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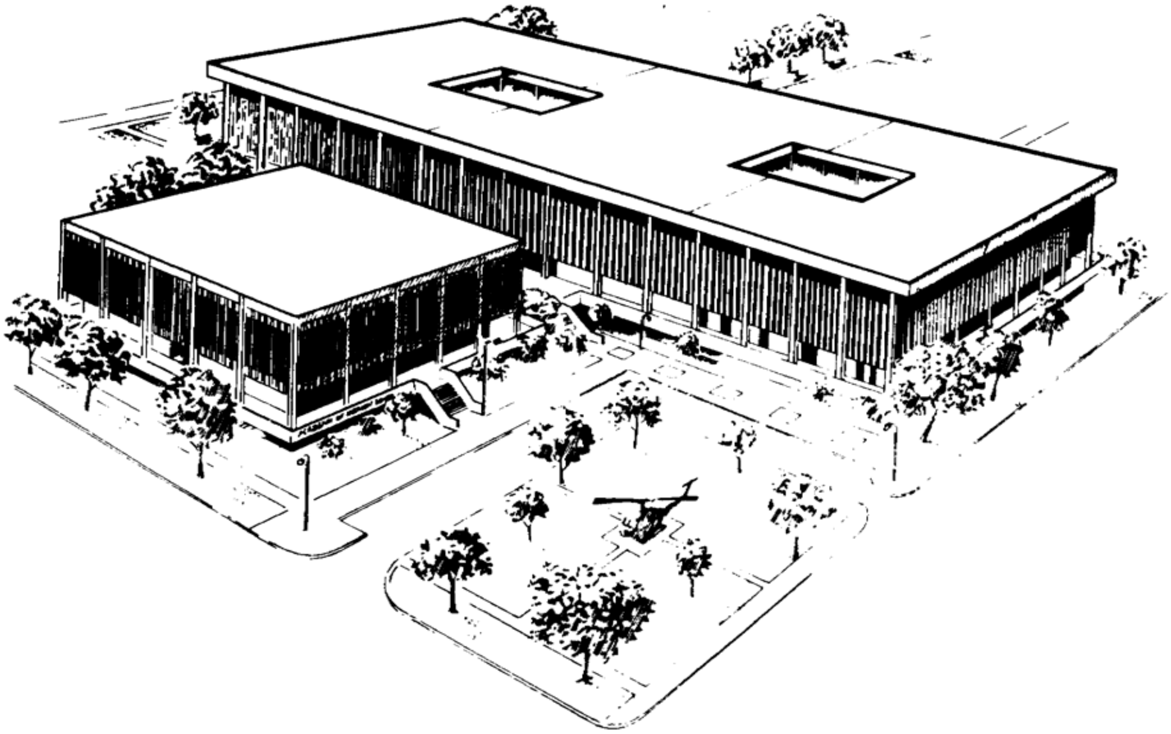

LABORATORY MATHEMATICS

SUBCOURSE MD0837 EDITION 100

DEVELOPMENT

This subcourse is approved for resident and correspondence course instruction. It reflects the current thought of the Academy of Health Sciences and conforms to printed Department of the Army doctrine as closely as currently possible. Development and progress render such doctrine continuously subject to change.

ADMINISTRATION

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Approved students whose enrollments remain in good standing may apply to the Nonresident Instruction Branch for subsequent courses by telephone, letter, or e-mail.

Be sure your social security number is on all correspondence sent to the Academy of Health Sciences.

CLARIFICATION OF TRAINING LITERATURE TERMINOLOGY

When used in this publication, words such as "he," "him," "his," and "men" are intended to include both the masculine and feminine genders, unless specifically stated otherwise or when obvious in context.

.

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CORRESPONDENCE COURSE OF THE
U.S. ARMY MEDICAL DEPARTMENT CENTER AND SCHOOL

SUBCOURSE MD0837
LABORATORY MATHEMATICS

INTRODUCTION

This subcourse was developed to prepare you and sustain you in mathematical skills for the Advanced Medical Laboratory (AML) Course (92B30). The emphasis is upon computations related to solutions and their concentrations. If you feel that you need a more basic review of mathematics before taking this subcourse, you should request Subcourse MD0900, Basic Mathematics, which covers addition, subtraction, multiplication, and division of whole numbers; decimals and fractions; and conversions to and from the metric system. If you need a basic review of chemistry, you should request Subcourse MD0903, General Chemistry.

You must be able to solve the types of problems included in the lesson exercises and the examination if you are to perform well in the AML course and later as an Advanced Medical Laboratory Specialist.

Subcourse Components:

This subcourse consist of nine lessons and Appendices. The lessons and appendices are:

Lesson 1. General Mathematics Review.

Lesson 2. Introduction to Solution Mathematics.

Lesson 3. Molar Solutions.

Lesson 4. Equivalent Solutions.

Lesson 5. Conversion of Concentration Units.

Lesson 6. Dilutions.

Lesson 7. Titration.

Lesson 8. Concentrated Acids and Bases.

Lesson 9. pH and Buffers

APPENDIX A. Review of Dimensional Analysis

Credit Awarded:

To receive credit hours, you must be officially enrolled and complete an examination furnished by the Nonresident Instruction Branch at Fort Sam Houston, Texas. Upon successful completion of the examination for this subcourse, you will be awarded 13 credit hours.

You can enroll by going to the web site http://atrrs.army.mil and enrolling under "Self Development" (School Code 555).

A listing of correspondence courses and subcourses available through the Nonresident Instruction Section is found in Chapter 4 of DA Pamphlet 350-59, Army Correspondence Course Program Catalog. The DA PAM is available at the following website: http://www.usapa.army.mil/pdffiles/p350-59.pdf.

LESSON ASSIGNMENT

LESSON 1	General Mathematics Review.
TEXT ASSIGNMENT	Paragraphs 1-1 through 1-49.
LESSON OBJECTIVES	After completing this lesson, you should be able to:

1-1. Identify and apply the properties of real numbers when performing computations.

1-2. Use properties of equality to solve equations for unknowns.

1-3. Round off numbers according to instructions in this lesson.

1-4. Express numbers in standard scientific notation and perform arithmetic calculations for same.

1-5. Determine the number of significant figures in a number and appropriately report these figures for computations.

1-6. Find the logarithm for a given number, find the antilogarithm for a given logarithm, and calculate using logarithms.

SUGGESTION After studying each section, complete the exercises at the end of each section. These exercises help you to achieve the lesson objectives.

LESSON 1

GENERAL MATHEMATICS REVIEW

Section I. PROPERTIES OF REAL NUMBERS

1-1. DISCUSSION

Real numbers can be described as being the set of all x, such that x is the coordinate of a point on a number line. The real numbers include the rational and irrational numbers. Rational numbers include integers and quotients of integers. Integers are whole numbers. Positive integers, such as 1, 2, 3, and so forth, are also called natural numbers. An operation is a method of combining two elements of the set of real numbers to form a third element (not exclusive). The properties of real numbers depend upon two basic operations, addition and multiplication. This section is designed to give you the concepts and skills necessary for the content and course work to follow.

1-2. SOME FUNDAMENTAL PROPERTIES OF REAL NUMBERS

For any real numbers a, b, c, and d, the additive and multiplicative properties are as follows:

a. **Commutative Properties**.

$a + b = b + a$ $\qquad$ $ab = ba$

Examples:

$3 + 5 = 5 + 3$ $\qquad$ $2 \text{ X } 4 = 4 \text{ X } 2$

b. **Associative Properties**.

$(a + b) + c = a + (b + c)$ $\qquad$ $(ab)c = a(bc)$

Examples:

$(2 + 1) + 4 = 2 + (1 + 4)$ $\qquad$ $(4 \text{ X } 3) \text{ X } 2 = 4 \text{ X } (3 \text{ X } 2)$

c. **Identity Properties**.

$a + 0 = a$ $\qquad$ $a \text{ X } 1 = a$

Examples:

$3 + 0 = 3$ $\qquad$ $7 \text{ X } 1 = 7$

d. **Inverse Properties**.

$a + (-a) = 0$

For each nonzero real number a, there is a unique real number 1/a such that:

$$a \text{ X } \frac{1}{a} = 1$$

Examples:

$3 + (-3) = 0$

$$12 \text{ X } \frac{1}{12} = 1$$

e. **Distributive Property**.

$a (b + c) = ab + ac$

Example: $2(3 + 2) = (2 \text{ X } 3) + (2 \text{ X } 2)$

1-3. DEFINITIONS OF SUBTRACTION AND DIVISION

a. Subtraction is a special case of addition.

$a - b = a + (-b)$

Example: $5 - 2 = 5 + (-2)$

b. Division is a special case of multiplication.

$\frac{a}{b} = a \text{ X } \frac{1}{b}$ provided that b is not equal to 0

Example: $\frac{6}{3} = 6 \text{ X } \frac{1}{3}$.

NOTE: Division by 0 is undefined, because any real number divided by 0 has no solution.

1-4. PROPERTIES OF ZERO (0)

a. $a \times 0 = 0$

Example: $5 \times 0 = 0$

b. If $\frac{a}{b} = 0$, then $a = 0$

Example: $\frac{0}{3} = 0$

1-5. PROPERTIES OF NEGATIVE NUMBERS

a. $(-a)(-b) = ab$

Example: $(-3)(-4) = 12$

NOTE: When no arithmetic operator (+, -, ÷) is present between two real numbers, one or both being placed in parentheses, multiplication is implied.

b. $(-a)b = -(ab)$

Example: $(-2)3 = -6$

c. $-a = (-1)a$

Example: $-13 = (-1)13$

1-6. PROPERTIES OF FRACTIONS

a. $\frac{a}{b} = \frac{c.}{d}$ if and only if $ad = bc$

Example: $\frac{4}{2} = \frac{8}{4}$ since $(4)(4) = (2)(8)$

b. $\frac{a}{b} + \frac{c}{b} = \frac{a + c}{b}$

Example: $\frac{6}{2} + \frac{4}{2} = \frac{6 + 4}{2} = \frac{10}{2} = 5$

c. $\frac{a}{b} + \frac{c}{d} = \frac{ad + bc}{bd}$

Example: $\frac{4}{3} + \frac{1}{2} = \frac{(4)(2) + (3)(1)}{(3)(2)} = \frac{8+3}{6} = \frac{11}{6}$

d. $\frac{a}{b} \text{ X } \frac{c}{d} = \frac{ac}{bd}$

Example: $\frac{2}{3} \text{ X } \frac{6}{5} = \frac{12}{15}$

e. $\frac{a}{b}\ .\div \frac{c}{d} = \frac{a}{b} \text{ X } \frac{d}{c}$

Example: $\frac{2}{3} \div \frac{4}{5} = \frac{2}{3} \text{ X } \frac{5}{4} = \frac{10}{12}$

f. $-\frac{a}{b} = \frac{-a}{b} = \frac{a}{-b}$

Example: $-\frac{6}{3} = \frac{-6}{3} = \frac{6}{-3} = -2$

1-7. ORDER OF OPERATIONS

a. If parentheses or brackets are present, evaluate the expressions within first, starting with the innermost set of parentheses using the rules that follow.

b. Evaluate any exponential expressions.

c. Evaluate any multiplications or divisions from left to right.

d. Evaluate any additions or subtractions from left to right.

e. If an expression is in fractional form, perform stated operations in the numerator and denominator, then simplify if possible.

Examples:

$$\frac{3 + 10}{2 - 4} = \frac{13}{-2} \qquad \frac{15}{9 - 6} = \frac{15}{3} = 5$$

NOTE: A good way to remember the order of operations is to use the memory aid, "**P**lease **e**xcuse **m**y **d**ear **A**unt **S**ally." The order for solving algebra equations is: parenthesis, exponents, multiplication and division, addition and subtraction.

1-8. EXERCISES, SECTION I

After you have completed these exercises, turn to the end of the lesson and check your answers with the solutions.

FIRST REQUIREMENT: Identify the property of real numbers that justifies each of the following mathematical statements:

a. $2 + (p - 1) = (2 + p) - 1$

____

b. $4 + a = a + 4$

____

c. $\frac{1}{7 + e}(7 + e) = 1$

____

d. $3 \times 1 = 3$

____

e. $r(1 + 8) = r + 8r$

____

f. $-11 + 11 = 0$

g. $5 + 0 = 5$

h. $12 - 6 = (-6) + 12$

i. $3(2 - 5) = 6 - 15$

j. $v - v = 0$

SECOND REQUIREMENT: Evaluate the following expressions:

k. $-6(-3)$

l. $-12 - 3$

m. $\frac{15}{3} + 4 + 7$

n. $\frac{24}{8 - 6}$

o. $\frac{2}{5} - \frac{5}{3}$

____

p. 6(30y + 4)

____

q. $4 + 2 \times 4 - \frac{9}{3}$

____

r. 150 X 0

____

s. $\frac{2}{9} + \frac{5}{9}$

____

t. $\frac{2}{8} - \frac{4}{8}$

____

Section II. PROPERTIES OF EQUALITY

1-9. DISCUSSION

An equation is a statement that two expressions or quantities are equal in value. The statement of equality 2x + 3 = 11 is an equation. It is algebraic shorthand for "the sum of two times some number plus three is equal to eleven." Of course this algebraic shorthand makes problem solving much easier. Solving an equation is determining the values of the unknown numbers in that equation that makes the equation true. Various techniques may be employed in the solving of equations. The form of the equation and the degree (highest power a variable is raised in the equation) will determine the method utilized in problem solving. Most of the problem solving in the following course work will involve first-degree equations; that is equations whose variables are raised to a power of one, so the following methods will develop techniques to solve equations of this type.

NOTE: Some second-degree equations will be encountered in the pH and buffers problems. However, methods taught here will be sufficient for any problem solving.

1-10. PROPERTIES OF EQUALITY

For any real number a, b, or c, the properties are as follows.

a. **Reflexive Property**.

$a = a$

Example: $8 = 8$

b. **Symmetric Property**.

If $a = b$, then $b = a$

Example: If $x + 5 = y$, then $y = x + 5$

c. **Transitive Property**.

If $a = b$ and $b = c$, then $a = c$

Example: If $a + b = c$, and $c = 3a$, then $a + b = 3a$

d. **Substitution Property**.

If $k = x$, then k may replace x in any equation without changing the truth of the equation.

Example: If $2k - 4 = x$, and $x + k = 9$,
then $2k - 4 + k = 9$

1-11. RULES OF EQUALITY

a. Adding or subtracting the same quantity on both sides of an equation produces an equivalent equation.

b. Multiplying or dividing both sides of an equation by the same nonzero quantity produces an equivalent equation.

c. Simplifying an expression on either side of an equation produces an equivalent equation.

1-12. SOLVING FIRST DEGREE EQUATIONS

a. Eliminate any fractions by multiplication.

b. Simplify each side of the equation as much as possible by combining like terms.

c. Use the addition property to simplify the equation so that all terms with the desired variable are on one side of the equation and all numbers and variables other than the one being solved for are on the other side.

d. Use multiplication or division to get an equation with just the desired variable on one side.

e. Check your answer by substituting the solution back into the original equation and evaluating it for truth when practical.

1-13. PROBLEM—SOLVING TECHNIQUES

a. Solve for x:

$3[1 - 2(x + 1)] = 2 - x$	
$3[1 - 2x - 2] = 2 - x$	simplify
$3(-1 - 2x) = 2 - x$	
$-3 - 6x = 2 - x$	
$-3 - 5x = 2$	add x to both sides
$-5x = 5$	add 3 to both sides
$x = -1$	divide both sides by -5
Check the Solution.	
$3[1 - 2(-1 + 1)] = 2 - (-1)$	substitute -1 for all x
$3[1 - 2(0)] = 2 - (-1)$	evaluate
$3(1) = 2 + 1$	true
$3 = 3$	

b. Solve for x:

$$\frac{x}{2} - \frac{3x}{4} = 1$$

$$4 \left| \frac{x}{2} - \frac{3x}{4} \right| = (4)1$$ multiply both sides by the least common denominator

$2x - 3x = 4$ simplify

$-x = 4$

$x = -4$ multiply both sides by -1

Check the Solution.

$$\frac{(-4)}{2} - \frac{3(-4)}{4} = 1$$ substitute -4 for all x

$-2 + 3 = 1$ true

$1 = 1$

c. Solve for y:

$9x + abc = yz - mno$

$9x + abc + mno = yz$ adding mno to both sides

$$\frac{9x + abc + mno}{z} = y$$ dividing both sides by z

NOTE: With problems that contain more than one variable, it may not be practical to check for truth. However, double-check your work for accuracy.

1-14. EXERCISES, SECTION II

After you have completed these exercises, turn to the end of the lesson and check your answers with the solutions.

a $1 - (4c + 7) = -5c$

____

b. $8 + 6b = 3b - 4$

____

c. $r - (5 - [r - (4 + r)]) = 8$

____

d. $8z - 2z + 6 = z - 6$

____

e. $-\frac{5k}{9} = 2$

____

f. $4(v - 1) - 4(2v + 2) = 8$

____

g. $\frac{7r}{2} - \frac{9r}{2} = 16$

____

h. $\frac{y - 8}{5} + \frac{y}{3} = \frac{-8}{5}$

____

i. Solve for y:

$4x + 6y = 12$

____

j. $-7 - 5k = 10k - 7$

____

k. $p + 9 = 0$

____

l. $-7x + 1 = 50$

____

m. $3(x + 5) = 2x - 1$

____

n. $2 + \frac{m}{3} = \frac{m}{5}$

____

o. $5 - 8k = -19$

____

p. Solve for z:

$abz = pqr$

____

q. Solve for c:

$rst = abc - ghi$

____

r. Solve for x:

$$\frac{abc}{def} = \frac{rst}{xyz}$$

s. Solve for b:

b + c = 0

Section III. ROUNDING OFF NUMBERS

1-15. DISCUSSION

Rounding off of numbers is the dropping of one or more digits of a number to obtain the desired number of significant figures.

1-16. RULES

a. When the digit dropped is less than five, the last digit retained remains unchanged; e.g., 6.582 becomes 6.58.

b. When the digit dropped is greater than five, the last digit retained is increased by one (1); e.g., 6.586 becomes 6.59.

c. When the digit dropped is five alone or a five followed by only a zero or only zeros, the digit remaining is rounded to the nearest even number; e.g., 2.585 becomes 2.58. Since the remaining digit was already an even number it was not changed. However, the number 2.575 when rounded off becomes 2.58, since the remaining digit was seven.

NOTE: Although we will follow the above rule concerning the dropping of the number five, be aware that it is not always observed. It is quite common to round up if the digit dropped is five or greater and to leave the last digit unchanged if the digit dropped is four or less.

1-17. EXERCISES, SECTION III

After you have completed these exercises, turn to the end of the lesson and check your answers with the solutions.

FIRST REQUIREMENT: Round to the nearest tenths place (one digit to the right of the decimal point).

a. 78.99

____

b. 13.6

____

c. 56.019

____

d. 639.138

____

e. 1.27

____

f. 99.999

____

g. 24.351

____

h. 844.556

____

SECOND REQUIREMENT: Round to the nearest hundredths place (two digits to the right of the decimal point).

i. 6.7755

____

j. 100.999

____

k. 5.098

l. 9.999

m. 8.8151

n. 8.6555

THIRD REQUIREMENT: Round to the nearest thousandths place (three digits to the right of the decimal point).

o. 6.9895

p. 333.0003

q. 21.7666

r. 1.3375

s. 4.8525

t. 2.1999

Section IV. SCIENTIFIC NOTATION

1-18. DISCUSSION

In the clinical laboratory, elementary mathematics are frequently used by the laboratory specialist to calculate the amount of a specific chemical needed to prepare a reagent or to determine the concentration of a chemical constituent in a clinical sample. To enable the laboratory specialist to perform calculations with greater efficiency, the use of scientific notation must be mastered.

1-19. EXPONENTS

An exponent is a superscript number written to the right of a base number. An exponent indicates the number of times the base number is to be used as a multiplicative factor to produce a product equal to the exponential expression.

a. **General Examples**.

(1) 32 = 3 X 3 = 9

In the above example, the exponent is the number two and indicates that the number three, referred to as the base number, is to be used as a multiplicative factor twice.

(2) 23 = 2 X 2 X 2 = 8

In the above example, the exponent is the number three and indicates that the base number two is to be used as a multiplicative factor three times.

b. **Examples Using Powers of Ten (10)**.

