_4 \\ &= 2 \times H + 1 \times S + 4 \times O \text{ (in grams)} \\ &= 2 \times 1.01 + 1 \times 32.06 + 4 \times 16.00\text{ g} \\ &= 98.08\text{ g}\end{aligned}$$

Example 3: What is the weight of 0.435 mol of water?

Solution:

$$
\begin{aligned}
0.435 \text{ mol of } H_2O &= 0.435 \text{ mol} \times \text{gmw of } H_2O \\
&= 0.435 \text{ mol} \times (2 \times 1.01 + 1 \times 16.00)\ \frac{\text{g}}{\text{mol}} \\
&= 0.435\ \cancel{\text{mol}} \times 18.02\ \frac{\text{g}}{\cancel{\text{mol}}} \\
&= 7.84 \text{ g}
\end{aligned}
$$

Example 4: How many moles of sodium chloride is 15.0 g of the substance?

Solution:

$$1.00 \text{ mol of NaCl} = 1 \times 23.0 + 1 \times 35.45 \text{ g} = 58.45 \text{ g}$$

$$15.0\ \cancel{\text{g}} \times \frac{1.00 \text{ mol}}{58.45\ \cancel{\text{g}}} = 0.257 \text{ mol of NaCl}$$

NAME____________________________ DATE________________

Molecular Weight and Moles

Practice Problems (Level 1)

1. Find the molecular or formula weight of each of the following compounds.

a. KCl
b. NaF
c. HI
d. LiBr
e. RbBr
f. $BaCl_2$
g. CaI_2
h. Na_2S
i. MgS
j. AlP

2. Find the weight of each of the following, expressed in grams.

a. 1.00 mol of CaO
b. 1.00 mol of BeSe
c. 1.00 mol of KF
d. 1.00 mol of SrO
e. 2.00 mol of MgI_2
f. 2.00 mol of Li_3P
g. 5.00 mol of $CaCl_2$
h. 0.50 mol of $FeBr_2$
i. 0.20 mol of Cu_2O
j. 0.40 mol of Hg_2Cl_2

3. Find the weight of each of the following, expressed in moles.

a. 18.02 g of H_2O
b. 80.92 g of HBr
c. 17.04 g of NH_3
d. 190.42 g of $MgCl_2$
e. 36.75 g of K_2S
f. 25.38 g of SnI_2
g. 11.98 g of FeO
h. 15.0 g of KBr
i. 25.0 g of SrS
j. 50.0 g of AlF_3

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Molecular Weight and Moles

Practice Problems (Level 2)

1. Find the molecular weight or formula weight of each of the following compounds.
   a. HNO_3
   b. Fe_2O_3
   c. H_3PO_4
   d. K_2SO_4
   e. Be_5As_2
   f. ammonium nitrate
   g. rubidium sulfite
   h. lithium carbonate
   i. magnesium hydroxide
   j. aluminum sulfate

2. Find the weight of each of the following, expressed in grams.
   a. 1.00 mol of $HC_2H_3O_2$
   b. 2.50 mol of K_2CrO_4
   c. 0.50 mol of $Ca(ClO_3)_2$
   d. 0.25 mol of $Ba(NO_3)_2$
   e. 0.375 mol of $Na_2Cr_2O_7$
   f. 0.25 mol of sodium acetate
   g. 0.152 mol of phosphoric acid
   h. 0.0582 mol of lithium sulfate
   i. 0.418 mol of iron (III) nitrate
   j. 1.872 mol of copper (II) acetate

3. Find the weight of each of the following, expressed in moles.
   a. 100.00 g of $CaCO_3$
   b. 100.00 g of $Ni(NO_3)_2$
   c. 50.00 g of $C_6H_{12}O_6$
   d. 25.00 g of K_3PO_4
   e. 15.57 g of $Bi(OH)_3$
   f. 3.50 g of arsenic trichloride
   g. 0.572 g of calcium phosphide
   h. 1.750 g of calcium acetate
   i. 4.904 g of aluminum nitrate
   j. 27.85 g of iron (II) phosphate

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Molecular Weight and Moles

Practice Problems (Level 3)

1. Find the molecular weight or formula weight of each of the following compounds.
   a. dinitrogen pentoxide
   b. copper (II) sulfate pentahydrate
   c. stannic hydroxide
   d. iron (III) acetate
   e. potassium aluminum sulfate
   f. ammonium phosphate
   g. scandium sulfate
   h. calcium bicarbonate
   i. ammonium dichromate
   j. aluminum carbonate

2. Find the weight of each of the following, expressed in grams.
   a. 1.00 mol of calcium nitrate
   b. 0.593 mol of calcium chlorate
   c. 0.277 mol of ammonium acetate
   d. 3.552 mol of beryllium nitrate
   e. 1.04 mol of potassium aluminum sulfate
   f. 0.994 mol of zinc phosphate
   g. 1.24 mol of barium perbromate
   h. 0.756 mol of magnesium acetate
   i. 0.391 mol of cadmium bicarbonate
   j. 2.45×10^{-3} mol of iron (II) ferricyanide

3. Find the weight of each of the following, expressed in moles.
   a. 3.00 g of ammonium iodate
   b. 0.504 g of magnesium bicarbonate
   c. 15.88 g of ammonium phosphate
   d. 2.52×10^{-6} g of sucrose ($C_{12}H_{22}O_{11}$)
   e. 1.45×10^{3} g of lanthanum nitrate
   f. 2.957 of sodium acetate
   g. 0.852 g of calcium phosphate
   h. 32.75 g of iron (III) carbonate
   i. 6.02×10^{7} g of lead (II) phosphate
   j. 3.55×10^{-4} g of barium phosphate

Molecular Weight and Moles

Teacher Notes

1. Explanation of Levels

Level 1: Primarily binary compounds used; formulas given for all compounds; mole problems involve simple multiplication and division.

Level 2: Some ternary compounds included; formulas given for some compounds, names for others; problems involve more difficult multiplication and division.

Level 3: Almost entirely ternary compounds included; names given for all compounds; difficult multiplication and division required in almost all cases; some problems involving exponential notation.

2. Answers

Level 1:

1. a. 74.557; b. 41.99; c. 127.92; d. 86.856; e. 165.40; f. 208.28; g. 293.90; h. 78.048; i. 56.39; j. 57.96

2. a. 56.082 g; b. 87.97 g; c. 58.10 g; d. 103.63 g; e. 556.28 g; f. 103.59 g; g. 554.95 g; h. 107.85 g; i. 28.62 g; j. 188.86 g

3. a. 1.001 mol; b. 1.000 mol; c. 6.964 mol; d. .3332 mol e. .0681 mol; f. .1667 mol; g. .126 mol; h. .208 mol; i. .595 mol; j. .01 mol

Level 2:

1. a. 63.016; b. 159.70; c. 97.99506; d. 174.266; e. 194.89; f. 80.048; g. 251.03; h. 73.891; i. 58.34; j. 342.16

2. a. 60.1 g; b. 485 g; c. $1\bar{0}0$ g (103 g); d. 65 g;
   e. 98.3g; f. 21 g; g. 14.9 g; h. 6.40 g; i. 75.2 g;
   j. 340.0 g