(1) 102 = 10 X 10 = 100

(2) 103 = 10 X 10 X 10 = 1000

1-20. RULES OF EXPONENTIATION

The following rules apply to exponential expressions that have the same base with the exception of addition and subtraction.

a. **Addition and Subtraction**. The expressions must first be evaluated. Then, addition or subtraction is performed in the usual manner.

Example.

102 + 101 = 100 + 10 = 110

b. **Multiplication**.

xaxb = xa+b

Example.

2322 = 23+2 = 25 = 2 X 2 X 2 X 2 X 2 = 32

c. **Division**.

$$\frac{xa}{xb} = xa-b$$

Example.

$$\frac{24}{22} = 24-2 = 22 = 2 X 2 = 4$$

d. **Exponential Expressions Raised to a Power**.

(xa)b = xab

Example.

(22)3 = 2(2)(3) = 26 = 2 X 2 X 2 X 2 X 2 X 2 = 64

1-21. EXPRESSING NUMBERS IN TERMS OF SCIENTIFIC NOTATION

a. **Division.** Scientific notation uses the exponential method of expressing very large and very small numbers in terms of an exponential expression. Utilizing this method, numbers are expressed as a product of two numbers. The first number is called the digit term and is a number greater than or equal to one but less than ten. The second number of the product is called the exponential term and is written as a power of ten.

b. **Examples.**

(1) 100 = 1 X 10 X 10 = 1 X 102

In the above example, the digit term is the number one and the exponential term is ten squared (102).

(2) 200 = 2 x 10 X 10 = 2 X 102

In this example, the digit term is the number two and the exponential term is ten squared (102).

(3) 745 = 7.45 X 10 X 10 = 7.45 X 102

(4) 0.00486 = 4.86 X 10-1 X 10-1 X 10-1 = 4.86 X 10-3

c. **Decimal Movement.** In examining the above examples, it is easily seen that the exponent of the number ten is determined by movement of the decimal point. The exponent of the number ten is positive when the decimal is moved to the left and negative when moved to the right. Standard scientific notation refers to the placement of the decimal to the right of the first nonzero integer.

(1) Positive exponents. When the decimal point is moved to the left, the exponent of the number ten is always positive.

(2) Negative exponents. When the decimal point is moved to the right, the exponent of the number ten is always negative.

(3) Examples.

214 = 2.14 X 102

0.102 = 1.02 X 10-1

In the first example using the number 214, the decimal was moved two places to the left and thus the exponent of the number ten is two. In the second example using the number 0.102, the decimal was moved one place to the right and thus the exponent of the number ten is minus one (-1).

d. **Powers of Ten (10) Expressed in Scientific Notation**.

1,000,000 = 1 X 10^{6}

100,000 = 1 X 10^{5}

10,000 = 1 X 10^{4}

decimal moved to the left

1,000 = 1 X 10^{3}

100 = 1 X 10^{2}

10 = 1 X 10^{1}

1 = 1 X 10^{0} --

0.1 = 1 X 10^{-1}

0.01 = 1 X 10^{-2}

0.001 = 1 X 10^{-3}

decimal moved to the right

0.0001 = 1 X 10^{-4}

0.00001 = 1 X 10^{-5}

0.000001 = 1 X 10^{-6}

1-22. CHANGING THE SIGN OF THE EXPONENT

a. To change the sign of the exponent, move the exponential term to the denominator or numerator as appropriate with a concurrent change in sign. This procedure yields an equivalent expression.

Example.

$$\begin{array}{l} \text{Numerator} \\ \\ \text{Denominator} \end{array} \quad \frac{10^{-6}}{10^{4}} = \frac{10^{-4}}{10^{6}} = 10^{-10}$$

b. When an exponential term is moved from one side of the division line to the other, the exponent changes signs.

Example. Write and equivalent expression of 102, changing the sign of the exponent.

Solution.

$$\begin{array}{ll} \text{Numerator} & \\ & \dfrac{102}{1} = \dfrac{1}{10\text{-}2} \\ \text{Denominator} & \end{array}$$

c. Write an expression that changes the sign of the exponential term in 1/106 without changing the value.

Solution.

$$\frac{1}{106} = \frac{10\text{-}6}{1} = 10\text{-}6$$

d. Write an expression that changes the signs of the exponential without changing the value.

$$\frac{4.8 \text{ X } 10\text{-}7}{1.3 \text{ X } 104}$$

Solution.

$$\frac{4.8 \text{ X } 10\text{-}7}{1.3 \text{ X } 104} = \frac{4.8 \text{ X } 10\text{-}4}{1.3 \text{ X } 107} = 3.7 \text{ X } 10\text{-}11$$

1-23. MULTIPLICATION OF NUMBERS EXPRESSED IN SCIENTIFIC NOTATION

To multiply numbers expressed in scientific notation, multiply the <u>digit terms</u> in the usual way and add the exponents of the <u>exponential terms</u>.

a. Example. Multiply 102 by 103

Solution. In multiplying exponential terms, the exponents of the base ten (10) are added (the base numbers must be the same), therefore:

102 X 103 = 102+3 = 105

b. Example. Multiply 3.2 X 102 by 2.0 X 103

Solution. Remember that when multiplying exponential expressions, the <u>digit terms</u> are multiplied in the usual manner and the exponents of the <u>exponential terms</u> are added, therefore:

(3.2 X 102) X (2.0 X 103) = 3.2 X 2.0 X 102 X 103 =

3.2 X 2.0 X 102+3 = 6.4 X 105

c. Example. Multiply 3.2 X 10-2 by 2.0 X 103

Solution.

(3.2 X 10-2) X (2.0 X 103) = 3.2 X 2.0 X 10-2 X 103 =

6.4 X 10-2+3 = 6.4 X 101

Equivalently:

3.2 X 10-2
2.0 X 103
6.4 X 101

Multiply the digit terms and add the exponents of the exponential terms.

d. Example**.** Multiply 0.00056 by 0.013 using scientific notation.

Solution. Before you multiply, the above numbers must be converted to scientific notation. Multiply the digit terms and add the exponentential terms.

Thus, 0.00056 = 5.6 X 10-4 and 0.013 = 1.3 X 10-2

5.6 X 10-4
1.3 X 10-2
7.28 X 10-6 ----> 7.3 X 10-6

1-24. DIVISION OF NUMBERS EXPRESSED IN SCIENTIFIC NOTATION

a. **Method One.** To divide exponential expressions, the digit term of the numerator is divided by the digit term of the denominator in the usual way, and the exponent of the denominator is subtracted from the exponent of the numerator.

(1) Example. Divide 103 by 102

Solution. In division, the exponent of the denominator is subtracted from the exponent of the numerator. Therefore:

103 divided by 102 = 103-2 = 101 = 10

(2) Example. Divide 5.0 X 103 by 2.5 X 102

Solution. By using method one, the laboratory specialist divides 5.0 by 2.5 and then subtracts the exponents of the exponential terms. Therefore:

$$\frac{5.0}{2.5} = 2.0 \text{ and } 103\text{-}2 = 101$$

Thus, 5.0 x 103 divided by 2.5 X 102 = 2.0 X 101

b. **Method Two.** In this method of dividing exponential expressions, the digit term is treated in exactly the same manner as in method one. The exponential term in the denominator is moved to the numerator, with the appropriate sign change, and the exponents of the exponential terms in the numerator are added.

(1) Example. Divide 5.0 X 103 by 2.5 X 102

Solution.

$$\frac{5.0 \text{ X } 103}{2.5 \text{ X } 102} = \frac{5.0 \text{ X } 103 \text{ X } 10\text{-}2}{2.5} = \frac{5.0 \text{ X } 103\text{-}2}{2.5} = 2.0 \text{ X } 101$$

(2) Example. Divide 0.000005 by 250

Solution. Before the above numbers can be divided using the methods previously demonstrated, they must be converted to scientific notation.

Thus, 0.0000050 = 5.0 X 10-6 and 250 = 2.5 X 102

$$\frac{5.0 \text{ X } 10\text{-}6}{2.5 \text{ X } 102} = \frac{5.0 \text{ X } 10\text{-}6 \text{ X } 10\text{-}2}{2.5} = 2.0 \text{ X } 10\text{-}8$$

1-25. EXERCISES, SECTION IV

After you have completed these exercises, turn to the end of the lesson, and check your answers with the Academy solutions.

FIRST REQUIREMENT: Express the following numbers in standard scientific notation:

a 0.002406

b. 4742

c. 1.463

d. 774.82

e. 91

f. 0.2534

SECOND REQUIREMENT: Evaluate the following expressions:

g. 102 + 19

h. 2324

i. 2232

j. (101)3

____

k. $\frac{33}{30}$

____

l. $\frac{82}{8-2}$

____

THIRD REQUIREMENT: Evaluate the following expressions:

m. 4.0 X 10-11 divided by 2.0 X 10-3

____

n. 8.4 X 106 divided by 2.1 X 10-3

____

o. 3.4 X 105 multiplied by 2.2 X 102

____

p. 6.119 X 101 multiplied by 1.112 X 10-1

____

FOURTH REQUIREMENT: Write expressions that change the signs of the following exponentials without changing the value:

q. $\frac{1}{7-4}$

____

r. 10-16

____

s. $$\frac{5.5 \text{ X } 10\text{-}3}{9.1 \text{ X } 106}$$

____

t. $$\frac{10\text{-}8}{24}$$

____

Section V. SIGNIFICANT FIGURES

1-26. DISCUSSION

The validity of a number depends upon the precision to which the number was determined.

a. For example, an individual counted the number of ants that were in an anthill and reported that there were 246,337 ants. This number is obviously inaccurate since during the period of the count, ants died and others hatched. A more realistic number of ants would have been 246,000, or 2.46 X 105.

b. A common pitfall is the manner in which laboratory values are reported. The specialist can use a calculator to determine the value of an unknown specimen and report a value to the physician that is inaccurate. For example, in using the photometric equation, the value for glucose in a clinical specimen was determined to be 99.1825734 using the calculator. It would be absurd to report this value since the limits of accuracy of the measuring device used to obtain the calculation data (accuracy of the test) allows, at best, only a value of 99.

c. In most cases, the number of significant figures in your least accurate piece of data used in your calculations determines the accuracy of reported results.

1-27. DETERMINING THE NUMBER OF SIGNIFICANT FIGURES IN A NUMBER

a. **Nonzero Integers.** All nonzero integers are significant figures.

b. **Use of Zero.** A zero in a number may or may not be significant, depending upon the manner in which it is used.

(1) <u>Zeros used to locate decimal points</u>. If one or more zeros are used to locate a decimal point (place holder) they are not significant. For example, the numbers 0.025, 0.0025, and 0.00025 each have only two significant figures, the two and the five.

(2) <u>Zeros appearing between numbers</u>. When a zero appears between

numbers it is significant. For example, 1.095 has four significant figures.

(3) Zeros appearing at the end of a number.

(a) If a number contains a decimal point and the last number (digit) is a zero, the zero is a significant figure. For example, 15.60 has four significant figures.

(b) If the last digit in the number is a zero and the number does not contain a decimal point, the zero may or may not be significant. For example, the number 1670 has four significant figures if the accuracy of the measurement included the zero as a significant digit. If the digit seven was estimated, then the zero is not significant and hence the number contains only three significant figures.

NOTE: For all course work that follows, any trailing zeros will be considered significant. For example, the number 1000 has four significant figures.

c. **Examples.**

Number	Number of significant figures
18	2
18.0	3
108	3
0.0018	2
0.0108	3
180	3 (for this subcourse)

1-28. IMPLIED LIMITS

If a laboratory result is reported as 3.6, it indicates that this value is accurate to the nearest tenth and that the exact value lies between 3.55 and 3.65.

1-29. LABORATORY APPLICATION

The real importance of significant figures lies in their application to fundamental laboratory calculations.

a. **Addition and Subtraction.** When adding or subtracting, the last digit retained in the sum or difference should correspond to the first doubtful decimal place of the addends (least accurate number).

Example. Add 5.683 plus 0.0052.

Solution. In the number 5.683, the three is the doubtful decimal place; i.e., the value of this measurement could vary from 5.6825 to 5.6835. Since the fourth digit after the decimal point is unknown, the answer is limited to four digits. Thus,

$$\begin{array}{r} 5.683 \\ \underline{+\ 0.0052} \\ 5.6882 \end{array} \text{ ----> } 5.688$$

b. **Multiplication and Division.** When multiplying or dividing, the product or quotient can contain no more significant digits than the least number of significant figures in the numbers involved in the calculation.

Example. Multiply 0.5823 by 8.2.

Solution. The number 0.5823 contains four significant figures whereas the number 8.2 contains only two significant figures. Thus, the product can contain only two significant figures.

Therefore 0.5823
X 8.2
4.77486 -----> 4.8

1-30. EXERCISES, SECTION V

After you have completed these exercises, turn to the end of the lesson and check your answers with the solutions.

FIRST REQUIREMENT: Indicate the number of significant figures in the following:

a. 8400 _____

b. 4370 _____

c. 2.045 _____

d. 10.360 _____

e. 0.8400 _____

f. 209 _____

g. 482.000 _____

h. 80 _____

i. 1.95001 _____

j. 101 _____

k. 0.004300 _____

l. 0.0085 _____

m. 370 _____

n. 999 _____

o. 110.0 _____

p. 0018 _____

q. 49.0611 _____

r. 101.0 _____

s. 0.0850100 _____

t. 0.00001 _____

SECOND REQUIREMENT: Solve the following problems and report the answers using the appropriate number of significant figures:

u. 6012.14 + 305.2

v. 310.221 — 6.1

w. 0.01154 + 0.23

x. 100.2 + 85

y. 66 — 2

z. 18.9 X 21

aa. 0.269 - 3

ab. 662 - 18.0

ac. 75 X 801

ad. 0.21 X 3.0233

Section VI. LOGARITHMS

1-31. DISCUSSION

The "common" logarithm (log) or the logarithm to the base ten (log10) of a number is the exponent (power) to which the number ten must be raised to equal that number. For example, the logarithm of 100 is equal to two, since the exponent (power) to which the number ten (10) must be raised to equal 100 is two or 100 = 102. Since logarithms are exponents, they follow the "Rules of Exponentiation" previously discussed. The laboratory specialist can utilize logarithms to perform multiplication, division, find roots, and raise a number to a power. A second use of logarithms is in the solving of a number of equations used in the clinical laboratory; e.g., pH = -log [H+] and absorbance = 2 - log %T

NOTE: The work with logarithms in this subcourse will consist of traditional manual methods with tables (see Appendix B) rather than the use of a calculator to find logarithms and related values. Please remember this when you do the exercises and the examination items. Logarithms are approximate values. Comparable operations using a calculator will yield slightly different results in most instances.

1-32. EXAMPLES

Number	Number expressed exponentially	Logarithm (log)
10000	104	4
1000	103	3
100	102	2
10	101	1
1	100	0
0.1	10-1	-1
0.001	10-3	-3
0.0001	10-4	-4

It is easily seen that the exponent (power) to which the number ten has been raised is the same as the logarithm.

1-33. PARTS OF A LOGARITHM

A logarithm is composed of two parts.

a. The characteristic is the portion of the logarithm that lies to the left of the decimal point. It is a whole number that may be negative or positive, depending upon the original number.

b. The mantissa is the portion of the logarithm that lies to the right of the decimal point. The mantissa is a decimal fraction of one (1) and it is always positive.

c. For example, the logarithm of 20 is equal to 1.3010, where the number one is the characteristic and the number .3010 is the mantissa.

1-34. DETERMINATION OF THE CHARACTERISTIC

In order to find the logarithm of a number, the characteristic must first be determined. The characteristic is always determined by inspection of the original number or by applying a simple rule.

a. **By Inspection.**

(1) If we take the number 20, we see that it lies between 10 and 100, which have logarithms of one and two, respectively. Therefore, the logarithm of the number 20 must be between one and two. The logarithm of 20 is 1.3010 and hence the characteristic (remembering that the characteristic is the number to the left of the decimal point) of the logarithm is one.

(2) An alternate method is to express the original number in scientific notation. The power of ten in the expression is the characteristic of the original number. The number 20 when expressed in scientific notation is 2.0 X 101. Hence, the characteristic of the original number is one.

NOTE: Both methods for determining the characteristic may be used for numbers less than or greater than one.

b. **By Rule.**

(1) Original numbers greater than one. If the number whose logarithm is being determined is greater than one, the characteristic for the logarithm of that number will be <u>positive</u> and <u>one less</u> than the number of digits found to the left of the decimal point in the original number. For example:

Original Number	Number of digits to left of decimal	Characteristic (digits - one)
2.0	1	0
20.0	2	1
200.0	3	2
2000.0	4	3
20000.0	5	4

(2) Original numbers that are less than one. If the number whose logarithm is being determined is less than one, the characteristic for the logarithm of that number will be <u>negative</u> and is a number <u>one more</u> than the number of zeros to the right of the decimal point that precede the first nonzero integer. For example:

Original Number	Number of zeros to right of decimal	Characteristic (num. of 0 + 1)
0.2	0	-1
0.02	1	-2
0.002	2	-3
0.0002	3	-4

(3) A negative characteristic may be distinguished from a positive characteristic by placing a negative sign <u>above</u> the characteristic. This is necessary since the mantissa is always positive. For example:

$$\log 0.2 = \bar{1}.3010$$

(4) Equivalently, the negative logarithm of 0.2 could be expressed as the difference between the characteristic and mantissa. This is done to facilitate certain methods of calculation, pH in particular. In this case, the negative sign is placed in front of the difference. For example:

$$\log 0.2 = \bar{1}.3010 = 0.3010 - 1 = -0.6990$$

This method is considered by most authors to be an incorrect method to express the logarithm of a number less than one because the mantissa may never be negative. However, this method is commonly used by electronic calculators and certain other methods of calculation. You should be able to use either method without error in the laboratory, as long as you consistently use either approach.

(5) A logarithm with a negative characteristic may also be changed to a positive form by adding ten to the characteristic and adding minus (-) ten <u>after</u> the mantissa. Since ten is both added and subtracted to the logarithm the operation does not change the value of the logarithm. For example:

$$\log 0.2 = \bar{1}.3010$$

Add ten to the characteristic

$$10 + (-1) = 9 \text{ ----> } 9.3010$$

Add minus ten (-10) <u>after</u> the mantissa

$$9.3010 - 10 = \bar{1}.3010$$

1-35. DETERMINATION OF THE MANTISSA

The mantissa of the logarithm is found in the table of "common" logarithms (see Appendix B) and is <u>independent</u> of the position of the decimal point in the original number. Thus, the numbers 0.02, 0.2, 2, 20, 200, 2,000, and 20,000 all have the same mantissa; i.e., 0.3010.

a. In determining the number to obtain the mantissa, do not consider preceding zeros.

b. The number for which a mantissa is desired, commonly referred to as the <u>natural number</u>, must contain at least three digits, the last two of which may be zeros.

c. For numbers that contain less than three digits, add enough zeros to yield a three digit number.

d. Numbers that contain more than three nonzero digits should be rounded to three significant figures.

e. Consider the following examples:

Original Number	Base Number
0.312	312
1.031	103
55.357	554
4	400
0.9	900

f. Examine the Appendix B, Table of Four-Place Logarithms. The table consists of a series of columns and rows of four—digit numbers bordered left by a column of two—digit numbers and uppermost by a row of single—digit numbers.

g. The first two digits of the natural number can be found in the first column.

h. The third digit of the number can be found in the uppermost row.

i. To determine the mantissa of a number, find the first two digits of the number in the left column. Move across this row to the column headed by the third digit of the number.

j. The four - digit number at the intersection is the mantissa of the given natural number.

1-36. DETERMINATION OF THE LOGARITHM OF A NUMBER

To determine the logarithm of a number, both the characteristic and the mantissa must be found. The characteristic must be determined by inspection (or rule), and the mantissa must be looked up in a table of "common" logarithms (a calculator may also be used). By close examination of the following table, the determination of the logarithm of a number should become clear.

Original Number	Characteristic	Mantissa	Logarithm
0.2	-1	3010	$\bar{1}.3010$
20	1	3010	1.3010
200	2	3010	2.3010
2000	3	3010	3.3010
20000	4	3010	4.3010
0.030	-2	4771	$\bar{2}.4771$
412	2	6149	2.6149
5490	3	7396	3.7396

1-37. LOGARITHMS OF NUMBERS EXPRESSED AS EXPONENTIALS

The determination of the logarithm becomes quite simple when the original number is written in scientific notation; i.e., as an exponential expression. This form of expressing the original number also should help you understand the meaning of the characteristic and the mantissa.

a. **Example 1.** Find the logarithm of the number 4120.

Solution. First write 4120 in exponential form.

4120 becomes 4.12 X 103

Take the logarithm of the exponential expression.

log (4.12 X 103)

When determining the logarithm of two numbers being multiplied together (4.12 X 103), the logarithm of <u>each number</u> is determined and the logarithms are then added.

log (4.12 X 103) = log 4.12 + the log 103

The logarithm of 4.12 is 0.6149 (mantissa) and the logarithm of 103 is three (characteristic). Thus, the logarithm of 4120 = 0.6149 + 3 or 3.6149.

b. **Example 2.** Find the logarithm of 0.000412.

Solution.

0.000412 = 4.12 X 10-4

log (4.12 X 10-4) =

log 4.12 + log 10-4 =

0.6149 + (-4) =

$\bar{4}$.6149 or equivalently, with a calculator, -3.3851

1-38. DETERMINATION OF ANTILOGARITHMS

The antilogarithm (antilog) is the number corresponding to a logarithm. For example, the antilogarithm of three is the number whose logarithm equals three. The number whose logarithm equals three is 1000; thus, the antilogarithm of three is 1000.

a. **Mantissa.** The mantissa of the logarithm can be found in the table of "common" logarithms, and the three digits corresponding to the mantissa is written down. In the event that the exact mantissa value is not in the table, find the closest mantissa in the columns without exceeding the value of the mantissa being worked with. Remember the mantissa is the number to the right of the decimal point.

Examples.

Mantissa	Corresponding Digits
0.8451	700
0.3010	200
0.2945	197
0.6981	499
0.9996	999

b. **Characteristic.** The characteristic of the logarithm determines where the decimal point will be placed in the number corresponding to the mantissa when determining the antilogarithm.

Examples.

Characteristic	Number of digits to the left of the decimal
0	1
1	2
2	3
3	4

c. **Example 1.** Find the antilogarithm of 2.000.

Solution. Determine the digits that correspond to the mantissa. The mantissa is .000 and the digits corresponding to this mantissa in the table of "common" logarithms are 100. The characteristic of two indicates that there are three digits or places to the left of the decimal point. Thus, antilog 2.000 = 100. To check your answer, take the logarithm of 100:

log 100 = 2

d. **Example 2.** Find the antilogarithm of 3.2989.

Solution. The digits corresponding to the mantissa (0.2989) are 199. The characteristic indicates that there are four digits to the left of the decimal point. Therefore, antilog 3.2989 = 1990.

e. **Example 3.** Find the antilogarithm of $\bar{3}.2989$.

Solution. Rewrite $\bar{3}.2989$ as 0.2989 - 3. This puts the logarithm in the form of an exponential expression. Take the antilogarithm of both numbers.

Antilog 0.2989 X antilog -3.

NOTE: The two antilog expressions are multiplied together since they represent numbers, not logs (refer to the Rules of Exponentiation in paragraph 1--20).

The antilogarithm of 0.2989 = 1.99

The antilogarithm of minus three (-3) is 10-3

Thus, the antilogarithm of $\bar{3}.2989$ = 1.99 x 10-3

f. **Example 4.** Use the logarithm table to find the antilogarithm of -2.5017, a logarithm from an electronic calculator.

Solution. In the calculator, the negative characteristic and positive mantissa were combined by algebraic addition to form the single negative number. To change a negative logarithm to the correct form, add the next highest positive whole number above the characteristic's (absolute value) to the negative log. This operation will yield the positive mantissa.

3 - 2.5017 = 0.4983

Use the additive inverse (negative form in this instance) of the whole number that was added as the characteristic of the corrected logarithm, and indicate its sign with a bar over the number.

$\bar{3}.4983$

Rewrite $\bar{3}.4983$ as 0.4983 - 3. This puts the logarithm in the form of an exponential expression. Take the antilog of both numbers.

Antilog 0.4983 X antilog (-3)

NOTE: The two antilog expressions are multiplied together since they represent numbers, not logs (refer to the Rules of Exponentiation in paragraph 1-20).

The antilog of 0.4983 = 3.15

The antilog of -3 = 10-3

Therefore, the antilog of -2.5017 = 3.15 X 10-3

1-39. USE OF LOGARITHMS IN MULTIPLICATION

To multiply two or more numbers together, the logarithm of each number is found and the logarithms of each are added together. The antilogarithm of the sum of the logarithms is then taken to give the product of the multiplication.

Example. Multiply 152 by 63 using logarithms.

Solution.

log 152 = 2.1818
log 63 = $\underline{+ 1.7993}$
Sum 3.9811

antilog 3.9811 = 9.57 X 103

Thus, 152 x 63 = 9.6 X 103, with 2 significant figures

NOTE: If you were to multiply these two numbers together by the usual method you would have obtained 9576 instead of 9570 as was determined by the use of logarithms due to the accuracy of the logarithm tables. <u>Logarithms are approximate values</u>.

1-40. USE OF LOGARITHMS IN DIVISION

To divide two numbers, the logarithms of the numbers are subtracted; i.e., the logarithm of the divisor (denominator) is subtracted from the logarithm of the dividend (numerator). The antilogarithm of the difference of logarithms is then taken to give the quotient.

Example**.** Divide 152 by 63 using logarithms.

Solution.

log 152 = 2.1818
log 63 = <u>-1.7993</u>
Difference 0.3825

antilogarithm of 0.3825 = 2.41

Thus, 152 divided by 63 = 2.4, with 2 significant figures

1-41. USE OF LOGARITHMS TO FIND ROOTS OF NUMBERS

To find the root of a number, the logarithm of the number is determined; the logarithm of the number is next divided by the root desired; e.g., if the square root of a number is wanted, the logarithm of the number is divided by two; if the cube root is required, divide by three, etc. The antilogarithm of the quotient is taken; the resulting number is the root of the number.

a. **Example 1.** Find the square root of 625.

Solution.

Square root 625 = (625)1/2

log (625)1/2 = 1/2 log 625 = 1/2 X 2.7959 = 1.3980

antilogarithm 1.3980 = 25.0

Thus, the square root of 625 = 25.0

b. **Example 2.** Find the fourth root of 625.

Solution.

Fourth root 625 = (625)1/4

log (625)1/4 = 1/4 log 625 = 1/4 X 2.7959 = 0.6990

antilog 0.6990 = 5.00

Thus, the fourth root of 625 = 5.00

1-42. USE OF LOGARITHMS TO FIND THE PRODUCT OF NUMBERS WITH EXPONENTS

To find the product of a number that has an exponent, determine the logarithm of the number and multiply by the exponent. The antilogarithm of the product is taken. For example, if a number has been squared, find the logarithm of the number and multiply by two, then find the antilog.

a. **Example 1.** Find the product of (625)2 using logarithms.

Solution.

log 6252 = 2 log 625 = 2 X 2.7959 = 5.5918

antilog 5.5918 = 391,000 = 3.91 X 105

NOTE: If 625 is squared with a calculator, the resulting number is 390,625.

b. **Example 2.** Find the product of 62510 using logarithms.

Solution.

log 62510 = 10 log 625 = 10 X 2.7959 = 27.959

antilog 27.959 = 9.10 X 1027

1-43. EXERCISES, SECTION VI

After you have completed the following exercises, turn to the end of the lesson, and check your answers with the solutions.

FIRST REQUIREMENT: Determine the logs for the following numbers.

a. 695,100 ______________________

b. 821 ______________________

c. 0.00063 ______________________

d. 3421 ______________________

e. 0.0458 ______________________

SECOND REQUIREMENT: Determine the antilogs of the following logarithms.

f. 8.6990 – 10 ______________________

g. -2.1864 ______________________

h. 1.7774 ______________________

i. 1.5400 ______________________

j. 1.1075 ______________________

THIRD REQUIREMENT: Evaluate the following expressions using logarithms.

k. 44 X 89 ______________________

l. 512 X 600 ______________________

m. 100,000 X 1000 ______________________

n. 0.011 X 0.54 ______________________

o. 0.00213 X 18 ______________________

FOURTH REQUIREMENT: Evaluate the following expressions using logarithms.

p. 2259 – 8

q. 384 – 8

r. 0.00854 - 0.127

s. 1620 - 67

t. 2 - 83

FIFTH REQUIREMENT: Find the square roots using logarithms.

u. 0.000625

v. 2,250,000

w. 169

x. 0.0001

y. 121

Section VII. SOLUTIONS TO EXERCISES

1-44. SOLUTIONS TO EXERCISES, SECTION I (PARAGRAPH 1-8)

a. Associative property (para 1-2b)

b. Commutative property (para 1-2a)

c. Inverse property for multiplication (para 1-2d)

d. Identity property for multiplication (para 1-2c)

e. Distributive property (para 1-2e)

f. Inverse property (para 1-2d)

g. Identity property for addition (para 1-2c)

h. Commutative property (para 1-2a)

i. Distributive property (para 1-2e)

j. Inverse property for addition (para 1-2d)

k. 18 (para 1-5a)

l. -15 (para 1-3a)

m. 16 (para 1-7)

n. 12 (para 1-7)

o. -1.27 or —19/15 (para 1-7)

p. 180 y + 24 (para 1-7)

q. 9 (para 1-7)

r. 0 (para 1-7)

s. 0.778 or 7/9 (para 1-7)

t. -2/8 or -1/4 (para 1-7)

1-45. SOLUTIONS TO EXERCISES, SECTION II (PARAGRAPH 1-14)

Reference paras 1-10 through 1-12:

a. $c = 6$

b. $b = -4$

c. $r = 17$

d. $z = -2.4$

e. $k = -3.6$ or $-18/5$

f. $v = -5$

g. $-r = 16$ or $r = -16$

h. $y = 0$

i. $$Y = \frac{12 - 4x}{6}$$

j. $k = 0$

k. $p = -9$

l. $x = -7$

m. $x = -16$

n. $m = -15$

o. $k = 3$

p. $$z = \frac{pqr}{ab}$$

q. $$\frac{rst + ghi}{ab} = c$$

r. $$x = \frac{defrst}{abcyz}$$

s. $b = -c$

1-46. SOLUTIONS TO EXERCISES, SECTION III (PARAGRAPH 1-17)

Reference paragraph 1-16:

a. 79.0

b. 13.6

c. 56.0

d. 639.1

e. 1.3

f. 100.0

g. 24.4

h. 844.6

i. 6.78

j. 101.00

k 5.10

l. 10.00

m. 8.82

n. 8.66

o. 6.990

p. 333.000

q. 21.767

r. 1.338

s. 4.852 (para 1-16c)

t. 2.200

1-47. SOLUTIONS TO EXERCISES, SECTION IV (PARAGRAPH 1-25)

a. 2.406 X 10-3 (para 1-21)

b. 4.742 X 103 (para 1-21)

c. 1.463 X 100 (para 1-21)

d. 7.77482 X 102 (para 1-21)

e. 9.1 X 101 (para 1-21)

f. 2.534 X 10-1 (para 1-21)

g. 119 (para 1-20a)

h. 128 (para 1-20b)

i. 36 (para 1-19)

j. 1,000 (para 1-20d)

k. 27 (para 1-20c)

l. 4,096 (para 1-20c)

m. 2.0 X 10-8 (para 1-24)

n. 4.0 X 109 (para 1-24)

o. 7.48 X 107 (para 1-23)

p. 6.804 X 100 (para 1-23)

q. $\frac{74}{1}$ (para 1-22)

r. $\frac{1}{1016}$ (para 1-22)

s. $\frac{5.5 \text{ X } 10\text{-}6}{9.1 \text{ X } 103}$ (papa 1-22)

t. $\frac{2\text{-}4}{108}$ (para 1-22)

1-48. SOLUTIONS TO EXERCISES, SECTION V (PARAGRAPH 1-30)

Reference paragraph 1-27:

a. 4

b. 4

c. 4

d. 5

e. 4

f. 3

g. 6

h. 2

i. 6

j. 3

k. 4

l. 2

m. 3

n. 3

o. 4

p. 2

q. 6

r. 4

s. 6

t. 1

Reference is paragraph 1—29:

u. 6317.3

v. 304.1

w. 0.24

x. 185

y. 64

z. 4.0 X 102 (only 2 significant figures)

aa. -2.731 = -3 (only 1 significant figure)

ab. 644

ac. 6.0 X 104 (only 2 significant figures)

ad. 0.63

1-49. SOLUTIONS TO EXERCISES, SECTION VI (PARAGRAPH 1-43)

a. 5.8420 (paras 1-34 thru 1-36)

b. 2.9143 (paras 1-34 thru 1-36)

c. by calculator: -3.2007 (paras 1-34 thru 1-36)

by tables: $\bar{4}.7993$

d. 3.5340 (paras 1-34 thru 1-36)

e. by calculator: -1.3391 (paras 1-34 thru 1-36)

by tables: $\bar{2}.6609$

f. 0.0500 (para 1-38)

g. by calculator: 6.51 X 10-3

by tables: -2.1864 = $\bar{3}.8136$

antilog $\bar{3}.8136$ = 6.51 X 10-3 (para 1-38)

h. 59.9 (para 1-28)

i. 34.7 (para 1-38)

j. 12.8 (para 1-28)

k. 3.9 X 103 (para 1-39)

l. 3.07 X 105 (para 1-39)

m. 100,000,000 or 108 (para 1-39)

n. 0.0059 (para 1-39)

o. 0.038 (para 1-39)

p. 282 (para 1-40)

q. 48 (para 1-40)

r. 0.067 (para 1-40)

s. 24.2 (para 1-40)

t. 0.024 (para 1-40)

u. 0.025 (para 1-41)

v. 1500 (para 1-41)

w. 13 (para 1-41)

x. 0.01 (para 1-41)

y. 11 (para 1-41)

End of Lesson 1

LESSON ASSIGNMENT

LESSON 2	Introduction to Solution Mathematics.
TEXT ASSIGNMENT	Paragraphs 2-1 through 2-33.
LESSON OBJECTIVE	After completing this lesson, you should be able to: 2-1. Select the statement which best describes the types and the requirements of standards and reagents. 2-2. Select the statement that best describes the parts of a solution and the different types that can be made for laboratory use. 2-3. Calculate percentages of solution concentrations by weight/volume, volume/volume, and composition.
SUGGESTIONS	After studying the assignment, complete the exercises at the end of this lesson. These exercises will help you to achieve the lesson objectives. Before studying this lesson, please read Appendix A, Review of Dimensional Analysis. The information provides useful guidelines and strategies to help you throughout the rest of this subcourse.

LESSON 2

INTRODUCTION TO SOLUTION MATHEMATICS

Section I. GRADES OF LABORATORY CHEMICALS

2-1. DISCUSSION

Chemicals available from various manufacturers are usually assigned one of several common grades. The particular grade assigned to a chemical depends upon the purity. However, the differentiation between the various grades is unofficial. There seems to be no agreement between chemical manufacturing companies concerning the various degrees of purity.

2-2. STANDARDS

A standard is a solution that contains a known amount of analytically pure (weighed or titrated) substance in a specific volume of solvent.

a. **Primary Standards.** Primary standards are highly purified chemicals which assay to be at least 99.95 from the National Bureau of Standards (NBS), U.S. Department of Commerce. A primary standard must fulfill the following criterion: (1) it can be accurately weighed or measured, (2) it must be stable, (3) it must be extremely pure, and (4) it must be nonhygroscopic (not absorb water readily).

b. **Secondary Standards.** If a standard is prepared from a substance whose concentration cannot be determined accurately either by weighing or measuring, it is called a secondary standard. The normality (concentration) of the secondary standard must be determined by titration (or other means) using a primary standard. HCl, H_2SO_4, and NaOH are laboratory reagents that fall into the category of secondary standards. If an exact concentration is desired, these reagents must be standardized prior to use.

2-3. REAGENT GRADE OR ANALYTICAL REAGENT GRADE

The chemicals, reagent grade or analytical reagent grade (ACS), that are assigned this grade are manufactured to meet specifications needed in both quantitative and qualitative analysis and are generally the purest form of laboratory chemicals available. This grade of chemical appears in two different forms: (1) lot analyzed, in which each lot is individually analyzed and the amount of impurities noted on the bottle, and (2) maximum impurities, which lists the maximum allowed impurities on the bottle. This grade of chemical also contains certain types of chemicals and solvents which are designed for specific laboratory determinations. These specialized reagents are:

a. **High Performance Liquid Chromatography Solvents.** This grade of solvent is designed to be used in high performance liquid chromatography (HPLC).

b. **Nanogram Solvents.** These solvents are highly purified for electron capture gas liquid chromatography (GLC) techniques such as pesticide residue analysis and steroid determinations.

c. **Liquid Scintillation Chemicals.** These are specifically formulated to meet the needs of scintillometric radioisotope counting techniques.

d. **Spectrophotometric Solvents.** These solvents contain low levels of contaminants and give reproducible absorption curves throughout the entire spectrum.

2-4. ORGANIC GRADE

Organic laboratory chemicals are of suitable purity for most research work and for general laboratory purposes such as standard and reagent prepa- ration. Specifications are usually determined by melting or boiling point.

2-5. CHEMICALLY PURE GRADE

Chemicals that are chemically pure grade (CP) cannot be used as standards in either the clinical laboratory or in research without first being analyzed. There are no set specifications for this grade of chemical and the quality varies from manufacturer to manufacturer. This grade of chemical is usually satisfactory for most reagents used in the chemical laboratory.

2-6. UNITED STATES PHARMACOPEIA AND NATIONAL FORMULARY GRADES

The United States Pharmacopeia (U.S.P.) or the National Formulary (N.F.) sets forth the specifications for the chemicals that are assigned the above listed grades. The specifications for this grade of chemical state that the impurities present may not be of such concentration as to be harmful to one's health. This grade of chemical is usually unsatisfactory for use in preparing standards or reagents. It should only be used in emergency situations.

2-7. PURIFIED, PRACTICAL, OR PURE GRADES

There are no definite specifications for these grades of chemicals; they should not be used in the preparation of standards or reagents.

2-8. TECHNICAL OR COMMERCIAL GRADES

The purity of this grade of chemical is dependent upon the ease with which a certain impurity is removed. This grade of chemical is only suited for industrial use.

Section II. GRADES OF LABORATORY WATER

2-9. DISCUSSION

Tap water contains various impurities, both organic and inorganic. These impurities may interfere with many of the chemical tests that are routinely performed and consequently contributed to invalid test results. For example, calcium is a commonly performed laboratory test, and some tap water contains considerable amounts of calcium. To eliminate the possibility of interfering substances in water, either deionized or distilled water must be used in the preparation of all standards and reagents. Water for use in the laboratory is broken down into three basic types: Type I, Type II, and Type III.

2-10. TYPE I WATER

This is the purest type of water and is used for laboratory tests where the greatest accuracy is desired. This type of water is designed for use with atomic absorption, flame photometry, pH determinations, and fluorometry.

2-11. TYPE II WATER

This type of water is for general laboratory use and can be used for any type of laboratory determination that does not specify Type I water. Of the type of procedures listed under Type I water, only atomic absorption and fluorometry require the use of Type I water.

2-12. TYPE III WATER

Type III water is used for rinsing glassware and is suitable for <u>most</u> qualitative determinations.

2-13. MINIMUM SPECIFICATIONS FOR THE THREE TYPES OF WATER

Specifications	I	II	III
Specific conductance microhms, maximum	0.1	2.0	5.0
Specific resistance megohms, minimum	10.0	0.5	0.2
Silicate, mg/dL, maximum	0.01	0.01	0.01
Heavy metals, mg/dL, maximum,	0.01	0.01	0.01
Permanganate reduction	pass test	pass test	pass test

2-14. REQUIREMENTS FOR STANDARDS AND REAGENTS

The following are recommended requirements for all water used in the laboratory for the preparation of standards and reagents.

Requirement	Level
Ammonia	Less than 0.1 mg/L
Dissolved carbon dioxide	Less than 3 mg/L
PH	6.0 to 7.0
Sodium	Less than 0.1 mg/L

2-15. CARBON-DIOXIDE-FREE WATER

a. Carbon-dioxide-free water can be prepared by boiling Type II water for 20 minutes in an acid washed chemically resistant glass flask (such as Pyrex or Kimax brand borosilicate glass).

b. Remove the heat source and immediately close securely with a stopper fitted with a soda lime drying tube or a NaOH air trap and allow to cool.

c. Prevent atmospheric exposure and contamination.

d. Carbon-dioxide-free water should be used for prothrombin times, certain reference buffers, and National Bureau of Standards (NBS) standards. It is also recommended that lyophilized material be reconstituted with carbon-dioxide-free water.