3. a. 0.99909 mol; b. 0.54727 mol; c. .2775 mol;
   d. 0.1178 mol; e. 0.05988 mol; f. 0.0193 mol;
   g. 0.00314 mol; h. 0.01106 mol; i. 0.02302 mol;
   j. 0.07790 mol

Level 3:

1. a. 108.016; b. 249.69; c. 186.7192; d. 232.988
   e. 258.20; f. 149.095; g. 378.0846; h. 162.118;
   i. 252.10; j. 233.993

2. a. 164 g; b. 123 g; c. 21.4 g; d. 472.5 g; e. 242 g;
   f. 384 g; g. 527 g; h. 108 g; i. 91.7 g; j. 1.45 g

3. a. 0.0155 mol; b. 0.00344 mol; c. 0.1065 mol;
   d. 7.36×10^{-9} mol; e. 4.46 mol; f. 0.03604 mol;
   g. 0.00275 mol; h. 0.1123 mol; i. 7.42×10^{4} mol;
   j. 5.90×10^{-7} mol

Percentage Composition

The term **percentage composition** refers to the number of parts per hundred of each element in a compound. Percentage composition can be determined in one of two ways. First, if the correct chemical formula for the compound is known, then the percentage of each element is equal to the weight of that element divided by the total weight of the compound. Or, in the compound A_xB_y, the percentage of each element is:

$$\%A = \frac{x \text{ (atomic weight of A)}}{\text{molecular weight of } A_xB_y} \times 100\%$$

$$\%B = \frac{y \text{ (atomic weight of B)}}{\text{molecular weight of } A_xB_y} \times 100\%$$

The percentage composition of a compound can also be determined from laboratory data. If the weight of each element present in a compound can be determined, then the percentage of each is equal to the weight of the element present in the compound divided by the total weight of the compound. The second example (next page) illustrates this method.

Solved Examples

Example 1: What is the percentage composition of calcium chloride, $CaCl_2$?

Solution: As indicated above,

$$\%Ca = \frac{\text{atomic weight of Ca}}{\text{molecular weight of } CaCl_2} \times 100\%$$

$$= \frac{40.08}{110.98} \times 100\% = 36.1\%$$

$$\%Cl = \frac{2 \times \text{atomic weight of Cl}}{\text{molecular weight of } CaCl_2} \times 100\%$$

$$= \frac{2 \times 35.46}{110.98} \times 100\% = 63.90\%$$

Example 2: A sample of an unknown gas is found to consist of 10.48 g of nitrogen and 11.96 g of oxygen. What is the percentage composition of this gas?

Solution: The total weight of the compound = 10.48 g + 11.96 g
= 22.44 g

$$\%N = \frac{\text{weight of nitrogen}}{\text{total weight of gas}} \times 100\%$$

$$\%O = \frac{\text{weight of oxygen}}{\text{total weight of gas}} \times 100\%$$

$$\%N = \frac{10.48\text{ g}}{22.44\text{ g}} \times 100\% = 46.70\%$$

$$\%O = \frac{11.96\text{ g}}{22.44\text{ g}} \times 100\% = 53.30\%$$

NAME ______________________ DATE ____________

Percentage Composition

Practice Problems (Level 1)

Find the percentage composition of each compound listed below. In the first ten problems, the correct formula is given. In the next four problems, laboratory data for the compound are presented.

1. FeO
2. $MgCl_2$
3. CH_4
4. CS_2
5. NO_2
6. HgO
7. SnI_4
8. Cu_2O
9. NH_3
10. CH_2O
11. What is the percentage composition of an ore which contains 1.85 g of aluminum and 1.65 g of oxygen?
12. A compound of silver contains 4.35 g of that metal combined with 0.65 g of sulfur. What is the percentage composition of this compound?
13. A certain sample of gas weighing 22.8 g contains 6.22 g of carbon and an unweighed quantity of oxygen. What is the percentage composition of this gas?
14. A compound consisting of aluminum and chlorine weighs 17.82 g. The aluminum in the compound weighs 3.60 g. What is the weight of chlorine in the compound? What is the percentage composition of the compound?

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Percentage Composition

Practice Problems (Level 2)

Find the percentage composition of each compound listed below. In the first eight problems, the correct formula or name is given. In the next six problems, laboratory data for the compound are presented.

1. KNO_2
2. NH_4Cl
3. $SrCl_2$
4. $KMnO_4$
5. sulfuric acid
6. potassium phosphate
7. ammonium bromide
8. barium hydroxide
9. Analysis of a compound shows that it consists of 43.40 g of copper and 10.95 g of sulfur. What is the percentage composition of this compound?
10. A sample of benzene is analyzed and found to consist of 13.74 g of carbon and 1.15 g of hydrogen. What is the percentage composition of benzene?
11. Analysis of an unknown compound shows that it consists of 21.8 g of oxygen, 4.09 g of aluminum, and 6.36 g of nitrogen. What is the percentage composition of this compound?
12. A compound consisting of carbon, hydrogen, and oxygen weighs 40.85 g. Analysis shows that the compound contains 10.90 g of carbon and 0.90 g of hydrogen. What is the percentage composition of the compound?
13. An organic compound consisting of carbon, hydrogen, and oxygen only weighs 13.669 g. Analysis shows that the compound contains 0.547 g and 8.707 g of the last two of these elements, respectively. What is the percentage composition of this compound?
14. Analysis of an ore of calcium shows it to contain 13.61 g of calcium and 21.77 g of oxygen in a sample weighing 46.28 g. What is the percentage composition of this compound?

NAME ______________________ DATE ______________

Percentage Composition

Practice Problems (Level 3)

Find the percentage composition of each compound listed below. In the first eight problems, the correct name is given. In the next six problems, laboratory data for the compound are presented.

1. zinc carbonate

2. ammonium sulfate

3. iron (III) oxide

4. calcium phosphate

5. ammonium dichromate

6. aluminum nitrate

7. calcium acetate

8. ammonium carbonate

9. A sample of water weighing 89.6 g is found to contain 9.98 g of hydrogen. How much oxygen is present in the water? What is the percentage composition of water?

10. An organic compound is found to contain 10.18 g of carbon, 1.47 g of hydrogen, and 0.93 g of oxygen. What is the percentage composition of this compound?

11. A sample of organic material weighing 0.005587 g is analyzed and found to contain 2.51×10^{-4} g of hydrogen and 1.98×10^{-3} g of oxygen. What is the percentage composition of this compound?

12. Analysis of 0.005644 g of a compound shows that it contains 4.177×10^{-3} g of carbon and 9.76×10^{-4} g of nitrogen. The hydrogen in the compound cannot be determined directly. What is the percentage composition of this compound?

13. Laboratory analysis of a sample of paraldol shows that it contains 4.67 g of carbon, 0.77 g of hydrogen, and 3.12 g of oxygen. What is the percentage composition of paraldol?

14. An organic compound separated into its elements produces 8.2 g of carbon and 1.44 g of hydrogen from a sample weighing 32.8 g. What is the weight of oxygen in this compound? What is the percentage composition of this compound?

Percentage Composition

Teacher Notes

1. Explanation of Levels

Level 1: Formulas for all compounds given; all compounds are binary; only binary compounds used in lab data problems.

Level 2: Compounds listed in first set are given both by formula and by name; both binary and ternary compounds used in both sections.

Level 3: Compounds listed by name only; more complex problems in both sections.