2-16. FREQUENCY OF WATER TESTING

The specifications described by the College of American Pathologists (CAP) on water quality should be applied in performing a water quality check. These specifications should be checked weekly on both flowing and stored water. The check site of the test should be identified and the results compared to established limits and logged on an appropriate chart. If results exceed established limits, corrective measures must be taken according to CAP guidelines and a record of action taken placed in the log.

Section III. PARTS OF A SOLUTION

2-17. INTRODUCTION

In order to discuss a solution, one first needs a definition of the term. A <u>solution</u> is a homogeneous mixture of gases, solids, or liquids in which the individual molecules of two or more substances are evenly dispersed throughout the medium (solution). For example, if one dissolves sucrose in water, the sugar granules will break down into individual molecules. This is to say, if one were to sample any two sections of the solution, the same number of sugar molecules would be found per unit volume. Solutions are not always liquids; that is, if liquid mercury is rubbed on solid gold, the solid gold will dissolve the liquid mercury resulting in a solid solution.

2-18. SOLVENT

A solution consists of two parts. The first part is the solvent, which is defined as that portion of the solution that is either in greatest concentration or the portion of the solution that is doing the dissolving. Water is usually considered to be the universal solvent, no matter what proportion of the solution it is.

2-19. SOLUTE

The solute is the second part of a solution. One can define the solute as the substance that is being dissolved or the substance that is in lower concentration.

2-20. EXAMPLES

If one were to take 5 grams of sodium hydroxide and dissolve it in 100 milliliters of deionized water, what would be the solvent and what would be the solute? If you selected sodium hydroxide as the solute and water as the solvent, you were correct. The reason the sodium hydroxide is the solute is because it is present in the lower concentration and it is also the substance being dissolved. Water is the solvent because it is (1) the universal solvent; (2) present in the greatest concentration; and (3) the substance that is doing the dissolving.

Section IV. TYPES OF SOLUTIONS

2-21. INTRODUCTION

As stated, a solution is a homogeneous mixture of gases, liquids, or solids in which the molecules of two or more substances are evenly dispersed throughout the medium. This can be more specifically defined as a true solution because the substances are mixed on a molecular basis. There are two other types of solutions with which one should be familiar--the colloidal solution and the emulsion.

2-22. COLLOIDAL SOLUTION

A colloidal solution is not a true solution since the individual molecules are not mixed on a molecular basis; hence, the solution is not homogeneous throughout. This type of solution contains aggregates of molecules with each individual aggregate being denoted as a colloid or as a colloidal particle. The number of molecules in each aggregate may be from several hundred to a few thousand. The size of each colloidal particle is usually in the range of from 1 to 200 nanometers (nm) in diameter. This range has been arbitrarily set since particles with a diameter of less than 1 nm do not scatter light and particles greater than 200 nm can be seen with the light microscope. As was mentioned above, there is no homogeneity of particles in this type of solution. The particles are actually suspended in the solvent. A colloidal solution is broken down into two parts: (1) the dispersed phase (analogous to the solute of a true solution) and (2) the dispersion medium (analogous to the solvent of a true solution). The most important colloidal systems are those involving a solid dispersed in a liquid and are commonly denoted as colloidal suspensions. In some determinations of amylase, the starch substrate is a colloidal suspension.

2-23. EMULSIONS

Emulsions are colloidal solutions in which the dispersed phase and the dispersion medium are immiscible. If one mixes water and oil together and then shakes the container, the oil will break into very small particles that will disperse within the water. Upon standing, the oil droplets (dispersed phase) will tend to coalesce, forming larger and larger droplets. These droplets eventually become too large to remain within the dispersion medium (water) and the two liquids will separate. If one wishes to stabilize an emulsion, an emulsifying agent such as a bile salt must be added.

2-24. SATURATED SOLUTIONS

The laboratory specialist may on occasion be required to prepare a saturated solution. This type of solution is defined as a solution in which the dissolved solute is in equilibrium with undissolved solute. At a given temperature, most solutions can dissolve (hold) only a given number of solute particles. Once this solution contains all the solute particles that it can "hold," any additional solute will settle to the bottom of the container. Once this point has been reached, the solution is said to be saturated. As stated, there is an equilibrium between dissolved and undissolved solute. More clearly, this is to say that dissolved solute is continuously leaving the solution and entering into the crystal lattice of the undissolved solute. At the same time, undissolved solute is leaving the crystalline state and entering into solution.

2-25. UNSATURATED SOLUTIONS

An unsaturated solution is one that does not contain all the solute molecules that the solution could possibly "hold."

Section V. PERCENT CONCENTRATION

2-26. DISCUSSION

The simplest solution to prepare in the laboratory is the percent solution. The main reason that this type of solution is commonly used is the ease in calculating the amount of solute needed in its preparation. The term percent is a shortened form of percentage, which means parts per 100 total parts. This method is used with all three of the basic comparisons of percent concentration statements, that is, weight/weight, weight/volume, and volume/volume. The percent weight/weight is not commonly used in the clinical laboratory.

2-27. WEIGHT/VOLUME PERCENT SOLUTIONS

a. The weight per unit volume system of percent concentration is the most frequently used method in the clinical laboratory. This method is used when a solid solute is mixed with a liquid solvent. Since percent is parts per 100 parts of total solution, then a weight/volume percent solutions (w/v) unit would be grams of solute per 100 milliliters of total solution (g/100 mL). In the clinical laboratory and related situations, any time a w/v solution is described by a number followed by a percent symbol (grams per 100 mL (g/100 mL). Since 100 milliliters is equal to one deciliter, then a percent w/v concentration can be expressed as grams per deciliter (g/dL).

b. Another expression of w/v percent concentration used in the laboratory involves the use of fractional parts of the gram while retaining the unit of volume of solution. Examples of this kind would be milligram per deciliter or milligram percent. These may be expressed as mg/dL, mg/100 mL, or mg

c. The following problems may be solved using dimensional analysis (see Appendix A); however, simple ratio and proportion will suffice in some instances. Both methods of problem solving will yield comparable results.

d. Example. If a 5.0 g/dL solution of NaCl is to be prepared with a volume of 50 mL, how much NaCl would be required?

Solution. Based on definition the same problem could be solved using appropriate factors.

By dimensional analysis, we would solve as follows:

After reading the problem carefully it should be quite evident that the desired quantity is the amount of solute expressed in grams.

Express the volume in dL

$$50.\ 0 \text{ mL X} = \frac{1 \text{ dL}}{100 \text{ mL}} = 0.500 \text{ dL}$$

Multiply the volume expressed in deciliters times the percent concentration to determine the amount of solute contained in the specified amount of total solution.

$$0.500 \text{ dL X} = \frac{5.0 \text{ g}}{\text{dL}} = \underline{2.5 \text{ g}}$$

e. Examining the solution reveals that the given volume in milliters is first converted to deciliters, using the appropriate conversion factor. The volume of solution expressed in deciliters is then multiplied by the weight per volume concentration to yield the desired quantity, mass of solute expressed in grams per given volume. Notice that if the correct factor is used the unwanted units cancel out.

2-28. SOLVING WEIGHT/VOLUME PROBLEMS

The first step in solving any solution problem is to carefully read the problem. Remember that a unit of concentration consists of two components-- a weight (or volume) and a volume, thus w/v or v/v. Also notice that in reading the problem, the wording will suggest that a solution is to be prepared. Words such as "prepare" or "make" will be used in this type of problem. When making a percent solution, three things are necessary, that is, mass of solute needed, concentration of the solution to be prepared, and total volume of solution. Consider and see that all three of these requirements are met. Mass of solute is usually expressed in grams or milligrams. Concentration of the solution in g/dL, mg/dL, or (mL/dL). Total volume of solution is expressed in deciliters. Use the information given in the problem or calculated from information given in the problem as appropriate factors.

a. **Example 1**. How much NaCl is needed to make 2.0 liters of a 3.0 g/dL NaCl solution?

Solution. The first step of any problem is to read the problem carefully. Determine what the problem is asking for.

Grams of solute (NaCl).

Express the volume in deciliters.

$$2.0\text{ L X }\frac{10\text{ dL}}{1\text{ L}} = 20\text{ dL}$$

Multiply the volume expressed in deciliters times the percent concentration to determine the grams of NaCl contained in 2.0 liters of a 3.0 g/dL solution.

$$20\text{ dL X }\frac{3.0\text{ g}}{\text{dL}} = \underline{60\text{ g}}$$

b. **Example 2**. How many milliliters of solution can be prepared from 62.5 mg of K_2SO_4 if the desired concentration is 12.5 mg/dL?

Solution. Read the problem carefully and determine what the problem is asking for.

Total volume of solution expressed in milliliters.

Ratio and proportion is well suited for solving this type of problem. Set the problem up with respect to the considerations discussed in Appendix A concerning ratio and proportion.

$$\frac{12.5 \text{ mg}}{1 \text{ dL}} = \frac{62.5 \text{ mg}}{x \text{ dL}}$$

Solve for the unknown quantity.

$$(12.5 \text{ mg})(x \text{ dL}) = (62.5 \text{ mg})(1 \text{ dL})$$

$$\frac{(62.5 \text{ mg})(1 \text{ dL})}{12.5 \text{ mg}} = x \text{ dL} = 5.00 \text{ dL}$$

To completely satisfy the problem, convert deciliters to milliliters.

$$5.00 \text{ dL X } \frac{100 \text{ mL}}{1 \text{ dL}} = \underline{500 \text{ mL}}$$

c. **Example 3**. What is the prepared by adding 50.0 g $CaCO_3$ to a 250.0 mL flask and adjusting the volume to the mark?

Solution. Read the problem carefully and determine what the problem is asking for.

Percent (Solving this type of problem involves expressing the mass of solute per unit volume.

Since the unit volume is deciliters the first step in problem solving would be to convert the given volume in milliliters to deciliters.

$$250.0 \text{ mL X } \frac{1 \text{ dL}}{100 \text{ mL}} = 2.50 \text{ dL}$$

The most direct approach to conclude solving the problem would be to simply express the given information as weight per volume and to evaluate the numerical data as follows.

$$\frac{50.0 \text{ g}}{2.50 \text{ dL}} = \underline{20.0 \text{ g/dL}} \text{ or equivalently, } \underline{20.0}$$

Ratio and proportion could also be used to determine the concentration of the solution obtaining the same results.

$$\frac{50.0 \text{ g}}{2.50 \text{ dL}} = \frac{x \text{ g}}{1 \text{ dL}}$$

$$(50.0 \text{ g})(1 \text{ dL}) = (2.50 \text{ dL})(x \text{ g})$$

$$x \text{ g} = \frac{(50.0 \text{ g})(1 \text{ dL})}{2.50 \text{ dL}} = 20.0 \text{ g}$$

Substituting for x in the original expression yields the desired results.

$$\frac{x \text{ g}}{1 \text{ dL}} = \frac{20.0 \text{ g}}{1 \text{ dL}} = \underline{20.0 \text{ g/dL}}$$

2-29. HYDRATES

Some salts come in several forms, such as the anhydrous form (no water) and in the form of one or more hydrates. In a hydrate, a number of water molecules are attached to each molecule of salt. The water that is attached to the salt contributes to the molecular weight. Thus, if we were to weigh out equal amounts of a desired chemical and one of its hydrates, the hydrate would not yield as much of the desired chemical, per unit weight, as the anhydrous form, since some of the weight is attributable to the water molecules. In some cases, the prescribed form of a salt may not be available for the preparation of a solution. One must be able to determine how much of the available form is equivalent to the quantity of the form prescribed. To do this, we must first determine the amount of the prescribed form that is needed. Then, using the molecular weights of both substances involved, use ratio and proportion to determine the amount of available form needed.

a. **Example 1**. A procedure requires that 100 mL of a 10.0 solution be prepared. Only $CuSO_4.H_2O$ is available. How much of the hydrate is needed to prepare this solution?

Solution. Read the problem carefully and determine the desired quantity.

Grams of $CuSO_4.H_2O$.

Calculate the amount of the anhydrous form needed.

$$100\text{ mL } \times \frac{1\text{ dL}}{100\text{ mL}} \times \frac{10.0\text{ g}}{1\text{ dL}} = 10.0\text{ g}$$

Calculate the gram molecular weight of each substance. (**NOTE:** You will learn more about gram molecular weights and moles in Lesson 3.)

(1) List each element in a column.

(2) Write the atomic weight of each element. (See Appendix C.)

(3) Multiply the weight of each element by the number of that element in the chemical formula. For example, if there are 2 atoms of hydrogen present in the formula, multiply the weight of hydrogen by two.

(4) Sum the weights to obtain the total weight, and express the total weight in grams/mole.

$CuSO_4$ (anhydrous)

$$\begin{array}{lll}
\text{Cu} & 63.5 \times 1 = & 63.5 \\
\text{S} & 32.1 \times 1 = & 32.1 \\
\text{O} & 6.0 \times 4 = & \underline{+\ 64.0} \\
 & & 159.6\text{ g/mole}
\end{array}$$

$CuSO_4{\cdot}H_2O$ (hydrate)

$$\begin{array}{lll}
\text{Cu} & 63.5 \times 1 = & 63.5 \\
\text{S} & 32.1 \times 1 = & 32.1 \\
\text{O} & 6.0 \times 4 = & 64.0 \\
\text{H} & 1.0 \times 2 = & 2.0 \\
\text{O} & 6.0 \times 1 = & \underline{+\ 16.0} \\
 & & 177.6\text{ g/mol}
\end{array}$$

If the amount of anhydrous salt needed is now multiplied by the ratio of the molecular weights of the hydrate to anhydrate the desired results will be obtained.

$$10.0\text{ g } CuSO_4 \times \frac{177.6\text{ g/mol } CuSO_4{\cdot}H_2O}{159.6\text{ g/mol } CuSO_4} = 11.1\text{ g CuSO4.H2O}$$

Ratio and proportion could also be used to solve this type of problem.

$$\frac{159.6 \text{ g/mol anhydrous}}{177.6 \text{ g/mol hydrate}} = \frac{10.0 \text{ g anhydrous}}{x \text{ g hydrate}}$$

$$\frac{(177.6 \text{ g/mol hydrate})(10.0 \text{ g anhydrous})}{159.6 \text{ g/mol anhydrous}} = \underline{11.1 \text{ g}}$$

NOTE: Notice in this example that you needed a larger weight of the hydrated form to prepare a solution of equal strength. This is because a hydrated molecule weighs more than its anhydrous counterpart. The key is that the solution using a hydrate instead of the prescribed form will still react in the same manner in solution if it is prepared properly. Therefore, care must be taken when performing the calculation to ensure the proper amount of hydrate is used.

b. **Example 2**. How much $Fe_2(SO_4)_3 \cdot 4H_2O$ is needed to prepare 500.0 mL of a 200.0 mg/dL $Fe_2(SO_4)_3$ solution?

Solution. Read the problem carefully and determine the unknown quantity.

Grams of $Fe_2(SO_4)_3 \cdot 4H_2O$.

Calculate the amount of the anhydrous form needed.

$$500.0 \text{ mL} \times \frac{1 \text{ dL}}{100 \text{ mL}} \times \frac{200.0 \text{ mg}}{1 \text{ dL}} = 1000 \text{ mg}$$

Calculate the gram molecular weight of each substance. (See Appendix C.)

$Fe_2(SO_4)_3$

Fe 55.8 X 2 = 111.6
S 32.1 X 3 = 96.3
O 16.0 X 12 = $\underline{+ 192.0}$
399.9 g/mol

$Fe_2(SO_4)_3 \cdot 4H_2O$

Fe 55.8 X 2 = 111.6
S 32.1 X 3 = 96.3
O 16.0 X 12 = 192.0
H 1.0 X 8 = 8.0
O 16.0 X 4 = $\underline{+ 64.0}$
471.9 g/mol

Use ratio and proportion to determine the amount of hydrate needed.

$$\frac{399.9 \text{ g/mol anhydrous}}{471.9 \text{ g/mol hydrate}} = \frac{1000 \text{ mg anhydrous}}{x \text{ mg hydrate}}$$

$$x \text{ mg hydrate} = \frac{(1000 \text{ mg anhydrous})\ (471.9 \text{ g/mol hydrate})}{399.9 \text{ g/mol anhydrous}}$$

x mg hydrate = <u>1180 mg</u>

c. **Example 3**. How many grams of $CaCl_2 \cdot 2H_2O$ are needed to make 1.00 liter of a 20.0 mg/dL $CaCl_2$ solution?

Solution. Read the problem carefully and select the unknown quantity.

Grams of $CaCl_2 \cdot 2H_2O$.

Calculate the amount of the anhydrous form needed.

Calculate the amount of the anhydrous form needed

$$1.00 \text{ L} \times \frac{10 \text{ dL}}{1 \text{ L}} \times \frac{20.0 \text{ mg}}{1 \text{ dL}} = 200 \text{ mg}$$

Calculate the gram molecular weight of each substance. (See Appendix C.)

$CaCl_2$

Ca	40.1 X 1 =	40.1
Cl	35.5 X 2 =	+ 71.0
		111.1 g/mol

$CaCl_2 \cdot 2H_2O$

Ca	40.1 X 1 =	40.1
Cl	35.5 X 2 =	71.0
H	1.0 X 4 =	4.0
O	16.0 X 2 =	+ 32.0
		147.1 g/mol

Use ratio and proportion to determine the amount of hydrate needed.

$$\frac{111.1 \text{ g/mol anhydrous}}{147.1 \text{ g/mol hydrate}} = \frac{0.200 \text{ g anhydrous}}{x \text{ g hydrate}}$$

$$x \text{ g hydrate} = \frac{(147.1 \text{ g/mol hydrate})\ (0.200 \text{ g anhydrous})}{111.1 \text{ g/mol anhydrous}}$$

x g hydrate = $\underline{0.265 \text{ g}}$

2-30. VOLUME/VOLUME PERCENT SOLUTION PROBLEMS

When a solution has a liquid solute in a liquid solvent, percent concentration is expressed as volume per unit volume (v/v). When a number expressing a liquid solute in a liquid solvent is followed by a percent symbol (%) then the concentration is a v/v concentration.

a. **Problem Solving**. Methods developed earlier for solving weight per volume solutions may be applied to volume per volume solutions.

b. **Example 1**. How much alcohol is required to make 100 mL of a 20.0 mL/dL alcohol solution?

Solution. Read the problem carefully and determine the unknown quantity.

Milliters of alcohol.

Express the volume in deciliters.

$$100 \text{ mL X } \frac{1 \text{ dL}}{100 \text{ mL}} = 1.00 \text{ dL}$$

Multiply the volume expressed in deciliters times the percent concentration to determine the volume of solute needed to prepare the solution.

$$1.00 \text{ dL X } \frac{20.0 \text{ mL}}{\text{dL}} = \underline{20.0 \text{ mL}}$$

c. **Example 2**. What is the concentration of a solution that was prepared by adding 75 mL of alcohol to 175 mL of water?

Solution. Read the problem carefully and determine the unknown quantity.

Percent concentration.

Solving this type of problem involves expressing the volume of solute per unit volume.

Determine the total volume of solution.

175 mL water
$\underline{\ \ 75}$ mL alcohol
250 mL total volume

Since the unit volume is deciliters, convert the given volume in milliliters to deciliters.

$$250.0 \text{ mL} \times \frac{1 \text{ dL}}{100 \text{ mL}} = 2.50 \text{ dL}$$

To conclude solving the problem express the volume of solute per unit volume and evaluate the numerical data as follows.

$$\frac{75 \text{ mL}}{2.50 \text{ dL}} = \underline{30 \text{ mL/dL}} \text{ or equivalently, } \underline{30}$$

2-31. WEIGHT/WEIGHT PERCENT SOLUTIONS

This type, weight/weight percent solutions (w/w), of percent solution is usually expressed as (w/w), where "w" denotes weight (usually grams) in both cases. An example of a correct designation for this type of solution is as follows: 10 g/100 g (w/w), which indicates to the technician that there are 10 grams of solute for every 100 grams total solution. The (w/w) denotes that the solution is a "weight in weight" percent solution. This type of solution is rarely if ever prepared in the clinical laboratory since it is easier to measure volumes of liquids rather than weigh the liquid on an analytical balance. One must be familiar with w/w solutions because the common laboratory acids and bases are manufactured on a weight-in-weight basis. For more discussion on this, please refer to Lesson 8, Concentrated Acids and Bases.