2. Answers

Level 1:

1. 77.7% Fe; 22.3% O
2. 25.4% Mg; 74.6% Cl
3. 74.9% C; 25.1% H
4. 15.8% C; 84.2% S
5. 30.4% N; 69.6% O
6. 92.6% Hg; 7.4% O
7. 19.0% Sn; 81.0% I
8. 88.8% C; 11.2% O
9. 82.2% N; 17.8% H
10. 40.0% C; 6.7% H; 53.3% O
11. 52.9% Al; 47.1% O
12. 87.0% Ag; 13.0% S
13. 27.3% C; 72.7% O
14. 20.2% Al; 79.8% Cl

Level 2:

1. 45.9% K; 16.5% N; 37.6% O
2. 26.2% N; 7.5% H; 66.3% Cl
3. 55.3% Sr; 44.7% Cl;
4. 24.7% K; 34.8% Mn; 40.5% O
5. 2.0% H; 32.7% S; 65.3% O
6. 55.2% K; 14.6% P; 30.2% O
7. 14.3% N; 4.1% H; 81.6% Br
8. 80.1% Ba; 18.7% O; 1.2% H
9. 79.9% Cu; 20.1% S;
10. 92.3% C; 7.7% H

11. 67.6% O; 12.7% Al; 19.7% N
12. 26.7% C; 2.2% H; 71.1% O
13. 29.4% Ca; 23.6% S; 47.0% O
14. 29.4% Ca; 47.0% O; 23.6% unknown

Level 3:

1. 52.1% Zn; 9.6% C; 38.3% O
2. 21.2% N; 6.1% H; 24.3% S; 48.4% O
3. 69.9% Fe; 30.1% O
4. 38.8% Ca; 20.0% P; 41.2% O
5. 11.1% N; 3.2% H; 41.3% Cr; 44.4% O
6. 12.7% Al; 19.7% N; 67.6% O
7. 25.3% Ca; 30.4% C; 3.8% H; 40.5% O
8. 29.1% N; 8.4% H; 12.5% C; 50.0% O
9. 79.6 g; 11.1% H; 88.9% O
10. 80.9% C; 11.7% H; 7.4% O
11. 60.1% C; 4.5% H; 35.4% O
12. 74.0% C; 8.7% H; 17.3% N
13. 54.6% C; 9.0% H; 36.4% O
14. 23.2 g; 25.0% C; 4.4% H; 70.6% O

Simplest (Empirical) Formula

Chemical formulas can often be predicted on the basis of valences of the elements present in the compound. However, the only certain way to know a compound's formula is to determine its composition in the laboratory. Knowing the percentage composition of a compound is, therefore, the first step in determining the chemical formula for that compound.

The **simplest** (or **empirical**) formula of a compound is that formula which shows the simplest ratio of all elements present in the compound, as shown by laboratory analysis. The **simplest** formula may or may not also be the **correct** (or **molecular**) formula (see "Molecular [True] Formula," page 88).

Solved Examples

Example 1: Determine the simplest formula of a compound whose percentage composition is 87.5% nitrogen and 12.5% hydrogen.

Solution: To determine simplest formula from percentage composition, follow these two steps:

Step 1: Divide the percentage composition of each element by that element's atomic weight.

$$\text{For N: } \frac{0.875}{14.01} = 0.0625$$

$$\text{For H: } \frac{0.125}{1.01} = 0.124$$

Step 2: Express these two quotients as the smallest whole number ratio possible. One way to do this is to divide each quotient by the smallest of the quotients.

$$\frac{0.0625}{0.0625} = 1 \qquad \frac{0.124}{0.0625} = 1.98$$

The nearest whole number ratio to these numbers (1 and 1.98) is 1:2. This ratio expresses the ratio of atoms in the compound. The formula for the compound, then, is NH_2.

Example 2: Determine the simplest formula of a compound for which a sample contains 3.60 g of potassium and 1.47 g of sulfur.

Solution: Dimensional analysis allows us to solve this problem directly if we use the gram atomic weight for each element.

$$3.60\text{ g K} \times \frac{1\text{ atom}}{39.10\text{ g}} = 0.0920\text{ atom}$$

$$1.47\text{ g S} \times \frac{1\text{ atom}}{32.06\text{ g}} = 0.0459\text{ atom}$$

The ratio of atoms in this compound, then, is 0.0920:0.0459 or:

$$\text{K:}\ \frac{0.0920}{0.0459} = 2.01 \qquad \text{S:}\ \frac{0.0459}{0.0459} = 1.00$$

This makes the simplest formula: K_2S.

NAME ____________________ DATE ____________

Simplest (Empirical) Formula

Practice Problems (Level 1)

Determine the simplest formula for each compound listed below. In the first eight problems, the percentage composition of the compound is given. In the last two, laboratory data for the compound are presented.

1. 80.0% carbon; 20.0% hydrogen
2. 71.5% calcium; 28.5% oxygen
3. 82.2% nitrogen; 17.8% hydrogen
4. 85.7% carbon; 14.3% hydrogen
5. 6.6% aluminum; 93.4% iodine
6. 92.2% carbon; 7.8% hydrogen
7. 23.5% potassium; 76.5% iodine
8. 50.0% sulfur; 50.0% oxygen
9. An oxide of arsenic contains 3.26 g of arsenic and 1.04 g of oxygen. What is the empirical formula for this oxide?
10. A sample of sodium oxide weighing 12.57 g contains 9.34 g of sodium. What is the empirical formula for this compound?

NAME ______________________ DATE ______________

Simplest (Empirical) Formula

Practice Problems (Level 2)

Determine the simplest formula for each compound listed below. In the first six problems, the percentage composition of the compound is shown. In the last four, laboratory data for the compound are given.

1. 63.5% silver; 8.2% nitrogen; 28.2% oxygen
2. 14.3% nitrogen; 4.1% hydrogen; 81.6% bromine
3. 24.7% potassium; 34.7% manganese; 40.5% oxygen
4. 56.6% potassium; 8.68% carbon; 34.7% oxygen
5. 9.92% carbon; 58.7% chlorine; 31.4% fluorine
6. 37.8% carbon; 6.4% hydrogen; 55.8% chlorine
7. 35.0% nitrogen; 5.0% hydrogen; 60.0% oxygen
8. 27.4% sodium; 1.2% hydrogen; 14.3% carbon; 57.1% oxygen
9. Analysis of a sample of a sulfur acid shows it to contain 0.17 g of hydrogen, 2.82 g of sulfur, and 5.67 g of oxygen. What is the simplest formula for this compound?
10. Analysis of a salt results in the following composition: 3.47 g of sodium; 2.12 g of nitrogen; and 7.27 g of oxygen. What is the empirical formula for this salt?
11. A barium salt is found to contain 21.93 g of barium, 5.12 g of sulfur, and 10.24 g of oxygen. What is the simplest formula of this compound?
12. An ore containing zinc, carbon, and oxygen and weighing 485.35 g is analyzed and found to contain 46.59 g of carbon and 186.37 g of oxygen. What is the simplest formula for this compound?

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Simplest (Empirical) Formula

Practice Problems (Level 3)

Determine the simplest formula for each compound listed below. In the first six problems, the percentage composition of the compound is given. In the last four, laboratory data for the compound are given.