2-32. PERCENT COMPOSITION PROBLEMS

At times it may be necessary to prepare a percent solution that requires only part of the entire molecule to be considered in the preparation of the solution. For example, you may wish to prepare a 2.5 g/dL sodium (Na^+) solution from NaCl. Since the entire molecule is used in the preparation of the solution, ratio and proportion will have to be used, based on the molecular weight of the entire molecule versus the molecular weight of the portion of the molecule being considered.

a. **Example 1**. How much NaCl is required to prepare 500 mL of a 2.5 g/dL sodium solution?

Solution. Read the problem carefully and determine the desired quantity.

Grams of NaCl.

Calculate the amount of sodium needed.

$$500 \text{ mL} \times \frac{1 \text{ dL}}{100 \text{ mL}} \times \frac{2.5 \text{ g}}{1 \text{ dL}} = 12.5 \text{ g}$$

Calculate the gram molecular weight of each substance. (See Appendix C.)

$$\begin{array}{lr} \text{Na} & 23.0 \\ \text{Cl} & +\ \underline{35.5} \\ & 58.5 \text{ g/mol} \end{array} \qquad \text{Na } 23.0 \text{ g/mol}$$

Use ratio and proportion to determine the amount of available form needed.

$$\frac{23.0 \text{ g/mol Na}}{58.5 \text{ g/mol NaCl}} = \frac{12.5 \text{ g Na}}{x \text{ g NaCl}}$$

$$x \text{ g NaCl} = \frac{(12.5 \text{ g Na})\ (58.5 \text{ g/mol NaCl})}{23.0 \text{ g/mol Na}}$$

x g NaCl = $\underline{31.8 \text{ g}}$

b. **Example 2**. How many milligrams of Na_2CO_3 are needed to prepare 100 mL of a 10.0 mg/dL sodium standard?

Solution. Read the problem carefully and determine the unknown quantity.

Milligrams of Na_2CO_3.

Calculate the amount of the sodium needed.

$$100 \text{ mL X } \frac{1 \text{ dL}}{100 \text{ mL}} \text{ X } \frac{10.0 \text{ mg}}{1 \text{ dL}} = 10.0 \text{ mg}$$

Calculate the gram molecular weight of each substance.

$$\begin{array}{llr} \text{Na} & 23.0 \text{ X } 2 = & 46.0 \\ \text{C} & 12.0 \text{ X } 1 = & 12.0 \\ \text{O} & 16.0 \text{ X } 3 = & \underline{+\ 48.0} \\ & & 106.0 \text{ g/mol} \end{array} \qquad \begin{array}{llr} \text{Na} & 23.0 \text{ X } 2 = & \underline{46.0} \\ & & 46.0 \text{ g/mol} \end{array}$$

Use ratio and proportion to determine the amount of available form needed.

$$\frac{46.0 \text{ g/mol Na}}{106.0 \text{ g/mol } Na_2CO_3} = \frac{10.0 \text{ mg Na}}{x \text{ mg } Na_2CO_3}$$

$$x \text{ mg } Na_2CO_3 = \frac{(106.0 \text{ g/mol NaCl})\ (10.0 \text{ mg Na})}{46.0 \text{ g/mol Na}}$$

$$x \text{ mg } Na_2CO_3 = \underline{23.0 \text{ mg}}$$

2-33. ADVANTAGES AND DISADVANTAGES OF PERCENT SOLUTIONS

The main advantage of percent solutions is the ease in calculations. There are, however, two distinct disadvantages associated with percent solutions: (1) the technician does not know the number of molecules or ions that are present in the solution, and (2) the technician has no idea of the reacting strength of the solution.

Continue with Exercises

EXERCISES, LESSON 2

INSTRUCTIONS: Answer the following exercises by writing the answer in the space provided at the end of the question. After you have completed all the exercises, turn to "Solutions to Exercises" at the end of the lesson and check your answers. For each exercise answered incorrectly, reread the material referenced with the solution.

1. Which basic type of water is used for general laboratory procedures and has a minimum 0.5 specific resistance of megohoms?

____

2. Which solution is not homogenous throughout?

____

3. What amount of Na2SO4 is needed to make 2000 mL of a 30.0solution?

____

4. How many milligrams of NaCl are present in 5.0 mL of a 20.0 mg/dL NaCl solution?

____

5. How much HCl is needed to prepare 1.0 L of a 0.50 mL/dL HCl solution?

____

6. How many mL of water must be added to 30.0 mL of toluene in order to prepare a 10.0 mL/dL toluene solution?

____

7. How many mL of 5.0acid?

____

8. A solution was prepared by adding 920 mg of KCl to a 250-mL flask. The flask was then filled to the mark with water. What is the percent concentration of the solution?

____

9. How many grams of Mg3(PO4)2 are needed to prepare 500 mL of a 4.0 mg/dL Mg standard?

____

10. What amount of Na2SO4.4H2O is needed to make 2.0 liters of a 25.0 mg/dL Na2SO4 solution?

____

11. How many grams of MgCl2.8H2O are needed to make 250 mL of a 3.00 g/dL MgCl2 solution?

____

12. You are asked to make the following solutions. How much solute would you need for each?

a. 600 mL of a 2.0

____

b. 1900 mL of a 0.80

____

c. 770 mL of a 0.60 g/dL solution.

____

d. 16.0 mg/dL solution of 250 mL volume.

____

e. 140 mL of a 3.0 mg/dL strength.

____

f. 300 mL of 30.0

____

13. How much Fe2(SO4)3 is needed to make 5000 mL of a 0.500 g/dL iron standard?

____

14. How much KCl is needed to prepare 250 mL of a 4.5 g/dL potassium standard?

____

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 2

1. Type II water (paras 2-11, 2-13)

2. Colloidal solution (para 2-22)

3. 600 g (paras 2-27, 2-28)

4. 1.0 mg (paras 2-27, 2-28)

5. 5.0 mL (paras 2-27, 2-30)

6. 270.0 mL (para 2-30)

7. 800.0 mL (paras 2-27, 2-28)

NOTE: First, calculate total volume (300 mL). This includes both the water and the toluene. Subtract the amount of solute (30 mL) to get the volume of water needed.

8. .368

9. .072 g (para 2-32)

10. 0.75 g (para 2-29)

11. 18.8 g (para 2-29)

12. a. 12.0 g (paras 2-27, 2-28)
    b. 15.2 g (paras 2-27, 2-28)
    c. 4.6 g (paras 2-27, 2-28)
    d. 40.0 mg (paras 2-27, 2-28)
    e. 4.2 mg (paras 2-27, 2-28)
    f. 390.0 mL (para 2-30)

13. 89.6 g (para 2-32)

14. 21 g (para 2-32)

NOTE: When preparing a solution in which acid is diluted, always fill the volumetric flask approximately half full with deionized water, then slowly add the acid. This is an exothermic reaction (heat is produced), and the solution must be cooled to room temperature before adjusting the volume to the mark.

End of Lesson 2

LESSON ASSIGNMENT

LESSON 3	Molar Solutions.
TEXT ASSIGNMENT	Paragraphs 3-1 through 3-11.
LESSON OBJECTIVES	After completing this lesson, you should be able to:

3-1. Calculate the weight of solute in a molar solution.

3-2. Calculate the number of moles in a compound.

3-3. Calculate the mol/L or mmol/L of a solute when the weight/volume of the concentration is known.

3-4. Calculate the solution volume of a molar concentration when the weight of the solute is known.

SUGGESTION After studying the assignment, complete the exercises at the end of this lesson. These exercises will help you to achieve the lesson objectives.

LESSON 3

MOLAR SOLUTIONS

Section I. MOLE CONCEPT

3-1. DISCUSSION

In a chemical reaction, atoms and molecules are either combined or separated during the reaction. In other words, chemical reactions take place at the level of the atoms and molecules of the reactants.

a. A method that would allow the medical laboratory specialist to know the relative number of reactant particles involved in a chemical reaction would be useful. The concept of the mole and molarity provides such a method.

b. One mole of any substance will contain approximately 6.02×10^{23} particles (Avogadro's number). A mole of a particular substance indicates the number of grams proportional to the atomic or molecular weight of the substance. This weight is often referred to as its gram molecular weight.

c. The importance of the mole concept stems from the fact that a mole of any given element or compound will yield Avogadro's number of particles. If the masses of samples of two elements or compounds have the same ratio as the ratio of their atomic weights, the samples contain identical numbers of atoms or molecules.

3-2. CALCULATING GRAM MOLECULAR WEIGHT

The gram molecular weight (GMW) is the mass in grams of one mole of substance. The first step in solving for GMW is to list each element of the compound as represented by the chemical formula separately. Next, we must take into account the atomic weight of each individual atom along with the total number of each individual atom or element present. (See Appendix C for atomic weights and symbols of elements.) Finally, add all the weights of the atoms making up the formula and express the answer in grams/mole.

a. **Example 1**. What is the GMW of NaCl?

Solution.

NaCl

$$\begin{array}{ll} \text{Na} & \ 23.0 \\ \text{Cl} & \underline{+\ 35.5} \\ & \ 58.5 \text{ g/mol} \end{array}$$

b. **Example 2**. What is the GMW of H_2SO_4?

Solution. Notice in this problem that the subscripts apply to the hydrogen and oxygen atoms. In this problem, we must multiply these elements by their corresponding subscript to obtain the correct GMW for this molecule.

H_2SO_4

H 1.0 X 2 = 2.0
S 32.1 X 1 = 32.1
O 16.0 X 4 = + 64.0
98.1 g/mol

c. **Example 3**. What is the GMW of $(NH_4)_2CO_3.3H_2O$?

Solution. In this problem, the NH_4 radical is enclosed by parentheses and a subscript of two is assigned. In this situation, all elements contained within the parentheses are multiplied by two. Also, the hydrogen atom of the NH_4 radical has a subscript of four. Therefore, the total number of hydrogens present are eight. The attached water of hydration has a coefficient of three. This must be considered, and each individual atom within the water molecule must be multiplied by a coefficient of three in addition to any subscripts.

$(NH_4)_2CO_3.3H_2O$

N 14.0 X 2 = 28.0
H 1.0 X 8 = 8.0
C 12.0 X 1 = 12.0
O 16.0 X 3 = 48.0
H 1.0 X 6 = 6.0
O 16.0 X 3 = + 48.0
150.0 g/mol

3-3. MOLES AND GRAM MOLECULAR WEIGHT

Remember from the discussion on moles that a mole of any substance is the number of grams proportional to the atomic or molecular weight of the substance. If this is true, then one mole of NaCl weighs 58.5 grams. One mole of any substance has a mass equal to one gram molecular weight of that substance.

a. **Example 1**. How many moles are contained in 117.0 grams of NaCl?

Solution. Determine the mass of one gram molecular weight of the substance.

$$\begin{array}{ll} \text{Na} & \,23.0 \\ \text{Cl} & \underline{+\,35.5} \\ & 58.5 \text{ g/mol} \end{array}$$

Now use ratio and proportion to determine the unknown quantity.

$$\frac{58.5 \text{ g}}{1 \text{ mol}} = \frac{117.0 \text{ g}}{x \text{ mol}}$$

$$(58.5 \text{ g})(x \text{ mol}) = (117.0 \text{ g})(1 \text{ mol})$$

$$x \text{ mol} = \frac{(117.0 \text{ g})(1 \text{ mol})}{58.5 \text{ g}}$$

x mol = $\underline{2.00 \text{ mol}}$ There are 2.00 moles contained in 117.0 g of NaCl.

Equivalently:

$$\text{moles} = \frac{\text{the number of grams given}}{\text{GMW}}$$

b. **Example 2**. How many moles are contained in 71.0 grams of Na_2SO_4?

Solution. Calculate the GMW of the compound.

$$\begin{array}{lll} \text{Na} & 23.0 \times 2 = & 46.0 \\ \text{S} & 32.1 \times 1 = & 32.1 \\ \text{O} & 16.0 \times 4 = & \underline{+\,64.0} \\ & & 142.1 \text{ g/mol} \end{array}$$

Use ratio and proportion to determine the unknown quantity.

$$\frac{142.1 \text{ g}}{1 \text{ mol}} = \frac{71.0 \text{ g}}{x \text{ mol}}$$

$$(142.1 \text{ g})(x \text{ mol}) = (71.0 \text{ g})(1 \text{ mol})$$

$$x \text{ mol} = \frac{(71.0 \text{ g})(1 \text{ mol})}{142.1 \text{ g}}$$

x mol = 0.500 mol. There are 0.500 moles contained in 71.0 grams of Na_2SO_4

NOTE: Keep in mind that moles are not an expression of concentration. The unit mole is an expression of mass indicating the number of gram molecular weights. Take caution not to confuse mole with molarity (mol/L). When reporting moles, write the word moles, DO NOT abbreviate by using an M!!

Section II. MOLARITY AND MOLAR SOLUTIONS

3-4. DEFINITION

Molarity (mol/L) is defined as the number of moles of a substance per liter of solution.

3-5. SOLVING MOLARITY PROBLEMS

a. **Example 1**. How much Na_2SO_4 is needed to make 300 mL of a 2.00 mol/L Na_2SO_4 solution?

Solution. Notice in this problem that the wording suggests that a solution is to be prepared. If you examine the problem, you should find that the type of solution to be prepared is a molar solution. Read the problem carefully and determine the unknown quantity.

Grams of Na_2SO_4.

Regardless of the type, most molarity problems require the gram molecular weight, so calculating the gram molecular weight is a good starting point.

Na_2SO_4

$$\begin{array}{lll} \text{Na} & 23.0 \text{ X } 2 = & 46.0 \\ \text{S} & 32.1 \text{ X } 1 = & 32.1 \\ \text{O} & 16.0 \text{ X } 4 = & \underline{+\ 64.0} \\ & & 142.1 \text{ g/mol} \end{array}$$

Express the desired volume in liters.

$$300 \text{ mL X } \frac{1 \text{ L}}{1000 \text{ mL}} = 0.300 \text{ L}$$

Multiply the volume expressed in liters times the molar concentration times the gram molecular weight to determine the amount of salt, in grams, needed to prepare the solution.

$$0.300 \text{ L X } \frac{2.00 \text{ mol}}{\text{L}} \text{ X } \frac{142.1 \text{ g}}{\text{mol}} = \underline{85.3 \text{ g}}$$

b. **Example 2**. What volume of $CaCO_3$ solution can be prepared from 25.0 g of $CaCO_3$ if a 0.50 mol/L $CaCO_3$ is desired?

Solution. Read the problem carefully and determine the desired quantity.

Volume of $CaCO_3$ solution.

Calculate the GMW of the compound.

$CaCO_3$

$$\begin{array}{lrl} \text{Ca} & 40.1 \times 1 = & 40.1 \\ \text{C} & 12.0 \times 1 = & 12.0 \\ \text{O} & 16.0 \times 3 = & \underline{+\ 48.0} \\ & & 100.1 \text{ g/mol} \end{array}$$

Determine the number of moles contained in 25.0 grams of solute.

$$25.0 \text{ g} \times \frac{1 \text{ mol}}{100.1 \text{ g}} = 0.250 \text{ mol}$$

Use ratio and proportion to determine the amount of solution that can be prepared.

$$\frac{0.50 \text{ mol}}{1 \text{ L}} = \frac{0.250 \text{ mol}}{x \text{ L}}$$

$$(0.50 \text{ mol})(x \text{ L}) = (0.250 \text{ mol})(1 \text{ L})$$

$$x \text{ L} = \frac{(0.250 \text{ mol})(1 \text{ L})}{0.50 \text{ mol}} = \underline{0.50 \text{ L}}$$

Section III. MILLIMOLES PER LITER

3-6. INTRODUCTION

Another expression with which one should be familiar is millimoles per liter (mmol/L). It is a variation of the mole/liter concentration unit.

3-7. DISCUSSION

a. A millimole = 0.001 mole and 1 mole = 1000 millimoles.

b. A millimole has the same numerical value as one gram molecular weight, however, it is expressed in mg/mmol.

c. To convert mol/L to mmol/L and vice versa, multiply the concentration by the conversion factor 1000 mmol/1 mol as appropriate.

Example: (0.25 mol/L)(1000 mmol/1 mol) = 250 mmol/L

d. To convert mmol/L to mol/L, multiply mmol/L by (1 mol/1000 mmol).

Example: (2000 mmol/L)(1 mol/1000 mmol) = 2.000 mol/L

e. Use the following formula to solve millimole/liter problems:

(1) Formula. mg = (mmol/L)(mg/mmol)(L)

(2) Example. How much NaCl is needed to make 250 mL of a 300 mmol/L solution?

Solution. Read the problem carefully and select the formula that generates the desired quantity.

mg = (mmol/L)(mg/mmol)(L)

Determine the millimolar weight of the compound.

NaCl

Na 23.0
Cl + 35.5
58.5 mg/mmol

Substitute the given information.

mg = (300 mmol/L)(58.5 mg/mmol)(250 mL)

Make any necessary conversions.

mg = (300 mmol/L)(58.5 mg/mmol)(0.250 L)

Evaluate the expression to determine the unknown quantity.

mg = 4390 mg

3-8. SOLVING MILLIMOLE PER LITER PROBLEMS

Calculations involving millimole per liter solution are very similar to the calculations involving simple mole per liter solutions.

a. **Example 1**. What is the mmol/L concentration of a 0.25 mol/L NaOH solution?

Solution. Read the problem carefully and determine the desired unit of concentration.

mmol/L.

Multiply the given mol/L concentration times the appropriate conversion factor.

$$\frac{0.25 \text{ mol}}{\text{L}} \times \frac{1000 \text{ mmol}}{\text{mol}} = \underline{250 \text{ mmol/L}}$$

b. **Example 2**. What is the mol/L concentration of a 2000 mmol/L KCl solution?

Solution. Read the problem carefully and determine the desired unit of concentration.

mol/L.

Multiply the given mmol/L concentration times the appropriate conversion factor.

$$\frac{2000 \text{ mmol}}{\text{L}} \times \frac{1 \text{ mol}}{1000 \text{ mmol}} = \underline{2.000 \text{ mol/L}}$$

c. **Example 3**. How much NaCl is needed to make 250 mL of a 300 mmol/L solution?

Solution**.** Read the problem carefully and determine the desired quantity.

Milligrams of NaCl.

Determine the millimolar weight of the compound.

NaCl

Na	23.0
Cl	<u>+ 35.5</u>
	58.5 mg/mmol

Express the desired volume of solution in liters

$$250 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.250 \text{ L}$$

Multiply the desired volume in liters times the millimolar concentration times the millimolar weight to determine the amount of salt needed to prepare the solution.

$$0.250\text{ L} \times \frac{300\text{ mmol}}{\text{L}} \times \frac{58.5\text{ mg}}{\text{mmol}} = \underline{4390\text{ mg}}$$

Section IV. VARIATIONS OF MOLARITY PROBLEMS

3-9. PROBLEMS INVOLVING HYDRATES

Since molarity is based on the number of moles (Avogadro's number) per liter of solution, then one mole per liter of $CuSO_4$ will contain the same number particles as one mole per liter of $CuSO_4 \cdot 5H_2O$. When calculating molarity problems involving hydrates, use the gram molecular weight of the substance being used (weighed) in the preparation of the solution.

Example. How much $CuSO_4 \cdot 5H_2O$ is required to prepare 250 mL of a 2.00 mol/L $CuSO_4$ solution?

Solution. Read the problem carefully and determine the desired quantity.

Grams of $CuSO_4 \cdot 5H_2O$.

Determine the GMW of the <u>available compound</u>.

$CuSO_4 \cdot 5H_2O$

Cu 63.5 X 1 = 63.5
S 32.1 X 1 = 32.1
O 6.0 X 4 = 64.0
H 1.0 X 10 = 10.0
O 16.0 X 5 = + 80.0
249.6 g/mol

Express the desired volume in liters.

$$250\text{ mL} \times \frac{1\text{ L}}{1000\text{ mL}} = 0.250\text{ L}$$

Multiply the volume expressed in liters times the molar concentration times the gram molecular weight to determine the amount of hydrate salt, in grams, needed to prepare the solution.

$$0.250\text{ L} \times \frac{2.00\text{ mol}}{\text{L}} = \frac{249.6\text{ g}}{\text{mol}} = \underline{125\text{ g}}$$

3-10. PROBLEMS INVOLVING ONLY PART OF THE MOLECULE

When preparing solutions using compounds formed by ionic bonding, realize that these molecules will ionize when in solution. For example, one mole of Na_2SO_4 will ionize to two moles of sodium and one mole of sulfate radical. If we were to prepare a one mol/L sodium solution using Na_2SO_4, then only 0.5 moles of sodium sulfate would be needed for every liter of solution since two moles of sodium are contained in one mole of sodium sulfate. When solving this type of problem use an appropriate factor to account for the moles of desired substance per mole of compound.