1. 26.6% potassium; 35.4% chromium; 38.0% oxygen
2. 74.0% carbon; 8.7% hydrogen; 17.3% nitrogen
3. 69.8% iron; 30.2% oxygen
4. 83.7% carbon; 16.3% hydrogen
5. 50.8% zinc; 16.0% phosphorus; 33.2% oxygen
6. 21.2% nitrogen; 6.1% hydrogen; 24.3% sulfur; 48.4% oxygen
7. Chemical analysis of a 10.000 g of oil of wintergreen shows that it consists of 6.32 g of carbon, 0.53 g of hydrogen, and 3.16 g of oxygen. What is the simplest formula for oil of wintergreen?
8. An acid is analyzed in the laboratory and the following results are obtained: 3.1% hydrogen, 31.6% phosphorus, 65.3% oxygen. What is the simplest formula of this acid?
9. Examination of 3.2×10^{-2} g of an unknown white powder shows that the powder consists of an unknown amount of nitrogen, 2.6×10^{-3} g of hydrogen, 6.7×10^{-3} g of phosphorus, and 1.37×10^{-2} g of oxygen. What is the simplest formula for this compound?
10. A rock sample weighing 5.88×10^{4} g is known to contain calcium, phosphorus, and oxygen. The amount of the first two elements in this rock is found to be 2.28×10^{4} g and 1.18×10^{4} g respectively. What is the formula for the compound in this rock sample?

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Simplest (Empirical) Formula

Teacher Notes

1. Explanation of Levels

Level 1: Binary compounds in which components are in 1:1, 1:2, or 1:3 ratio only are included.

Level 2: Binary and ternary compounds containing no more than four atoms of any one kind are included.

Level 3: Binary and ternary compounds of higher complexity, containing up to 12 atoms, are included.

2. Answers

Level 1: 1. CH_3; 2. CaO; 3. NH_3; 4. CH_2; 5. AlI_3; 6. CH; 7. KI; 8. SO_2; 9. As_2O_3; 10. Na_2O

Level 2: 1. $AgNO_3$; 2. NH_4Br; 3. $KMnO_4$; 4. K_2CO_3; 5. CCl_2F_2; 6. C_2H_4Cl; 7. $N_2H_4O_3$; 8. $NaHCO_3$; 9. H_2SO_4; 10. $NaNO_3$; 11. $BaSO_4$; 12. $ZnCO_3$

Level 3: 1. $K_2Cr_2O_7$; 2. C_5H_7N; 3. Fe_2O_3; 4. C_3H_7; 5. $Zn_3P_2O_8[Zn_3(PO_4)_2]$; 6. $N_2H_8SO_4[(NH_4)_2SO_4]$; 7. $C_8H_8O_3$; 8. H_3PO_4; 9. $N_3H_{12}PO_4[(NH_4)_3PO_4]$; 10. $Ca_3P_2O_8[Ca_3(PO_4)_2]$

Molecular (True) Formula

The simplest, or empirical, formula of a compound can be determined from the percentage composition of that compound. (See "Simplest [Empirical] Formula," page 82.) But the simplest formula may or may not be the correct formula. For example, compounds with the true molecular formulas of CH, C_2H_2, C_3H_3, and C_6H_6 all have the same simplest formula: CH. But the true formulas obviously differ from each other.

The additional piece of information needed to determine the correct (molecular, true) formula of a compound is its **molecular** (or **formula**) **weight**. If we know, for example, that the molecular weight of the compound whose simplest formula is CH is 39, we are then able to select from all possible formulas of the general form $(CH)_X$. We can do that as follows:

- molecular weight given by simplest formula $= 12 + 1 = 13$
- true molecular weight of compound (given) $= 39$
- ratio of true molecular weight to molecular weight of simplest formula: $39 : 13 = 3 : 1$
- therefore, the true formula must have three times as many atoms of each kind: C_3H_3

The true molecular weight of a compound is often not easy to determine. The most common methods by which that is done is by finding the molar volume of gases or determining certain colligative properties (boiling point elevation or freezing point depression) of non-ionic substances. The solved examples illustrate some of these points.

Solved Examples

Example 1: The percentage composition of a certain gas is: 42.9% carbon and 57.1% oxygen. 1.0 L of this gas weighs 1.25 g. What is the molecular formula for this gas?

Solution: From the percentage composition of the gas, we find its simplest formula is CO. Its molecular weight can be found by:

$$1.25\ \frac{g}{\cancel{L}} \times 22.4\ \frac{\cancel{L}}{mol} = 28.0\ \frac{g}{mol}$$

Since the true molecular weight is the same as the molecular weight of the simplest formula, the simplest formula is also the correct molecular formula.

Example 2: The simplest formula of a compound is C_2H_3O. 86 g of this compound dissolved in 1000 g of water has a boiling point of 100.52° C. What is the compound's true formula?

Solution: From the boiling point data, the molecular weight of the compound is 86. Since this is twice the value given by the simplest formula ($2 \times 12 + 3 \times 1 + 1 \times 16 = 43$), the correct formula must be $C_4H_6O_2$.

NAME________________________ DATE________________

Molecular (True) Formula

Practice Problems (Level 1)

1. A compound has the following percentage composition: 26.7% carbon; 2.2% hydrogen; 71.1% oxygen. The molecular weight of this compound is 90. What is the compound's true formula?

2. A certain compound was analyzed and found to have the following composition: 54.6% carbon; 9.0% hydrogen; 36.4% oxygen. The true molecular weight for the compound is 176. What is the molecular formula of the compound?

3. The percentage composition of ethane gas is 80.0% carbon and 20.0% hydrogen. The molecular weight for ethane is 30. What is the correct formula for this compound?

4. Analysis of a compound shows that it consists of 24.3% carbon, 4.1% hydrogen, and 71.6% chlorine. The molecular weight of the compound is determined to be 89.8. What molecular formula corresponds to these data?

5. An unknown compound is analyzed and found to consist of 49.0% carbon, 2.7% hydrogen, and 48.2% chlorine. Boiling point data suggest that the molecular weight of the compound is about 150. What molecular formula would you predict for this compound?

6. A gaseous compound is found to have the following composition: 30.5% nitrogen and 69.5% oxygen. The molecular weight of the gas is found to be 91.8. What molecular formula corresponds to these data?

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Molecular (True) Formula

Practice Problems (Level 2)

1. Chemical analysis of a gaseous compound shows its composition to be 36.4% carbon, 57.5% fluorine, and 6.1% hydrogen. A sample of 1.00 L of this gas weighs 2.96 g. What molecular formula do these data suggest for this compound?

2. A gaseous compound consists of 47.3% sulfur and 52.7% chlorine. 500. mL of this gas weighs 3.00 g. From these data, calculate the molecular formula for this gas.

3. One of the oxides of carbon has a composition of 42.8% carbon and 57.2% oxygen. 250. mL of this gas weighs 0.313 g. Which oxide of carbon is this, and what is the compound's true molecular formula?

4. Analysis of an organic compound indicates that it has a percentage composition as follows: 40.7% carbon; 5.0% hydrogen; 54.3% oxygen. When this compound is vaporized, 35.0 mL of the vapor weighs 0.184 g. What molecular formula do you predict for this compound?

5. An organic compound with the following composition is dissolved in water: 51.5% carbon; 8.7% hydrogen; 39.8% nitrogen. When 35.00 g of this compound is dissolved in 250 g of water, the boiling point of the resulting solution is 100.52° C. What molecular formula does this suggest for the compound?

6. A sample of an organic compound weighing 4.585 g is analyzed and found to consist of 3.056 g of carbon, 0.257 g of hydrogen, 0.677 g of oxygen, and 0.595 g of nitrogen. When 86.4 g of this compound is dissolved in 400 g of water, the resulting freezing point is −3.72° C. What molecular formula do these data suggest for the compound?

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Molecular (True) Formula

Practice Problems (Level 3)

1. A sample of gas is analyzed and found to consist of 46.0% carbon and 53.9% nitrogen. The density of the gas at 708 torr and 20.0° C is 2.02 g/L. What molecular formula do you predict for this gas?

2. A sample of an unknown compound weighing 3.58 g is analyzed and found to consist of 3.18 g of carbon and 0.40 g of hydrogen. When 4.50 g of this compound is dissolved in 100.0 g of water, the freezing point of the resulting solution is −0.620° C. What molecular formula do these data suggest for this compound?

3. A monosaccharide of unknown composition is analyzed and found to consist of 40.2% carbon, 53.3% oxygen, and 6.6% hydrogen. A solution formed by dissolving 6.67 g of this compound in 150. g of water has a freezing point of −0.45° C. What formula do you predict for this compound?

4. Analysis of a gaseous organic compound shows it to consist of 81.9% carbon and 18.2% hydrogen. A sample of 65.0 mL of the gas is found to weigh 0.296 g at STP. What is the correct molecular formula for this compound?

5. An unidentified organic compound has a composition of 40.0% carbon, 6.7% hydrogen, and 53.2% oxygen. When 28.6 g of this compound is dissolved in 245 g of water, the solution has a boiling point of 100.33° C. What molecular formula corresponds to these data?

6. A sample of an unidentified compound weighing 2.9147 g is analyzed and found to consist of 2.458 g of carbon and 0.459 g of hydrogen. At 745 torr and 35.0° C, 287.7 mL of this gas weighs 1.272 g. What is the correct molecular formula for this gas?

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Molecular (True) Formula

Teacher Notes

1. Explanation of Levels

Level 1: Simplest formulas must be calculated; molecular weight given for all compounds.

Level 2: Simplest formulas must be calculated; molecular weights must be calculated, but all calculations are as simple as possible.

Level 3: Simplest formulas must be calculated; molecular weights must be calculated, and all calculations represent some degree of difficulty.

2. Answers

Level 1:
1. $C_2H_2O_4$
2. $C_8H_{16}O_4$
3. C_2H_6
4. $C_2H_4Cl_2$
5. $C_6H_4Cl_2$
6. N_2O_4

Level 2:
1. $C_2F_2H_4$
2. S_2Cl_2
3. CO
4. $C_4H_6O_4$
5. $C_6H_{12}N_4$
6. C_6H_6NO

Level 3:
1. C_2N_2
2. $C_{10}H_{15}$
3. $C_6H_{12}O_6$
4. C_9H_{24}
5. $C_6H_{12}O_6$
6. C_8H_{18}

Gas Law Problems

The **volume** of a gas is a function of both **temperature** and **pressure.** The relationships among volume, temperature, and pressure for gases are expressed by Boyle's Law (volume and pressure), Charles' Law (volume and temperature), the combined gas law (volume, pressure, and temperature), and the ideal gas law (volume, number of moles, pressure, and temperature).

Two other laws express important relationships among gases. Graham's Law describes the relative rates of diffusion between two gases of different densities. And Dalton's Law of partial pressures describes the pressures exerted by each of the gases present in a mixture.

Solved Examples

Example 1: The volume of a dry gas at 738 torr is 45.0 mL. What will be the volume of this gas at standard pressure?

Solution: Standard pressure is 760. torr. A change in pressure from 738 torr to 760. torr is an increase by a factor of $\frac{760.}{738}$. This change in pressure will cause a comparable **decrease** in the same ratio. Therefore, the new volume of gas will be

$$45.0 \text{ mL} \times \frac{738 \text{ torr}}{760. \text{ torr}} = 43.7 \text{ mL}$$

Example 2: What volume will 1.50 g of carbon dioxide occupy at a temperature of 27.0°C and a pressure of 1.077 atm?

Solution: This problem can be solved by using the ideal gas law: pv = nRT. All variables are given or can be found except for v. Solving the ideal gas law for v, then, gives:

$$v = \frac{nRT}{p}$$

We can now solve for v if we (1) are sure to express all quantities in proper units, and (2) recall that n (number of moles) = w (weight in g) ÷ m (gram molecular weight).

$$v\,(L) = \frac{\dfrac{1.50\ g}{44.0\ g/mol} \times 0.0821\ \dfrac{L\ atm}{mol\ K} \times 300\ K}{1.077\ atm}$$

$$= 0.780\ L$$

Example 3: The relative rate of diffusion of two gases is 1.5. The molecular weight of the more dense gas is 36.45. What is the molecular weight of the other gas?

Solution: According to Graham's Law:

$$\frac{r_1}{r_2} = \frac{\sqrt{mw_2}}{\sqrt{mw_1}} \text{ or, in this case, } 1.5 = \frac{\sqrt{36.45}}{\sqrt{x}}$$

$$\text{or } \sqrt{x} = \frac{\sqrt{36.45}}{1.5}$$

$$\text{or } x = 16.2$$

NAME ______________________ DATE ______________

Gas Law Problems

Practice Problems (Level 1)

1. A dry gas occupies a volume of 28.4 mL at 725 torr. What will be the volume of this gas at 800. torr?

2. A dry gas with a volume of 588.8 mL at a pressure of 1.00049 atm is subjected to a new pressure of 1.035 atm. What is its volume under the new pressure?

3. A dry gas occupies a volume of 35. 9 mL at a temperature of 22.0° C. What volume will the same gas occupy at a temperature of 28.0° C?

4. At a temperature of 24.46° C, a dry gas occupies a volume of 4.588 mL. What volume will the gas occupy at a temperature of 21.24° C?

5. At a pressure of 780. mm and 24.2° C, a certain gas has a volume of 350.0 mL. What will be the volume of this gas under standard conditions?

6. A dry gas at a temperature of 67.5° C and a pressure of 882 torr occupies a volume of 242.2 mL. What will be the volume of the gas at a new pressure of 840. torr and 80.0° C?

7. A sample of gas containing 0.089 mol is put into a 10.00 L container at a temperature of 30.0° C. What pressure does the gas exert on the container?

8. How many moles of gas are contained in a 50.0 L cylinder at a pressure of 100.0 atm and a temperature of 35° C?

9. What is the relative rate of diffusion between two samples of oxygen and hydrogen gas?

10. A sample of gas consists of 75% hydrogen and 25% oxygen. The total pressure exerted by the gas is 788 torr. What pressure is exerted by each of the gases individually?

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Gas Law Problems

Practice Problems (Level 2)

1. A gas occupies a volume of 34.2 mL at a temperature of 15.0°C and a pressure of 800.0 torr. What will be the volume of this gas at standard conditions?

2. At conditions of 785 torr of pressure and 15.0°C temperature, a gas occupies a volume of 45.5 mL. What will be the volume of the same gas at 745 torr and 30.0°C?

3. A dry gas has a volume of 100.0 mL at a pressure of 1600. torr. At what pressure would this volume be reduced to 50.0 mL?