Example**.** How much Na_2SO_4 is needed to prepare 2.00 liters of 1.50 mol/L sodium solution?

Solution. Read the problem carefully and determine the desired quantity.

Grams of Na_2SO_4.

Calculate the GMW of the compound.

Na_2SO_4

Na	23.0 X 2 =	46.0
S	32.1 X 1 =	32.1
O	16.0 X 4 =	+ 64.0
		142.1 g/mol

Multiply the volume expressed in liters times the molar concentration times the gram molecular weight times the ratio of ionized sodium per mole of compound to determine the amount of salt, in grams, needed to prepare the solution.

$$2.00\text{ L} \times \frac{1.50\text{ mol Na}}{\text{L}} \times \frac{142.1\text{ g}}{\text{mol}} \times \frac{1\text{ mol } Na_2SO_4}{2\text{ mol Na}} = \underline{213\text{ g}}\ Na_2SO_4$$

3-11. ADVANTAGES AND DISADVANTAGES OF MOLAR SOLUTIONS

a. Advantage: Number of particles in solution known.

b. Disadvantage: Reactive strength of solution unknown.

Continue with Exercises

EXERCISES, LESSON 3

INSTRUCTIONS: Answer the following exercises by writing the answer in the space provided at the end of the question.

After you have completed all the exercises, turn to "Solutions to Exercises" at the end of the lesson and check your answers. For each exercise answered incorrectly, reread the material referenced with the solution.

1. How much $CaCl_2$ is needed to make 200 mL of a 6.00 mol/L $CaCl_2$ solution?

   ____

2. What weight of NaOH would be required to prepare 3000 mL of a 1.50 mol/L NaOH solution?

   ____

3. In a solution containing 3000 mg of NaCl in 250 mL, what is the mol/L concentration?

   ____

4. How many moles of $CaCl_2$ are contained in 333.0 g of $CaCl_2$?

   ____

5. What is the mmol/L concentration of a KOH solution prepared by adding 14.0 g of KOH to a 500 mL volumetric flask and adjusting to the mark with deionized water?

   ____

6. How much 250 mmol/L solution can be prepared from 149 mg of KCl?

____

7. A solution of NaCl contains 300 grams per liter. What is the mol/L concentration of the solution?

____

8. What amount of $Na_2SO_4.4H_2O$ is needed to make 750 mL of a 1.25 mol/L solution of Na_2SO_4?

____

9. How much $CuSO_4.5H_2O$ is needed to prepare 250 mL of a 2.00 mol/L $CuSO_4$ solution?

____

10. How much $NaNO_3$ is needed to make 200 mL of a 30.0 mmol/L $NaNO_3$ solution?

____

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 3

1. 133 g (para 3-5)
2. 180 g (para 3-5)
3. 0.205 mol/L (para 3-5)
4. 2.997 moles (para 3-3)
5. 499 mmol/L (para 3-7)
6. 7.99 mL (para 3-7)
7. 5.13 mol/L (paras 3—3, 3-5)
8. 201 g (para 3-9)
9. 125 g (para 3-9)
10. 510 mg (para 3-7)

End of Lesson 3

LESSON ASSIGNMENT

LESSON 4	Equivalent Solutions.
TEXT ASSIGNMENT	Paragraphs 4-1 through 4-11.
LESSON OBJECTIVE	After completing this lesson, you should be able to: Calculate the gram equivalent weight, normality of a solution, milliequivalent per liter problems and hydrate, with variations, equations.
SUGGESTION	After completing the assignment, complete the exercises of this lesson. These exercises will help you to achieve the lesson objective.

EQUIVALENT SOLUTIONS

Section I. EQUIVALENT WEIGHT

4-1. INTRODUCTION

a. A gram equivalent weight by definition is the mass of substance that will combine with or displace one mole of hydrogen.

b. In solutions of ionic compounds, the compound dissociates into positive or negative ions to react with other ions. An example is KOH, which separates into one K^+ ion and one OH^- ion. One mole of K^+ will replace one mole of H^+ in a chemical reaction; hence, one equivalent weight of KOH is equal to one mole of KOH. Consider now the combining ability of the SO_4^{-2} ion in sulfuric acid, H_2SO_4. One mole of the $SO_4^{\text{Ä}2}$ ion will combine with two moles of H^+. Since one gram equivalent weight of a compound or element is the mass that will combine with one mole of H^+, then one gram equivalent weight of H_2SO_{44} equals 0.5 moles (1/2 the gram molecular weight) of H_2SO_4 because two H^+ ions will combine with one SO_4^{-2} ion. Total positive ionic valence (TPIV) expresses the number of equivalents per mole of substance.

NOTE: This is true for metathesis reactions only, where ionic valence does not change.

4-2. DETERMINING THE GRAM EQUIVALENT WEIGHT

In solving for gram equivalent weight (GEW), the first step is to find the GMW of the compound and then divide by the total positive ionic valence (TPIV). The TPIV is normally determined by finding the valence of the first ion in the chemical compound. (See Appendix C for the valences of common ions.) The valence of that ion is then multiplied by any subscripts pertaining to that ion. If two or more valences are possible for the ion, simply find the valence of the negative ion. Multiply the valence by any subscripts that pertain to that ion giving a total negative ionic valence. Since the sum of the positive valence and negative valence is zero, the TPIV is numerically equal to the total negative valence.

a. **Example 1**. Find the GEW of KOH.

Solution. Determine the GMW of the compound

KOH

K	39.1
O	16.0
H	+ 1.0
	56.1 g/mol

Multiply the GMW, expressed in grams per mole, times 1/TPIV, expressed in moles per equivalent, used as an appropriate factor.

$$\frac{56.1\ \text{g}}{\text{mol}} \times \frac{\text{mol}}{1\ \text{Eq}} = \underline{56.1\ \text{g/Eq}}$$

b. **Example 2**. Find the GEW of $CaCO_3.2H_2O$.

Solution. Determine the GMW of the compound.

NOTE: Water, being a relatively neutral substance, is not considered when determining the TPIV of a hydrate. However, its weight must be considered when determining the GMW.

$CaCO_3.2H_2O$

$$\begin{array}{lll}
\text{Ca} & 40.1 \times 1 = & 40.1 \\
\text{C} & 12.0 \times 1 = & 12.0 \\
\text{O} & 16.0 \times 3 = & 48.0 \\
\text{H} & 1.0 \times 4 = & 4.0 \\
\text{O} & 16.0 \times 2 = & \underline{+\ 32.0} \\
 & & 136.1\ \text{g/mol}
\end{array}$$

Multiply the GMW, expressed in grams per mole, times the 1/TPIV, expressed in moles per equivalent, used as an appropriate factor.

$$\frac{136.1\ \text{g}}{\text{mol}} \times \frac{\text{mol}}{2\ \text{Eq}} = \underline{68.05\ \text{g/Eq}}$$

c. **Example**. Calculate the GEW of $Fe_2(SO_4)_3$.

Solution. Determine the GMW of the compound.

NOTE: The iron in ferric sulfate is ambivalent. The term ambivalent refers to the iron's ability to have more than one valence. In order to determine the TPIV of this and like compounds it is necessary to examine the charge of the anion (negative ion). The TPIV will be equal to the absolute value (integer value, excluding the sign) of the negative charge of the anion.

$Fe_2(SO_4)_3$

Fe	55.8 X 2 =	111.6
S	32.1 X 3 =	96.3
O	16.0 X 12 =	+ 192.0
		399.9 g/mol

Multiply the GMW, expressed in grams per mole, times the 1/TPIV, expressed in moles per equivalent, used as an appropriate factor.

$$\frac{399.9\text{ g}}{\text{mol}} \times \frac{\text{mol}}{6\text{ Eq}} = \underline{66.65\text{ g/Eq}}$$

4-3. EXCEPTIONS IN CALCULATING GRAM EQUIVALENT WEIGHT

The only exact method of calculating the GEW is to examine the complete reaction that the molecule will undergo. For example:

$H_3PO_4 + NaOH$ <=======> $NaH_2PO_4 + HOH$

In this reaction, only one of the hydrogen ions on the phosphoric acid molecule reacted with the sodium hydroxide. The TPIV for the acid in this example is one.

$H_3PO_4 + 2NaOH$ <=======> $Na_2HPO_4 + 2HOH$

In this example, two moles of sodium hydroxide were reacted with one mole of phosphoric acid and the TPIV is two since two of the hydrogen ions from the acid reacted.

$H_3PO_4 + 3NaOH$ <=======> $Na_3PO_4 + 3HOH$

The TPIV in this reaction is three since three hydrogen ions from the acid reacted with the three moles of sodium hydroxide. A general rule to follow is to examine the reaction that is to occur. If the products of the reaction cannot be determined by inspecting the given material, then assume complete ionization unless other information is given. In the case of phosphoric acid, if no guidance is given as to the type of reaction that the molecule will undergo, or if the degree of ionization is not given, then assume the TPIV to be three.

4-4. EQUIVALENTS

An equivalent indicates the number of grams equal to the equivalent weight of the substance. One equivalent has the mass of one gram equivalent weight. The mechanism of solving for the number of equivalents is similar to that of solving for the number of moles except GEW is used instead of GMW.

Section II. NORMAL SOLUTIONS

4-5. DEFINITION

A normal solution is defined as the number of equivalents per liter of solution. Normality is used to account for different degrees of reactivity of chemicals. One gram equivalent weight of a substance will react exactly with one gram equivalent weight of another substance. A one normal solution contains one GEW of solute in one liter of total solution. Use the following information in solving normality problems.

4-6. SOLVING NORMALITY PROBLEMS

a. **Method**. Based on the definition given above, it follows that:

(1) Normality (Eq/L) = number of equivalents per liter of solution

(2) GEW = grams per equivalent

(3) Liter = total volume of solution

b. **Example**. How much NaOH is needed to prepare 400 mL of a 10.0 Eq/L NaOH solution?

Solution. Read the problem carefully and determine the desired quantity.

Grams of NaOH.

Regardless of the type, most normality problems require the gram equivalent weight, so calculating the gram equivalent weight is a good starting point.

NaOH

$$\begin{array}{ll} \text{Na} & 23.0 \\ \text{O} & 16.0 \\ \text{H} & \underline{+\ 1.0} \\ & 40.0\ \text{g/mol} \end{array}$$

$$\frac{40.0\ \text{g}}{\text{mol}} \times \frac{\text{mol}}{1\ \text{Eq}} = 40.0\ \text{g/Eq}$$

Express the desired volume in liters.

$$400 \text{ mL} \times \frac{1\text{L}}{1000 \text{ mL}} = 0.400 \text{ L}$$

Multiply the normal concentration times the gram equivalent weight times the volume expressed in liters to determine the amount of solute needed to prepare the solution.

$$\frac{10 \text{ Eq}}{\text{L}} \times \frac{40.0 \text{ g}}{\text{Eq}} \times 0.400 \text{ L} = \underline{160 \text{ g}}$$

c. **Example**. What is the Eq/L concentration of an $AlPO_4$ solution that was prepared by adding 120.0 grams of $AlPO_4$ to a 250 mL flask and adjusting the volume to the mark?

Solution. Read the problem carefully and determine the desired quantity.

Eq/L concentration.

Determine the GEW of the compound.

$AlPO_4$

$$\begin{array}{lrr} \text{Al} & 27.0 \times 1 = & 27.0 \\ \text{P} & 31.0 \times 1 = & 31.0 \\ \text{O} & 16.0 \times 4 = & \underline{+\ 64.0} \\ & & 122.0 \text{ g/mol} \end{array}$$

$$\frac{122.0 \text{ g}}{\text{mol}} \times \frac{\text{mol}}{3 \text{ Eq}} = 40.67 \text{ g/Eq}$$

Express the grams of solute as equivalents using the appropriate conversion factor.

$$120.0 \text{ g} \times \frac{\text{Eq}}{40.67 \text{ g}} = 2.951 \text{ Eq}$$

Express the desired volume in liters.

$$250 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.250 \text{ L}$$

Concentration is an expression of the relative amounts of solute and solvent present, a ratio. Express the equivalents per liter and simplify the expression.

$$\frac{2.951 \text{ Eq}}{0.250 \text{ L}} = \underline{11.8 \text{ Eq/L}}$$

d. **Example**. What is the maximum volume of Na_2CO_3 solution that can be prepared from 106.0 grams of Na_2CO_3 if the desired concentration of solution is 0.500 Eq/L?

Solution. Read the problem carefully and determine the desired quantity.

Volume in liters.

Determine the GEW of the compound.

Na_2CO_3

Na 23.0 X 2 = 46.0
C 12.0 X 1 = 12.0
O 16.0 X 3 = <u>+ 48.0</u>
106.0 g/mol

$$\frac{106.0 \text{ g}}{\text{mol}} \times \frac{\text{mol}}{2 \text{ Eq}} = 53.00 \text{ g/Eq}$$

Calculate the grams per liter by multiplying the equivalent per liter concentration times the grams per equivalent.

$$\frac{0.500 \text{ Eq}}{\text{L}} \times \frac{53.00 \text{ g}}{\text{Eq}} = 26.5 \text{ g/L}$$

The calculated value of 26.5 g/L is now used as a factor to determine the volume of solution that can be prepared. However, the reciprocal of the value must be used as a factor. If there are 26.5 grams per liter then for every liter there are 26.5 grams of solute. This relationship is true of all factors we have used.

$$106.0 \text{ g} \times \frac{\text{L}}{26.5 \text{ g}} = \underline{4.00 \text{ L}}$$

Section III. MILLIEQUIVALENT PER LITER (mEq/L) SOLUTIONS

4-7. INTRODUCTION

Milliequivalent per liter problems are based on the same principle as normality. If we examine the definition of normality, we will find that a normal solution contains one gram equivalent weight per liter of solution. Equivalently, we can say that a normal solution is one equivalent per liter of solution (Eq/L). A milliequivalent per liter solution would, therefore, contain one milliequivalent weight per liter of solution.

4-8. DISCUSSION

a. A milliequivalent = 0.001 equivalent and 1 equivalent = 1000 milliequivalents.

b. A milliequivalent has the numerical value of one equivalent weight expressed in mg/mEq.

c. To convert Eq/L to mEq/L and vice versa multiply the concentration by the conversion factor 1000 mEq/1 Eq as appropriate.

4-9. SOLVING MILLIEQUIVALENT PER LITER PROBLEMS

a. **Example 1**. How much NaCl is needed to prepare 500 mL of a 25.0 mEq/L NaCl solution?

Solution. Read the problem carefully and determine the desired quantity.

Milligrams of NaCl.

Determine the milliequivalent weight of the compound.

NaCl

$$\begin{array}{ll} \text{Na} & 23.0 \\ \text{C} & \underline{+\ 35.5} \\ & 58.5 \text{ mg/mmol} \end{array}$$

$$\frac{58.5\text{ mg}}{\text{mmol}} \times \frac{\text{mmol}}{1\text{ mEq}} = 58.5\text{ mg/mEq}$$

Express the desired volume in liters.

$$500\text{ mL} \times \frac{1\text{ L}}{1000\text{ mL}} = 0.500\text{ L}$$

Determine the amount of salt needed using the appropriate factors.

$$\frac{25.0\text{ mEq}}{\text{L}} \times \frac{58.5\text{ mg}}{\text{mEq}} \times 0.500\text{ L} = \underline{731\text{ mg}}$$

b. **Example 2**. How much $MgSO_4$ is needed to make 100 mL of a 50.0 mEq/L $MgSO_4$ solution?

Solution. Read the problem carefully and determine the desired quantity.

Milligrams of $MgSO_4$.

Determine the milliequivalent weight of the compound.

$MgSO_4$

$$\begin{array}{lll} \text{Mg} & 24.3 \times 1 = & 24.3 \\ \text{S} & 32.1 \times 1 = & 32.1 \\ \text{O} & 16.0 \times 4 = & \underline{+\ 64.0} \\ & & 120.4\text{ mg/mmol} \end{array}$$

$$\frac{120.4\text{ mg}}{\text{mmo}} \times \frac{\text{mmol}}{2\text{ mEq}} = 60.20\text{ mg/mEq}$$

Express the desired volume in liters.

$$100\text{ mL} \times \frac{1\text{ L}}{1000\text{ mL}} = 0.100\text{ L}$$

Calculate the amount of salt needed using the appropriate factors.

$$\frac{50.0 \text{ mEq}}{\text{L}} \text{ X } \frac{60.20 \text{ mg}}{\text{mEq}} \text{ X } 0.100 \text{ L} = \underline{301 \text{ mg}}$$

Section IV. VARIATIONS OF NORMALITY PROBLEMS

4-10. PROBLEMS INVOLVING HYDRATES

Like molar solutions, normal solutions contain a known number of particles per liter of solution. An one equivalent per liter (Eq/L) solution of $CuSO_4$ will have the same concentration as a one equivalent per liter $CuSO_4.5H_2O$ solution. When solving normal problems involving hydrates, use the gram equivalent weight of the substance being weighed in the preparation of the solution.

Example.
How much $CuSO_4.5H_2O$ is required to prepare 500 mL of a 1.50 Eq/L $CuSO_4$ solution?

Solution. Read the problem carefully and determine the desired quantity.

Grams of $CuSO_4.5H_2O$.

Determine the GEW of the compound that is available (actually being weighed out).

$CuSO_4.5H_2O$

Cu	63.5 X 1 =	63.5
S	32.1 X 1 =	32.1
O	16.0 X 4 =	64.0
HOH	18.0 X 5 =	<u>+ 90.0</u>
		249.6 g/mol

NOTE: When determining the TPIV of hydrates, the water (H_2O) molecules are not considered. Water has a low degree of reactivity. Also, note that copper is ambivalent, and we must consider the charge of the anion (-) in order to correctly determine the TPIV.

$$\frac{249.6 \text{ g}}{\text{mol}} \text{ X } \frac{\text{mol}}{2 \text{ Eq}} = 124.8 \text{ g/Eq}$$

Now determine the grams of hydrate needed using the appropriate factors.

$$\frac{1.50\ \text{Eq}}{\text{L}} \times \frac{24.8\ \text{g}}{\text{Eq}} \times 0.500\ \text{L} = \underline{93.6\ \text{g}}$$

4-11. PROBLEMS THAT CONSIDER ONLY PART OF THE MOLECULE

When preparing solutions using compounds where only part of the molecule will be involved in the reaction, a factor must be used to account for the reactivity of the ion of interest per mole of compound.

NOTE: TPIV expresses the number of equivalents per mole of substance.

Example**.** How much Na_2HPO_4 is required to prepare 2.00 liters of a 1.00 Eq/L sodium standard?

Solution. Read the problem carefully and determine the desired quantity.

Grams of Na_2HPO_4.

Calculate the GEW of the compound using the TPIV of the ion of interest.

Na_2HPO_4

Na 23.0 X 2 = 46.0
H 1.0 X 1 = 1.0
P 31.0 X 1 = 31.0
O 16.0 X 4 = <u>+ 64.0</u>
142.0 g/mol Na_2HPO_4

$$\frac{142.0\ \text{g}\ Na_2HPO_4}{\text{mol}\ Na_2HPO_4} \times \frac{\text{mol}\ Na_2HPO_4}{2\ \text{Eq Na}} = 71.00\ \text{g}\ Na_2HPO_4/\text{Eq Na}$$

Determine the amount of salt (hydrate) needed in the usual manner using the appropriate factors.

$$\frac{1.00\ \text{Eq Na}}{\text{L}} \times \frac{71.00\ \text{g}\ Na_2HPO_4}{\text{Eq Na}} \times 2.00\ \text{L} = \underline{142\ \text{g}\ Na_2HPO_4}$$

Continue with Exercises

EXERCISES, LESSON 4

INSTRUCTIONS: Answer the following exercises by writing the answer in the space provided at the end of the question.

After you have completed all the exercises, turn to "Solutions to Exercises" at the end of the lesson and check your answers. For each exercise answered incorrectly, reread the material referenced with the solution.

1. What is the GEW of K_2SO_4?

   ____

2. What weight of $Ba(OH)_2$ is needed to make 1500 mL of a 0.500 Eq/L $Ba(OH)_2$ solution?

   ____

3. What amount of $CuSO_4$ solution can be prepared from 80.1 grams of compound, if the desired concentration is 1.50 Eq/L?