4. A dry gas at a temperature of 18.0°C has a volume of 40.0 mL. What temperature change is needed to reduce this volume to 35.0 mL?

5. 40.0 mL of gas is collected over water on a day when the barometric pressure was 790.0 torr and the temperature 20.0°C. What would be the volume of this (dry) gas at standard conditions?

6. A sample of oxygen collected over water when the atmospheric pressure was 1.002 atm and the room temperature, 25.5°C occupied 105.8 mL. What would be the volume of this dry gas at standard conditions?

7. What is the relative rate of diffusion of two gases whose densities are 1.33×10^{-3} and 1.165×10^{-3} g/cm^3?

8. Find the density of a gas which diffuses at a rate 1.16 times greater than that of one with a density of 1.78×10^{-3} g/cm^3.

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Gas Law Problems

Practice Problems (Level 3)

1. A gas collected on a day when the atmospheric pressure was 1.12 atm had a volume of 252.4 mL. A day later, that volume had changed to 248.8 mL. What was the atmospheric pressure on the second day?
2. The volume of a dry gas originally at standard temperature and pressure was recorded as 488.8 mL. What volume would the same gas occupy when subjected to a pressure of 100. atm and temperature of -245° C?
3. Calculate the density of carbon monoxide gas at standard conditions if 25.0 mL of the gas weighs 0.0329 g at 800.00 mm of pressure.
4. The weight of 15.0 mL of gas at 5.00 atm of pressure and a temperature of -53.0° C is 8.35×10^{-3} g. What is the density of this gas under standard conditions?
5. A sample of gas occupying 5.000 mL is collected over water at a pressure of 1.35 atm and a temperature of 10.5° C. What volume would the gas occupy at standard conditions?
6. What weight of oxygen gas is contained in a 3.50 L tank where the temperature is 50.0° C and the pressure maintained at 4.5 atm?
7. Calculate the molecular weight of a gas if 35.44 g of the gas stored in a 7.50 L tank exerts a pressure of 60.0 atm at a constant temperature of 35.5° C.
8. What pressure will be exerted by each of the gases in the following mixture if the total pressure of the mixture amounts to 768.8 torr? 0.500 g of hydrogen; 0.245 g of oxygen; 0.335 g of nitrogen.
9. A mixture of gases containing equal weights of carbon dioxide, carbon monoxide, and ammonia exerts an overall pressure of 450.0 torr. What pressure is exerted by each gas individually?
10. The relative rate of diffusion between two gases is 1.89. If the lighter gas is methane (mw = 16), what is the molecular weight of the other gas?

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Gas Law Problems

Teacher Notes

1. Explanation of Levels

Level 1: Problems involving pressure **or** temperature changes only; calculation of volume changes only; combined gas law problems with dry gases only; simple ideal gas law and Graham's Law problems.

Level 2: Combined gas law problems for dry gases; determination of pressure and temperature changes; simple problems involving water vapor pressure; problems on Graham's Law.

Level 3: Determination of pressure and temperature changes; more difficult combined gas law problems; effects of temperature and pressure change on density; more difficult Graham's Law and Dalton's Law problems.

2. Answers

Level 1: 1. 25.8 mL; 2. 568.9 mL; 3. 36.6 mL; 4. 4.538 mL; 5. 330 mL; 6. 264 mL; 7. 0.22 atm; 8. 198 mol; 9. 0.25; 10. 591 torr H_2; 197 torr O_2

Level 2: 1. 34.1 mL; 2. 50.4 mL; 3. 3200 torr; 4. −18.4°C; 5. 37.9 mL; 6. 94.2 mL; 7. 0.934; 8. 1.32×10^{-3} g/cm^3

Level 3: 1. 1.14 atm; 2. 0.50 mL; 3. 1.25 g/L; 4. .0897 g/mL; 5. 6.47 mL; 6. 19 g; 7. 2.00 g/mol; 8. 711 torr hydrogen; 23. torr oxygen; 35 torr nitrogen; 9. 90 torr carbon dioxide; 141 torr carbon monoxide; 219 torr ammonia; 10. 57.3

Density and Specific Gravity

The **density** of a substance is defined as its mass per unit volume. The "unit volume" used in calculating densities may be **one** liter, **one** milliliter, **one** cubic centimeter, **one** cubic foot (in the English system), or **one** of any other convenient volume. Mathematically, density is defined as:

$$D = \frac{m}{v}$$

Notice that if any two of these variables (density, mass, volume) are known, the third can be calculated.

Specific gravity is a term now falling into disuse. The specific gravity of a substance is defined as the density of that substance compared to the density of some standard, or:

$$SG = \frac{\text{density of substance}}{\text{density of standard}}$$

For solids and liquids, the standard is water. Since the density of water is 1.00 g/cm^3, the numerical value of the specific gravity of a solid or liquid is identical to its density. Specific gravity has no dimensions. The standard for gases is usually air, with a density of 1.29 g/L.

Solved Examples

Example 1: Calculate the density and specific gravity of a liquid if 38.94 cm^3 of that liquid has a mass of 34.8 g.

Solution:

$$\text{density} = \frac{\text{mass of liquid}}{\text{volume of liquid}}$$

$$= \frac{34.8\ \text{g}}{38.94\ \text{cm}^3}$$

$$= 0.894\ \text{g/cm}^3$$

$$\text{specific gravity} = \frac{\text{density of liquid}}{\text{density of water}}$$

$$= \frac{0.894\ \text{g/cm}^3}{1.000\ \text{g/cm}^3} = 0.894$$

Example 2: What volume of ethyl alcohol has a weight of 125.0 g? The specific gravity of ethyl alcohol is 0.791.

Solution: To find the density of the ethyl alcohol:

$$SG = \frac{D_{\text{substance}}}{D_{\text{standard}}} \quad \text{or, } D_{\text{substance}} = SG \times D_{\text{standard}}$$

So, $D_{\text{substance}} = 0.791 \times 1.000\ \text{g/cm}^3 = 0.791\ \text{g/cm}^3$

Then, to find the volume of the ethyl alcohol:

$$D = \frac{m}{v} \qquad \text{or} \qquad v = \frac{m}{D}$$

So, $v = \dfrac{125.0\ \text{g}}{0.791\ \text{g/cm}^3} = 158.0\ \text{cm}^3$

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Density and Specific Gravity

Practice Problems (Level 1)

1. A sample of seawater weighs 158 g and has a volume of 156 mL. What are the density and specific gravity of this seawater?
2. A block of aluminum occupies a volume of 15.0 mL and weighs 40.5 g. What are the density and specific gravity of this metal?
3. A cylindrical box with a volume of 200. cm^3 holds 432.0 g of sodium chloride. From this information, calculate the density and specific gravity of sodium chloride. (Ignore the space between sodium chloride crystals.)
4. If the salt is emptied out of the box described in problem 3 and sugar is used to fill the box, the weight of sugar used is equal to 316.0 g. What are the density and specific gravity of sugar? (Ignore the space between sugar crystals.)
5. Mercury metal is poured into a graduated cylinder that holds exactly 22.5 mL. The mercury used to fill the cylinder weighs 306.0 g. From this information, calculate the density and specific gravity of mercury.
6. What is the weight of the ethyl alcohol that exactly fills a 200.0 mL container? The density of ethyl alcohol is 0.789 g/mL.
7. Calculate the density and specific gravity of helium from the fact that a balloon with a capacity of 5.00 L holds 0.890 g. (Use air as the standard for calculating specific gravity.)
8. A flask built to hold exactly 2.5000 L is filled with nitrogen. The weight of the nitrogen in the flask at standard conditions is 3.1250 g. Calculate the density and specific gravity (air standard) for nitrogen.
9. A flask that weighs 345.8 g is filled with 225 mL of carbon tetrachloride. The weight of the flask and carbon tetrachloride is found to be 703.55 g. From this information, calculate the density and specific gravity of carbon tetrachloride.
10. A rubber balloon weighing 144.85 g is filled with carbon dioxide gas and reweighed. The weight of the balloon plus gas is 153.77. The volume of the balloon filled with carbon dioxide is 4.55 L. What density and specific gravity (air standard) do these data yield for carbon dioxide?

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Density and Specific Gravity

Practice Problems (Level 2)

1. Calculate the density and specific gravity of sulfuric acid from the information that 35.4 mL of the acid weighs 65.14 g.

2. Calculate the density and specific gravity of ammonia gas from the information that 8.9 L of the gas weighs 6.86 g.

3. A block of lead has dimensions of 4.5 cm by 5.2 cm by 6.0 cm. The block weighs 1587 g. From this information, calculate the density and specific gravity of lead.

4. 28.5 g of iron shot is added to a graduate flask containing 45.5 mL of water. The water level rises to the 49.1 mL mark. From this information, calculate the density and specific gravity of iron.

5. A cylindrical glass tube of length 27.75 cm and radius 2.00 cm is filled with argon gas. The empty tube weighs 188.25 g. The tube filled with argon weighs 188.87 g. Use these data to calculate the density of argon gas.

6. Find the weight of 250.0 mL of benzene. The density of benzene is 0.90. g/mL.

7. What volume of silver metal will weigh exactly 2500.0 g? The specific gravity of silver is 10.5.

8. What is the weight of 215 L of hydrogen sulfide gas if the density of hydrogen sulfide is 1.54 g/L?

9. The hydrogen gas stored inside a large weather balloon weighs 13.558 g. What is the volume of this balloon if the density of hydrogen is 0.089 g/L?

10. Calculate the weight of 2250.0 L of ammonia gas if the specific gravity of ammonia is 0.597.

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Density and Specific Gravity

Practice Problems (Level 3)

1. From information on the molar volume of gases, calculate the theoretical density of chlorine gas.

2. A rectangular block of copper metal weighs 1896 g. The dimensions of the block are 8.4 cm by 4.6 cm by 5.5 cm. From these data, what is the specific gravity of copper?

3. The specific gravity of gold is 19.3. What size block of gold would have a mass of exactly 1.0000 kg?

4. A tiny droplet of mercury metal is measured to have a radius of 2.3×10^{-4} cm. What would be the weight of mercury contained in this droplet? The specific gravity of mercury is 13.6.

5. A droplet of liquid iron is held suspended by a magnetic field. The weight of the iron is determined to be 4.52×10^{-4} g. What is the radius of the droplet? The density of iron is 7.86 g/mL.

6. A piece of aluminum foil 12.8 cm long and 4.2 cm wide is found to weigh 0.319 g. What is the thickness of the foil if the density of aluminum is 2.70 g/cm^3?

7. What is the weight of the helium contained in a flask whose volume is 4.5000 mL? The specific gravity of helium is 0.01395.

8. A drop of oil weighing 0.0025 g is dropped on the surface of a large pool of water. The drop spreads out until it forms a circle 18.2 m in diameter. How thick is the circle of oil? The density of this oil is 0.725 g/cm^3.

9. Ethyl alcohol is added to a beaker weighing 204.88 g until the beaker and alcohol together weigh 253.12 g. What volume of ethyl alcohol was added to the beaker? The density of ethyl alcohol is 0.789 g/cm^3.

10. What molecular weight do you calculate for hydrogen fluoride if the specific gravity of this gas is 1.384?

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Density and Specific Gravity

Teacher Notes

1. Explanation of Levels

Level 1: Calculation of density and specific gravity from simple weight and volume calculations; two problems on finding weight of given volume with density known.

Level 2: Calculations of density and specific gravity when volume is not given directly; calculations of weights and volumes when density or specific gravity is known.

Level 3: More complex calculations of density and specific gravity; calculations of one dimension of a geometric figure when weight and density or specific gravity are known.

2. Answers

Level 1: 1. 1.01 g/mL; 1.01; 2. 2.70 g/mL; 2.70; 3. 2.16 g/mL; 2.16; 4. 1.58 g/mL; 1.58; 5. 13.6 g/mL; 13.6; 6. 157.8 g; 7. 0.178 g/L; 0.138; 8. 1.25 g/L; 0.969; 9. 1.59 g/L; 1.59; 10. 1.96 g/L; 1.52

Level 2: 1. 1.84 g/mL; 1.84; 2. 0.77 g/L; 0.60; 3. 11 g/mL; 11; 4. 7.86 g/mL; 7.86; 5. 1.78 g/L; 6. 225 g; 7. 238 mL; 8. 331 g; 9. 152 L; 10. 1730 g

Level 3: 1. 3.165 g/L; 2. 8.9; 3. 51.8 mL; 4. 6.9×10^{-10} g; 5. 2.39×10^{-2} cm; 6. 0.0022 cm; 7. 8.1×10^{-4} g; 8. 1.33×10^{-9} cm; 9. 61.1 cm^3; 10. 40.00

Energy

Every chemical change involves the loss or gain of energy. Most physical changes also give off or absorb energy. In most cases, this energy occurs in the form of heat. The amount of heat energy gained or lost in a chemical or physical change is measured in units of **calories**. A calorie is defined as the amount of heat needed to raise the temperature of one gram of water by one degree Celsius. Another unit used to measure energy is the **joule**. A joule is equivalent to 0.239 calorie.

Temperature is the condition of a body which determines the transfer of heat to or from other bodies. It is an indication of the average kinetic energy of the particles of which that body is made.

Heat energy applied to a body may produce one of two effects in that body: (1) it may raise the temperature of that body, or (2) it may bring about a change of state of that body. The amount of heat needed to change a substance from a solid to a liquid is called its **heat of fusion**. The amount of heat needed to change a substance from a liquid to a gas is known as its **heat of vaporization**.

The heat energy absorbed or released during a chemical change is known as the **heat of reaction**. The heat of reaction is usually expressed in units of kilocalories per mole.

Solved Examples

Example 1: Change a temperature reading of 36.5°F to its equivalent in degrees Celsius and degrees Kelvin.

Solution:

$$
\begin{aligned}
°C &= 5/9\,(°F - 32) \\
&= 5/9\,(36.5°F - 32) \\
&= 2.5°C
\end{aligned}
$$

$$K = {}^\circ C + 273$$
$$= 2.5^\circ C + 273 = 275.5\ K$$

Example 2: How much heat (H) is released when 52.5 g of water cools from 67.5°C to 23.2°C? The specific heat (C_p) of water is 1.00 cal/°C · g.