   ____

4. A solution contains 6.00 gram equivalents per liter of solution. What is the normality of the solution?

   ____

5. What is the Eq/L concentration of a 6 mEq/L solution?

   ____

6. How many milligrams are required to make 250 mL of a 75.0 mEq/L Na_2CO_3 solution?

   ____

7. A KCl solution was prepared by adding 745 mg and adjusting the volume to 1.00 liter. What is the mEq/L concentration of the solution?

____

8. What is the Eq/L concentration of 40.0 g of NaOH in 400 mL of solution?

____

9. How many grams of $LiNO_3$ are needed to make 1000 mL of a 300 mEq/L $LiNO_3$ solution?

____

10. What amount of $Fe_3(PO_4)_2.4H_2O$ is required to make 500 mL of a 2.50 Eq/L $Fe_3(PO_4)_2$ solution?

____

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 4

1. 87.15 g (para 4-2)
2. 64.2 g (para 4-5)
3. 0.669 L (para 4-5)
4. 6 Eq/L = 6 Normal (para 4-5)
5. 0.006 Eq/L (para 4-6)
6. 994 mg (para 4-6)
7. 9.99 mEq/L (para 4-6)
8. 2.50 Eq/L (para 4-5)
9. 20.7 g (para 4-6)
10. 89.5 g (para 4-7a)

End of Lesson 4

LESSON ASSIGNMENT

LESSON 5	Conversion of Concentration Units.
TEXT ASSIGNMENT	Paragraphs 5-1 through 5-4.
LESSON OBJECTIVES	After completing this lesson, you should be able to: 5-1. Calculate conversions from mol/L to g/dL and from g/dL to mol/L. 5-2. Calculate conversions from Eq/L to g/dL and from g/dL to Eq/L. 5-3. Calculate conversions from mol/L to Eq/L and from Eq/L to mol/L.
SUGGESTION	After completing the assignment, complete the exercises of this lesson. These exercises will help you to achieve the lesson objectives.

LESSON 5

CONVERSION OF CONCENTRATION UNITS

5-1. DISCUSSION

At times, units of concentration will be given, but another unit of concentration will be required to perform calculations. In this lesson, we will discuss methods for conversion from one unit of concentration to another.

5-2. CONVERSION FROM mol/L TO g/dL AND FROM g/dL TO mol/L

Before discussing the methods for converting the above units, let's review the definition for each unit.

EQUIVALENT WAYS OF EXPRESSING
CONCENTRATION (NO NUMERICAL CONVERSION NECESSARY)

g/dL = parts of solute per 100 parts

g/L = parts of solute per 1000 parts

molarity = moles per liter

The g/dL concentration is defined as parts of solute per 100 parts total solution. The morlarity of a solution is defined as the number of moles per liter of solution (mol/L).

CONVERSIONS				
	CONCENTRATION EXPRESSED IN	BY	IN UNITS	TO GET CONCENTRATION EXPRESSED IN
Multiply	mol/L	GMW	g/mol	g/L or parts solute per 1000 parts
Multiply	g/dL	10	d/L/L	g/L or parts solute per 1000 parts
Multiply	g/dL	$\frac{10}{\text{GMW}}$	$\frac{\text{dL}}{\text{L}} \cdot \frac{\text{mol}}{\text{g}}$	mol/L

First, let us consider the mol/L concentration. Note that mol/L times the gram molecular weight (g/mol) yields g/L (mol/L X g/mol = g/L). Thus, g/L can also be referred to as parts solute per 1000 total solution. And the g/dL concentration times 10 dL/1 L yields parts per 1000. Equivalently, the g/dL concentration times 10 equals the molar concentration times the gram molecular weight.

a. **Example 1**. What is the molarity of a 4.0 g/dL NaCl solution?

Solution. Read the problem carefully and determine the desired unit of concentration.

Molarity (mol/L).

Comparable to problems encountered before that involved molarity, calculating the gram molecular weight is a good place to start.

NaCl

$$\begin{array}{ll} \text{Na} & \ 23.0 \\ \text{Cl} & \underline{+\ 35.5} \\ & \ 58.5\text{ g/mol} \end{array}$$

Now that the number of grams per mole has been determined, use it as an appropriate factor along with the conversion factor that will convert deciliters to liters.

$$\frac{4.0\text{ g}}{\text{dL}} \text{ X } \frac{1\text{ mol}}{58.5\text{ g}} \text{ X } \frac{10\text{ dL}}{1\text{ L}} = \underline{0.68\text{ mol/L}}$$

b. **Example 2**. You are directed to prepare 500 mL of a 5.0 mol/L $FeSO_4$ solution. What is the percent concentration of the solution?

Solution. Read the problem carefully and determine the desired unit of concentration.

Percent concentration (%).

Calculate the GMW of the compound.

$FeSO_4$

$$\begin{array}{llr} \text{Fe} & 55.8 \text{ X } 1 = & 55.8 \\ \text{S} & 32.1 \text{ X } 1 = & 32.1 \\ \text{O} & 16.0 \text{ X } 4 = & \underline{+\ \ 64.0} \\ & & 151.9\text{ g/mol} \end{array}$$

The given volume is a distracter. It is not needed to solve this problem.

Use the GMW and an appropriate volume conversion factor to express the concentration in g/dL.

$$\frac{5.0\text{ mol}}{\text{L}} \text{ X } \frac{151.9\text{ g}}{\text{mol}} \text{ X } \frac{1\text{ L}}{10\text{ dL}} = \underline{76\text{ g/dL}}$$

5-3. CONVERSION FROM Eq/L TO g/dL AND FROM g/dL TO Eq/L

In solving this type of conversion problem, the method is similar to the one used to solve for mol/L to g/dL. The g/dL concentration is, as previously stated, defined as parts of solute per 100 parts of total solution. The normality of a solution is defined as the number of equivalents per liter of solution (Eq/L).

a. **Example 1**. What is the g/dL concentration of a 1.5 Eq/L NaOH solution?

Solution. Read the problem carefully and determine the desired unit of concentration.

Percent concentration (g/dL).

Comparable to problems encountered before that involved normality, calculating the gram equivalent weight is a good place to start.

NaOH

Na	23.0
O	16.0
H	<u>+ 1.0</u>
	40.0 g/mol

$$\frac{40.0\text{ g}}{\text{mol}} \text{ X } \frac{1\text{ mol}}{1\text{ Eq}} = 40.0\text{ g/Eq}$$

Use the GEW and the appropriate volume conversion factor to express the concentration as g/dL.

$$\frac{1.5\text{ Eq}}{\text{L}} \text{ X } \frac{40.0\text{ g}}{\text{Eq}} \text{ X } \frac{1\text{ L}}{10\text{ dL}} = \underline{6.0\text{ g/dL}}$$

b. **Example 2**. What is the Eq/L concentration of a 5.0 g/dL $CaCO_3$ solution?

Solution. Read the problem carefully and determine the desired unit of concentration.

Normality (Eq/L).

Determine the GEW of the compound.

$CaCO_3$

Ca 40.1 X 1 = 40.1
C 12.0 X 1 = 12.0
O 16.0 X 3 = + 48.0
100.1 g/mol

$$\frac{100.1\ \text{g}}{\text{mol}} \times \frac{1\ \text{mol}}{2\ \text{Eq}} = 50.05\ \text{g/Eq}$$

Use the GEW and the appropriate volume conversion factor to express the Eq/L concentration.

$$\frac{5.0\ \text{g}}{\text{dL}} \times \frac{1\ \text{Eq}}{50.05\ \text{g}} \times \frac{10\ \text{dL}}{1\ \text{L}} = \underline{1.0\ \text{Eq/L}}$$

5-4. CONVERTING FROM mol/L TO Eq/L AND FROM Eq/L TO mol/L

The conversion is based on the use of TPIV expressed as Eq/mol and used as an appropriate factor.

a. **Example 1**. What is the mol/L concentration of a 4 Eq/L $Ba(OH)_2$ solution?

Solution. Read the problem carefully and determine the desired unit of concentration.

Molarity (mol/L).

Determine the number of equivalents per mole of compound and use it as an appropriate conversion factor to express the solution in the desired units.

$$\frac{4.0\ \text{Eq}}{\text{L}} \times \frac{1\ \text{mol}}{2\ \text{Eq}} = \underline{2.0\ \text{mol/L}}$$

b. **Example 2**. What is the Eq/L concentration of a 1 mol/L $AlPO_4$ solution?

Solution. Read the problem carefully and determine the desired unit of concentration.

Normality (Eq/L)

Determine the number of equivalents per mole of compound and use the appropriate conversion factor to express the solution in the desired units.

$$\frac{1\ \text{mo}}{\text{L}}\ \text{X}\ \frac{3\ \text{Eq}}{1\ \text{mol}} = \underline{3\ \text{Eq/L}}$$

Continue with Exercises

EXERCISES, LESSON 5

INSTRUCTIONS: Answer the following exercises by writing the answer in the space provided at the end of the question.

After you have completed all the exercises, turn to "Solutions to Exercises" at the end of the lesson and check your answers. For each exercise answered incorrectly, reread the material referenced with the solution.

1. Find the Eq/L concentration of a 4.0 mol/L $Fe_2(SO_4)_3$ solution.

   ____

2. What is the g/dL concentration of a 2.0 Eq/L NH_4NO_3 solution?

   ____

3. Calculate the mEq/L concentration of a 0.10 g/dL Na_2SO_4 solution.

   ____

4. A 1-liter flask contains 500 mL of a 0.75 Eq/L $CuWO_4$ solution. Calculate the mol/L concentration of the $CuWO_4$ solution.

   ____

5. Determine the mol/L concentration of a 20 g/dL $CaCO_3$ solution.

   ____

6. A $NaNO_3$ solution has a 300 mg/dL concentration. What is the mol/L concentration of the solution?

   ____

7. What is the Eq/L concentration of 100 mL of a 2.0 mol/L H_2SO_4?

____

8. Find the mg/dL concentration of a 800 mEq/L $(NH_4)_2SO_4$ solution.

____

9. What is the mol/L concentration of a 200 mEq/L $Cu_2CO_3.5H_2O$ solution?

____

10. What is the Eq/L concentration of a 745 mg/dL KCl solution?

____

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 5

1. 24 Eq/L (para 5-4)

2. 16 g/dL (para 5-3)

3. 14.1 mEq/L (para 5-3)

4. 0.38 mol/L (para 5-4)

5. 2.0 mol/L (para 5-2)

6. 0.0353 mol/L (para 5-2)

7. 4.0 Eq/L (para 5-4)

8 5280 mg/dL (para 5-3)

9. 0.100 mol/L (para 5-4)

10. 0.0999 Eq/L (para 5-3)

End of Lesson 5

LESSON ASSIGNMENT

LESSON 6	Dilutions.
TEXT ASSIGNMENT	Paragraphs 6-1 through 6-5.
LESSON OBJECTIVE	After completing this lesson, you should be able to: Calculate simple dilutions, serial dilutions, working solutions made from stock solutions, and new standards by mixing known standard concentrations.
SUGGESTION	After completing the assignment, complete the exercises of this lesson. These exercises will help you to achieve the lesson objective.

LESSON 6

DILUTIONS

6-1. DISCUSSION

A dilution is a laboratory procedure in which the concentration of a sample or solution is reduced by the addition of solvent (diluent). In the laboratory, a dilution is commonly performed when the concentration of an unknown is greater than the limits of linearity of a given quantitative procedure or when a working solution must be prepared from a stock.

6-2. SIMPLE DILUTIONS

Dilutions are expressed as a ratio between the volume of the original solution to the volume of final solution. The conventional form is to express the dilution as the ratio of one unit of the original volume to the final volume. Thus, when 2 mL of sample is added to 8 mL of diluent, the result is a 2:10 dilution, but this is preferably expressed as a 1:5 dilution. The reciprocal of the dilution is the dilution factor. This is the factor by which quantitative results are multiplied to give the final concentration for the original, undiluted sample. For a serum sample diluted 1:10 for a glucose determination, the test result must be multiplied by the dilution factor, 10, to determine the actual concentration of the undiluted serum sample. To calculate the concentration of a dilute solution, multiply the concentration of the original solution by the dilution, expressed as a fraction.

NOTE: A one to ten dilution may be expressed as 1:10 or equivalently as 1/10. This, of course, applies to all dilutions.

a. **Example 1**. A specimen is diluted by combining 3 mL of serum with 21 mL of saline. What is the dilution of the serum?

Solution. Read the problem carefully and calculate the dilution.

NOTE: Volume units must be the same when calculating dilutions.

$$\begin{array}{r l} 3 & \text{mL (parts serum)} \\ + \underline{\ 21} & \text{mL (parts saline)} \\ 24 & \text{mL (total parts)} \end{array}$$

Dilution = 3 mL ÷ 24 mL

Put the ratio in standard form by dividing each side of the ratio by the sample volume (3).

$$\frac{3\text{ mL}}{3\text{ mL}} \div \frac{24\text{ mL}}{3\text{ mL}} = \underline{1:8}$$

The dilution of this specimen is 1:8 and if this information were to be used in calculating the concentration of the specimen, the concentration of the diluted sample would be multiplied by 8 in order to determine the concentration of the original, undiluted sample.

NOTE: Dilutions and dilution factors have no unit of report, because the units divide out.

b. **Example 2**. A 5 mol/L standard solution has been diluted 1:5, what is the concentration of the dilute solution?

Solution. Read the problem carefully, and determine the unknown quantity. The unknown quantity is the concentration of the dilute solution. In order to solve this type of problem you must express the dilution as a fraction. If you have a fraction of a number, as is the case with this problem, you are in effect multiplying the fraction times the original number.

$$5\text{ mol/L} \times \frac{1}{5} = \underline{1\text{ mol/L}}$$

c. **Example 3**. A glucose determination was performed on a serum specimen that was diluted 1:10. The concentration of the diluted specimen was determined to be 73 mg/dL. What is the concentration of the undiluted serum?

Solution. Read the problem carefully and determine the unknown quantity. In this particular instance you need to determine the concentration of a serum specimen, based upon a glucose determination that was performed on a diluted sample.

The dilution factor is the reciprocal of the dilution. This is the factor by which the quantitative results are multiplied to determine the concentration of the undiluted serum sample.

$$(73\text{ mg/dL})(10) = \underline{730\text{ mg/dL}}$$

d. **Example 4**. Determine the amount of serum in 40 mL of a 1:5 dilution of serum with saline.

Solution. Read the problem carefully and use ratio and proportion to determine the amount of serum in the given volume of solution.

$$\frac{\text{1 part serum}}{\text{5 total parts}} \div \frac{\text{x mL serum}}{\text{40 mL total}}$$

$$\text{(5 total parts) (x mL serum)} = \text{(40 mL total)(1 part serum)}$$

$$\frac{\text{40 mL total)(1 part serum)}}{\text{5 total parts}} = \text{x mL serum}$$

$$\text{x mL serum} = \underline{\text{8 mL serum}}$$

6-3. SERIAL DILUTIONS

Serial dilution techniques are used when it is necessary to test several successive dilutions on the same sample. The procedure involves the preparation of an initial dilution and subsequent redilution of aliquots (portions) to give progressively more dilute solutions. In order to calculate the final or intermediate dilutions, use each dilution of interest as a multiplicative factor, of which the resulting product is the dilution.

a. **Example 1**. A serum specimen was successively diluted 1:2, 1:2, and 1:2 with saline. What is the final dilution of the specimen?

Solution. When determining the final dilution of a series of dilutions use each dilution as a multiplicative factor. The resulting product is the final dilution.

$$\frac{1}{2} \text{ X } \frac{1}{2} \text{ X } \frac{1}{2} = \frac{1}{8} \text{ or equivalently, } \underline{1:8}$$

b. **Example 2**. A 1:10 dilution of a substance is rediluted 3:5, 2:15, and 1:2. What is the final dilution?

Solution. When determining the final dilution of a series of dilutions, use each dilution as a multiplicative factor. The resulting product is the final dilution.

$$\frac{1}{10} X \frac{3}{5} X \frac{2}{15} X \frac{1}{2} = \frac{6}{1500}$$

Express the dilution in lowest terms.

<u>1:250</u>

c. **Example 3**. A 5 mol/L solution of HCl is diluted 1:5. The resulting solution is diluted 1:10. Determine the concentration of the final dilution.

Solution. When determining the final dilution of a series of dilutions use each dilution as a multiplicative factor. The resulting product is the final dilution.

$$\frac{1}{5} X \frac{1}{10} = \frac{1}{50}$$

Once the final dilution has been determined, the original concentration is multiplied by the dilution to yield the concentration of the dilute HCl solution.

$$\frac{1}{50} X 5 \text{ mol/L} = \frac{5}{50} \text{ mol/L or, eqivalently, } \underline{0.1 \text{ mol/L}}$$

6-4. PREPARATION OF A WORKING SOLUTION FROM A STOCK SOLUTION

In some instances, a laboratory may utilize a stock solution that will be diluted in order to be used as a working solution. In changing the concentration of the solution, a basic relationship is employed. This relationship is based on the fact that the amount of solute remains constant in both solutions and can be expressed as follows: The volume of the given solution times the concentration of that solution equals the volume of the resulting solution times the concentration of the second solution. An expression that describes this relationship is as follows:

$C_1V_1 = C_2V_2$

Where:

C_1 = Concentration of the stock solution

V_1 = Volume of the stock solution

C_2 = Concentration of the working solution

V_2 = Volume of the working solution

NOTE: The relationship between the volumes and concentrations involved is that of an inverse proportion; that is, when the volume of one solution is greater, the concentration of the solution of greater volume is less.

a. **Basic Rules for Solving $C_1V_1 = C_2V_2$ Problems**.

(1) Three of the 4 values must be known.

(2) The units of volume and concentration must be the same respectively.

(3) It is crucial that the volume and concentration that relate to one another are identified.

(4) Any unit of volume or concentration may be used.

b. **Example 1**. How much 30.03.0

Solution. Read the problem carefully and determine the relationship that will allow you to determine the unknown quantity.

$C_1V_1 = C_2V_2$

Ensure that the concentration and volume units are the same respectively. Using the appropriate expression, substitute the given values.

$C_1V_1 = C_2V_2$

(30.0

Solve for the unknown quantity.

$$V_1 = \frac{(3.0\%)\ (100\ \text{mL}}{30.0} = \underline{10\ \text{mL}}$$

10 mL of 30.0 % alcohol must be added to a 100 mL volumetric flask and q.s. to the mark with water in order to attain a concentration of 3.0 %.

Solution via dimensional analysis:

Calculate the amount of solute in the final (diluted) solution.

$$100\ \text{mL} \times \frac{1\ \text{dL}}{100\ \text{mL}} \times \frac{3.0\ \text{mL}}{\text{dL}} = 3.0\ \text{mL alcohol}$$

Multiply the amount of solute needed in the final solution times the percent concentration of the original solution, used appropriately as a factor.

$$3.0\ \text{mL alcohol} \times \frac{\text{dL}}{30.0\ \text{mL alcohol}} = 0.10\ \text{dL}$$

Expressing the volume of original solution in milliliters yields the same results as before.

$$0.10\ \text{dL} \times \frac{100\ \text{mL}}{\text{dL}} = \underline{10\ \text{mL}}$$

NOTE: A key for this particular type of problem is that in most instances the problem refers to the preparation of a dilute solution from a concentrated stock. Another key is in the fact that this type of problem will have either two concentrations and one volume or two volumes and one concentration.

c. **Example 2**. How much 0.100 Eq/L $Ca(OH)_2$ can be prepared from 10.0 mL of 6.00 mol/L $Ca(OH)_2$?

Solution. Read the problem carefully and determine the relationship that will allow you to solve the problem for the unknown quantity.

$C_1V_1 = C_2V_2$

Initial inspection of the problem reveals that the concentration units are different. Either concentration may be converted, the requirement is that both units must be the same.

Express 0.100 Eq/L $Ca(OH)_2$ in terms of mol/L

$$\frac{0.100 \text{ Eq}}{\text{L}} \text{ X } \frac{\text{mol}}{2 \text{ Eq}} = 0.0500 \text{ mol/L}$$

Using the appropriate expression substitute the values.