Solution: $H = C_p \times \text{mass of water} \times \text{temperature change}$

$$= 1.00\ \frac{\text{cal}}{^\circ\cancel{\text{C}}\cdot\cancel{\text{g}}} \times 52.5\,\cancel{\text{g}} \times 44.3^\circ\cancel{\text{C}}\,(67.5^\circ\text{C} - 23.2^\circ\text{C})$$

$$= 2330\ \text{cal}$$

Example 3: What mass (M) of aluminum can be melted by the addition of 250. cal of heat? The heat of fusion (H_f) of aluminum is 94.5 cal/g.

Solution:

$$H = H_f \times M \qquad \text{or} \qquad M = \frac{H}{H_f}$$

$$W = \frac{250.\,\cancel{\text{cal}}}{94.5\,\cancel{\text{cal}}/\text{g}}$$

$$= 2.65\ \text{g}$$

NAME ______________________ DATE ______________

Energy

Practice Problems (Level 1)

1. Each of the following is a temperature reading in one of three systems: Fahrenheit, Celsius, or Kelvin. Change the reading to its equivalent in both of the other scales.

   a. 70.°F
   b. 25°C
   c. 100.°F
   d. 373 K
   e. 85°C
   f. 215°C
   g. 90.°F
   h. 285 K
   i. 35°C
   j. 305 K

2. Calculate the number of calories of heat absorbed or released in each of the following changes.

   a. 40.0 g of water at 25.0°C raised to 60.0°C.
   b. 125 g of water at 10.0°C raised to 90.0°C.
   c. 75.0 g of water at 9.8°C raised to 22.4°C.
   d. 44.8 g of iron at 80.5°C cooled to 62.6°C. The specific heat of iron is 0.11 cal/°C · g.
   e. 64.82 g of aluminum metal at 100.0°C cooled to 82.5°C. The specific heat of aluminum metal is 0.215 cal/°C · g.

3. Calculate the amount of heat given off or taken on during each of the following changes:

   a. the melting of 25.0 g of iron; the heat of fusion of iron is 63.7 cal/g.
   b. the boiling of 125 g of antimony; the heat of vaporization of antimony is 380 cal/g.
   c. the melting of 235 g of bismuth; the heat of fusion of bismuth is 12.4 cal/g.
   d. the boiling of 350 g of chromium; the heat of vaporization of chromium is 1560 cal/g.

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Energy

Practice Problems (Level 2)

1. Each of the following is a temperature reading in one of three systems: Fahrenheit, Celsius, or Kelvin. Change the reading to its equivalent in both of the other scales.

   a. 215°C
   b. 6.0°F
   c. 145 K
   d. -95°C
   e. 2120°F
   f. -325°F
   g. 25.6°C
   h. 35 K
   i. 3565°C
   j. -166°F

2. Calculate the number of calories of heat lost or gained during each of the following changes:

   a. 114.32 g of water at 14.85°C raised to 18°C.
   b. 132 g of copper at 32.2°C raised to 45.0°C; the specific heat of copper is 0.092 g/°C · g.
   c. 24.5 g of ice at -10.0°C warmed to 42.5°C. The specific heat of ice is 0.50 cal/°C · g. The heat of fusion for ice is 80 cal/g.
   d. 354 g of ice at -40.0°C heated to steam at 112°C. The specific heats of ice and steam are 0.50 cal/°C · g. The heat of vaporization for water is 540 cal/g.

3. What is the final temperature of 250.0 g of water whose initial temperature is 25.0°C if 80.0 g of aluminum initially at 70.0°C is dropped into the water? The specific heat of aluminum is 0.215 cal/°C · g.

4. Calculate the specific heat of mercury metal if 250. g of the metal, warmed from 20.2°C to 37.4°C gives of 145 cal of heat.

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Energy

Practice Problems (Level 3)

1. A sample of mercury metal is heated from 25.5°C to 52.5°C. In the process, 187 cal of heat are absorbed. What mass of mercury was in the sample? The specific heat of mercury is 0.033 cal/°C · g.
2. A block of aluminum weighing 140 g is cooled from 98.4°C to 62.2°C with the release of 1080 cal of heat. From these data, calculate the specific heat of aluminum.
3. A cube of gold weighing 192.4 g is heated from 30.0°C to some higher temperature, with the absorption of 226 cal of heat. The specific heat of gold is 0.030 cal/°C · g. What was the final temperature of the gold?
4. A total of 54 cal of heat are absorbed as 58.3 g of lead is heated from 12.0°C to 42.0°C. From these data, what is the specific heat of lead?
5. A piece of erbium metal weighing 100.0 g and heated to 95.0°C is dropped into 200.0 g of water initially at 20.0°C. The final temperature of the mixture is 21.5°C. What is the specific heat of erbium metal?
6. A block of rhenium metal (specific heat = 0.0329 cal/g · °C) is heated to 88.2°C and then dropped into 100.0 g of water initially at 26.4°C. The final temperature of the mixture is 32.4°C. What was the mass of the block of rhenium?
7. When 258.6 g of benzene vapor is condensed to a liquid at its boiling point, 33,875 cal of heat are released. What is the heat of vaporization for benzene?
8. A sample of ethyl alcohol is converted from a liquid to a vapor with no temperature change. In the process, 30,640 cal of heat are absorbed. What mass of ethyl alcohol was in the sample? The heat of vaporization of ethyl alcohol is 210. cal/g.
9. The heat of combustion of methane is 212.8 kcal per mole. How much heat will be produced in the combustion of 100.0 g of methane?
10. The heat of combustion of toluene is 934.2 kcal per mole. How much heat will be released during the combustion of 250.0 g of toluene? The formula for toluene is $C_6H_5CH_3$.

Energy

Teacher Notes

1. Explanation of Levels

Level 1: Conversion of temperature readings at simple level; calculation of heat gain and loss for simple cases; calculation of heat gain and loss during melting and boiling.

Level 2: Conversion of temperature readings at more advanced levels (decimal, very large, and negative values); calculation of heat change, specific heat, and temperature for a variety of situations.

Level 3: Calculation of heat change, specific heat, weight, and temperature for complex changes; two simple problems on heat of combustion.

2. Answers

Level 1:

1. a. 21°C; 294°C; b. 77°F; 298 K; c. 37.8°C; 311.0 K; d. 100°C; 212°F; e. 185°F; 358 K; f. 419°F; 488 K; g. 32°C; 305 K; h. 12°C; 54°F; i. 95°F; 308 K; j. 32°C; 90°F
2. a. $14\bar{0}0$ cal; b. $10{,}\bar{0}00$ cal; c. 940 cal; d. −88 cal
3. a. 1,590 cal; b. 47,500 cal; c. 2,910 cal; d. 546,000 cal

Level 2:

1. a. 419°F; 488 K; b. −14°C; 259 K; c. −128°C; −198°F; d. −139°F; 178 K; e. 1160°C; 1433 K; f. −198°C; 75 K; g. 78.1°F; 298.8 K; h. −238°C; −396°F; i. 6449°F; 3838 K; j. −110°C; 163 K
2. a. 360 cal; b. 155 cal; c. 3120 cal; d. 264,000 cal
3. 27.9°C
4. 0.0337 cal/°C · g

Level 3: 1. 210 g; 2. 0.213 cal/°C · g; 3. 69°C; 4. 0.031 cal/°C · g; 5. .0408 cal/°C · g; 6. 327 g; 7. 131.0 cal/g; 8. 146 g; 9. 1326 kcal; 10. 2534 kcal