$C_1V_1 = C_2V_2$

$(6.00 \text{ mol/L})(10.0 \text{ mL}) = (0.0500 \text{ mol/L})V_2$

Solve for the unknown quantity.

$$V_2 = \frac{6.00 \text{ mol/L}) (10.0 \text{ mL})}{0.0500 \text{ mol/L}} = \underline{1200 \text{ mL}}$$

6-5. PREPARATION OF STANDARDS BY MIXING

When preparing a new standard by mixing two or more standards of known concentration, the concentration of the new standard times its volume is equal to the sum of the concentration of the first standard times its volume plus the concentration of the second standard times its volume and so on depending on the number of standards being mixed. Keep in mind that the concentration of the new standard will never be greater than the highest standard nor less than the lowest standard. All concentrations and volumes must be in the same units respectively. An expression that describes the relationship between the standards being mixed is as follows:

$C_1V_1 = C_2V_2 + C_3V_3 \ldots + C_nV_n$

Where:

C_1 = Concentration of the new standard

V_1 = Volume of the new standard (total volume)

C_2, C_3, C_n = Concentration of the standards used in preparation

V_2, V_3, V_n = Volume of the standards used in preparation

a. **Example 1**. You are preparing a series of standards for an automated glucose procedure. You already have a 50.0 mg/dL and a 300 mg/dL standard made, and are making intermediate concentrations by mixing various quantities of the two. If you mix 3.00 mL of the 50.0 mg/dL standard with 2.00 mL of the 300 mg/dL standard, what is the concentration of the new standard?

Solution. Read the problem carefully and select the formula that will allow you to solve the problem for the unknown quantity.

$$C_1V_1 = C_2V_2 + C_3V_3 \ . . . + C_nV_n$$

Ensure that all concentrations and volumes are expressed in the same units, respectively. (This problem does not require any concentration or volume unit conversion.)

To determine V_1, add the volumes of each of the standards being mixed. In this problem, V_1 equals 5.00 mL since 3.00 mL of the 50.0 mg/dL standard and 2.00 mL of the 300 mg/dL standard are being mixed.

Substitute the given information and solve for the unknown quantity.

$$C_1 \text{ (5.00 mL)} = \text{(50.0 mg/dL)(3.00 mL)} + \text{(300 mg/dL)(2.00 mL)}$$

$$C_1 = \frac{\text{(50.0 mg/dL)(3.00 mL)} + \text{(300 mg/dL)(2.00 mL)}}{\text{5.00 mL}}$$

$$C_1 = \underline{\text{150 mg/dL}}$$

b. **Example 2**. What is the concentration of a standard solution that was prepared by mixing 5.0 mL of a 6.0 Eq/L $Ca_3(PO_4)_2$ standard, 8.0 mL of a 2.0 mol/L $Ca_3(PO_4)_2$ standard, and 3.0 mL of a 31 g/dL $Ca_3(PO_4)_2$ standard?

Solution. Read the problem carefully and determine the relationship that will allow you to solve the problem for the unknown quantity.

$$C_1V_1 = C_2V_2 + C_3V_3 \ . . . + C_nV_n$$

Ensure that all concentrations and volumes are expressed in the same units, respectively.

NOTE: In this example, convert the concentrations to mol/L. <u>Consistency of the units is the only requirement</u>. One unit is not preferred over another in this type of calculation.

Convert 6.0 Eq/L $Ca_3(PO_4)_2$ to mol/L.

$$\frac{6.0\ \text{Eq}}{\text{L}} \times \frac{1\ \text{mol}}{6\ \text{Eq}} = 1.0\ \text{mol/L}$$

Convert 31 g/dL $Ca_3(PO_4)_2$ to mol/L.

Calculate the GMW of $Ca_3(PO_4)_2$.

$Ca_3(PO_4)_2$

Ca 40.1 X 3 = 120.3
P 31.0 X 2 = 62.0
O 16.0 X 8 = <u>+ 128.0</u>
310.3 g/mol

Use the appropriate factors to determine the mol/L concentration.

$$\frac{31\ \text{g}}{\text{dL}} \times \frac{1\ \text{mol}}{310.3\ \text{g}} \times \frac{10\ \text{dL}}{1\ \text{L}} = 1.0\ \text{mol/L}$$

Determine the total volume.

1.0 mol/L 5.0 mL
2.0 mol/L 8.0 mL
1.0 mol/L <u>+ 3.0 mL</u>
16.0 mL

Substitute the given information and solve for the unknown quantity.

$$C_1V_1 = C_2V_2 + C_3V_3 \ldots + C_nV_n$$

$$(C_1)(16.0\ \text{mL}) = (1.0\ \text{mol/L})(5.0\ \text{mL})+(2.0\ \text{mol/L})(8.0\ \text{mL})+(1.0\ \text{mol/L})(3.0\ \text{mL})$$

$$C_1 = \frac{(1.0\ \text{mol/L})(5.0\ \text{mL})+(2.0\ \text{mol/L})(8.0\ \text{mL})+(1.0\ \text{mol/L})(3.0\ \text{mL})}{6\ \text{mL}}$$

$C_1 = \underline{1.5\ \text{mol/L}}$

Continue with Exercises

EXERCISES, LESSON 6

INSTRUCTIONS: Answer the following exercises by writing the answer in the space provided at the end of the question.

After you have completed all the exercises, turn to "Solutions to Exercises" at the end of the lesson and check your answers. For each exercise answered incorrectly, reread the material referenced with the solution.

1. A 10.0 Eq/L NaOH solution was diluted 1:10, rediluted 1:5 and again to 3:15.

   a. Calculate the final dilution factor.

   b. Calculate the final concentration of the diluted NaOH.

2. How much 20 percent NaOH can be prepared from 1.5 mL of 70 percent NaOH?

3. How many milliliters of 6.0 mol/L HCl is needed to make 6.0 liters of 0.010 mol/L HCl?

4. How many milliliters of 3.0 Eq/L $LiCO_3$ must be diluted to prepare 5.0 L of a 20 mEq/L $LiCO_3$ solution?

5. A serum specimen was diluted by adding 100 L of serum to 0.50 mL of saline and 0.10 mL of deionized water. A glucose determination was performed and the concentration of the diluted sample was found to be 123 mg/dL. What is the concentration of the undiluted patient specimen?

6. A urea standard was prepared by mixing 2.0 mL of a 0.050 g/dL solution and 2.0 mL of a 10 mg/dL standard. What is the mg/dL concentration of the prepared standard urea standard?

7. How much 12.0 Eq/L H_2SO_4 is required to prepare 100 mL of a 60 mEq/L solution?

8. A 10 L solution of 0.50 mol/L concentration was prepared by adding 250 mL of stock solution and adjusting the volume to 10 L. What is the concentration of the stock solution?

9. How much water must be added to 500 mL of a 10 g/dL NaOH to dilute the concentration to 0.50 Eq/L?

10. What is the Eq/L concentration of a sodium standard that was prepared by mixing 2.0 mL of a 5.0 Eq/L Na_2HPO_4 with 4.0 mL of a 1.0 mol/L Na_2HPO_4?

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 6

1. a. 1:250 (para 6-3)
   b. 0.040 Eq/L (para 6-3)

2. 5.2 mL (para 6-4)

3. 10.0 mL (para 6-4)

4. 33.33 mL (para 6-4)

5. 861.0 mg/dL (para 6-2)

6. 30.0 mg/dL (para 6-5)

7. 0.5 mL (para 6-4)

8. 20.0 mol/L (para 6-4)

9. 2000.0 mL (para 6-4) The problem asks "how much water must be added."

10. 3.0 Eq/L (para 6-5)

End of Lesson 6

LESSON ASSIGNMENT

LESSON 7	Titration.
TEXT ASSIGNMENT	Paragraphs 7-1 through 7-5. (You are also responsible for the material in the previous lessons.)
LESSON OBJECTIVES	After completing this lesson, you should be able to: 7-1. Solve basic titration problems. 7-2. Convert concentrations and calculate dilutions for titration problems.
SUGGESTION	After completing the assignment, complete the exercises of this lesson. These exercises will help you to achieve the lesson objectives.

LESSON 7

TITRATION

7-1. DISCUSSION

One of the basic methods of analytical chemistry is titration. Titration is the incremental addition of a titrant into a known volume of titrating fluid in the presence of an indicator that identifies the neutralization point. Titration is based on the following principle: "One equivalent weight of a substance will react exactly with one equivalent weight of another substance." Equivalent weight refers to the reacting ability of a substance. When performing a titration, the concentration of the solutions involved MUST be expressed in terms of the concentration units that indicate reacting strength, Eq/L or mEq/L.

a. The number of equivalents or milliequivalents present in a given volume of solution indicates the reacting strength of a solution. Since titration is based on the reacting strength of one solution versus the reacting strength of another solution (usually an acid versus a base), the Eq/L concentration of the acid times the volume of the acid is equal to the Eq/L concentration of the base times the volume of the base. A familiar formula that states these facts is:

$$C_1V_1 = C_2V_2$$

Where:

C_1 = Eq/L or mEq/L concentration of the acid

V_1 = Volume of the acid

C_2 = Eq/L or mEq/L concentration of the base

V_2 = Volume of the base

NOTE: It is not crucial that the variables follow the above format. However, it is crucial that the appropriate concentration and volume be multiplied.

b. Associated Terms Follow:

(1) Acidimetry. Acidimetry is defined as the determination of the concentration of an unknown acid by using a standard base (a base in which the exact concentration is known).

(2) Alkalimetry. Alkalimetry is defined as the determination of the concentration of an unknown base by using a standard acid.

(3) Endpoint. The endpoint is defined as the point where equivalent amounts of two reactants have been brought together, the point of neutralization.

(4) Indicator. An indicator is a dye that changes color at the endpoint.

7-2. INDICATORS AND GENERAL RULES OF USE

a. Shown below are some of the indicators that may be used in acid-base titrations. These indicators vary in sensitivity and pH range and are chosen according to the equivalence point or stoichiometric endpoint of the reactants used in the titration.

		Color In:	
Indicator	pH Range	Acid	Base
Methyl violet 6B	1.5 - 3.2	Blue	Violet
Methyl orange	3.1 - 4.4	Red	Yellow
Ethyl orange	3.4 - 4.8	Red	Yellow
Bromcresol green	3.8 - 5.4	Yellow	Green
Methyl red	4.2 - 6.2	Red	Yellow
Bromthymol blue	6.0 - 7.6	Yellow	Blue
Phenol red	6.8 - 8.4	Yellow	Red
Phenolphthalein	8.2 - 10.0	Colorless	Red

b. The following general rules apply in choosing the proper indicator for titrations unless a special indicator is specified: (1) when titrating a strong acid with a strong base, use phenolphthalein, (2) When titrating a weak acid with a strong base, use phenolphthalein, (3) when titrating a weak base with a strong acid use methyl red, (4) never titrate a weak base with a weak acid because indicators will not give a sharp color change (when the endpoint is reached) in this combination. When using phenolphthalein and other colorless indicators, arrange the titration for the observation of the appearance of color and NOT the disappearance of color.

7-3. PERFORMING AN ACID-BASE TITRATION

The general method that is normally used in performing acid-base titrations is outlined as follows:

a. Place a specific volume of unknown into an Erlenmeyer flask and record the volume.

b. Place two to three drops of indicator into the solution to be titrated. The amount of indicator to be used may be stated in a specific procedure.

c. Fill a buret or serological pipet with the titrating solution of known concentration (Eq/L) and adjust the solution level until the bottom of the meniscus is at zero.

d. Add the titrant from the buret or pipet dropwise to the solution being titrated. It is extremely important to continuously swirl the titration reaction flask as the titrant is added; this allows the specialist to observe the endpoint. If you have sufficient unknown to perform the titration several times, it is recommended that you add the titrant quite rapidly the first time to obtain a rough approximation of the endpoint. This will save a considerable amount of time if several titrations have to be performed.

e. When the endpoint is reached (indicator changes color), record the volume of titrant used by reading the bottom of the meniscus in the buret or pipet.

7-4. SOLVING TITRATION PROBLEMS

a. **Basic Rules for Solving Titration Problems**.

(1) Three of the four values must be known.

(2) The units of concentration MUST be Eq/L or mEq/L.

(3) The units of volume and concentration must be the same respectively.

(4) It is crucial that the volume and concentration that relate to one another are identified.

b. **Example 1**. It took 5.2 mL of a 0.10 Eq/L NaOH to titrate 5.0 mL of a HCl solution. What is the mEq/L concentration of the acid?

Solution. Read the problem carefully and select the expression that will allow you to solve the problem for the unknown quantity.

$C_1V_1 = C_2V_2$

Ensure that the concentration of the solutions are expressed in terms of their normality and the units of volume are the same.

Substitute the given information into the appropriate expression.

$C_1V_1 = C_2V_2$

$(0.10 \text{ Eq/L})(5.2 \text{ mL}) = (C_2)(5.0 \text{ mL})$

Solve for the unknown quantity.

$$C_2 = \frac{(0.10 \text{ Eq/L})(5.2 \text{ mL})}{5.0 \text{ mL}} = 0.10 \text{ Eq/L}$$

Convert the concentration to mEq/L to satisfy the problem using the appropriate conversion factor.

$$\frac{0.10 \text{ Eq}}{\text{L}} \text{X} \frac{1000 \text{ mEq}}{1 \text{ Eq}} = \underline{100 \text{ mEq/L}}$$

Solution. Solving the problem with dimensional analysis yields the same results.

Determine the equivalents of base used in the titration.

$$5.2 \text{ mL X} \frac{1 \text{ L}}{1000 \text{ mL}} \text{X} \frac{0.10 \text{ Eq}}{\text{L}} = 0.00052 \text{ Eq}$$

According to the "Rule of Equivalence," an equal amount of acid is needed to neutralize the solution.

The desired number of equivalents of acid per given volume of acid used in the titration times the appropriate conversion factors yield the desired results.

$$\frac{0.00052 \text{ Eq}}{5.0 \text{ mL}} \text{X} \frac{1000 \text{ mL}}{1 \text{ L}} \text{X} \frac{1000 \text{ mEq}}{1 \text{ Eq}} = \underline{100 \text{ mEq/L}}$$

c. **Example 2**. A 0.25 mol/L H_2SO_4 solution was used to titrate 4.0 mL of a 1.0 Eq/L NaOH solution. What volume of acid was required to perform the titration?

Solution. Read the problem carefully and select the expression that will allow you to solve the problem for the unknown quantity.

$C_1V_1 = C_2V_2$

The concentration of the sulfuric acid in this problem is in mol/L and must be converted to Eq/L. Convert 0.25 mol/L H_2SO_4 to Eq/L.

Initial inspection reveals an inconsistent unit (mol/L). Convert the mol/L concentration of the acid to Eq/L.

$$\frac{0.25 \text{ mol}}{\text{L}} \times \frac{2 \text{ Eq}}{\text{mol}} = 0.50 \text{ Eq/L}$$

Substitute the information into the problem solving expression.

$C_1V_1 = C_2V_2$
$(0.50 \text{ Eq/L})(V_2) = (1.0 \text{ Eq/L})(4.0 \text{ mL})$

Solve for the unknown quantity.

$$V_2 = \frac{(1.0 \text{ Eq/L})(4.0 \text{ mL})}{0.50 \text{ Eq/L}} = \underline{8.0 \text{ mL}}$$

7-5. CALCULATING THE REACTING STRENGTH OF A SOLUTION

At times it may be necessary to calculate the reacting strength of a solution in order to adjust its concentration. For example, potassium acid phthalate (a primary standard) is often used to prepare secondary standards such as sodium hydroxide. In order to adjust a dilute NaOH solution to a higher concentration, it is necessary to determine the reacting strength of the dilute solution so that the proper adjustment can be made.

Example**.** You are preparing 100 mL of a 1.0 Eq/L NaOH standard solution. After titrating this solution against $KHC_8H_4O_4$, you discover that the concentration of the NaOH solution is <u>actually</u> 0.90 Eq/L. The flask containing the NaOH has 90 mL of solution remaining.

a. How many mEq are left in the flask? Read the problem carefully and determine the unknown quantity.

Milliequivalents of compound that remain in the flask.

The remaining volume, expressed in liters, times the actual Eq/L concentration times an appropriate factor for conversion of Eq/L to mEq/L will yield the desired result.

$$90 \text{ mL} \times \frac{\text{L}}{1000 \text{ mL}} \times \frac{0.90 \text{ Eq}}{\text{L}} \times \frac{1000 \text{ mEq}}{\text{Eq}} = 81 \text{ mEq}$$

b. How many mEq of NaOH must be added in order to adjust the volume and concentration to 100 mL of 1.0 Eq/L NaOH solution? In order to determine how many mEq of NaOH must be added you must first ascertain the number of mEq present in the desired volume and concentration of solution. You desire 100 mL of a 1.0 Eq/L NaOH. With this information the number of mEq desired can be calculated.

$$100 \text{ mL X } \frac{\text{L}}{1000 \text{ mL}} \text{ X } \frac{1.0 \text{ Eq}}{\text{L}} \text{ X } \frac{1000 \text{ mEq}}{\text{Eq}} = 100 \text{ mEq of NaOH desired}$$

mEq needed = mEq desired — mEq in flask

100 mEq — 81 mEq = $\underline{\text{19 mEq}}$ needed

Thus, 19 mEq of NaOH needs to be added to the flask and sufficient quantity (q.s.) volume to 100 mL in order to obtain a 1.0 Eq/L NaOH solution.

c. How much stock 10 Eq/L NaOH must be added to the flask in order to adjust the concentration to 1.0 Eq/L? Multiply the number of mEq needed times the concentration of the stock along with appropriate conversion factors for the volume and concentration to obtain the desired results.

$$19 \text{ mEq X } \frac{\text{Eq}}{1000 \text{ mEq}} \text{ X } \frac{\text{L}}{10 \text{ Eq}} \text{ X } \frac{1000 \text{ mL}}{\text{L}} = \underline{1.9 \text{ mL}}$$

1.9 mL of the stock solution must be added and the solution Q.S. to 100 mL to yield the desired solution.

Continue with Exercises

EXERCISES, LESSON 7

INSTRUCTIONS: Answer the following exercises by writing the answer in the space provided at the end of the question.

After you have completed all the exercises, turn to "Solutions to Exercises" at the end of the lesson and check your answers. For each exercise answered incorrectly, reread the material referenced with the solution.

1. A 100 mEq/L mercuric solution was used in a chloride determination. The assay required 4.0 mL of the specimen, whose concentration was determined to be 0.090 Eq/L. What volume of mercuric solution was used?

2. How much 0.50 Eq/L HCl is required to titrate 4.0 mL of a 0.75 mol/L $Ba(OH)_2$ solution?

3. 7.0 mL of a 3.65 g/dL HCl was used to titrate 5.0 mL of $Ca(OH)_2$. What is the mol/L concentration of the base?

4. You are directed to prepare 100 mL of a 500 mEq/L KOH standard. The primary standard used has a concentration of 0.400 Eq/L. When titrating your KOH solution, you discover that 4.0 mL of the primary standard reacted exactly with 5.0 mL of the KOH solution.

   a. What is the actual Eq/L concentration of the KOH solution?

   b. How many mEq are required to adjust the volume and concentration of the remaining 85 mL of KOH solution to 100 mL of a 500 mEq/L KOH solution?

   c. How many mL of 12 Eq/L KOH stock must be added in order to adjust the concentration of the KOH standard to 500 mEq/L?

5. Calculate the amount of 0.500 Eq/L NaOH needed to neutralize 8.0 mL of 2.0 mol/L H_3PO_4 solution.

6. In a titration, 7.2 mL of HCl of unknown strength neutralized 5.0 mL of 0.500 Eq/L NaOH. What is the concentration of the acid?

7. What is the reacting strength in milliequivalents of 85 mL of a 0.92 Eq/L HCl solution?

8. How many equivalents of NaCl are contained in 6.0 dL of a 0.85 g/dL NaCl solution?

9. A secondary standard, whose initial volume was 100 mL, was titrated against a primary standard. The concentration of the secondary standard was determined to be 0.46 Eq/L. 7.3 mL of the secondary standard was used in the titration. How many mEq of compound remain in the flask?

10. How many mL of 1.5 Eq/L $Ca(OH)_2$ are required to titrate 7.0 mL of a 5.0 g/dL HCl solution?

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 7

1. 3.6 mL (para 7-4)

2. 12 mL (para 7-4)

3. 0.70 mol/L (para 7-4)

4. a. 0.32 Eq/L (para 7-4)
   b. 22.8 mEg (para 7-5)
   c. 1.9 mL (para 7-5)

5. 96 mL (para 7-4)

6. 0.35 Eq/L (para 7-4)

7. 78.2 mEq (para 7-5)

8. 0.087 Eq (para 4-6)

9. 43 mEq (para 7-5)

10. 6.4 mL (para 7-4)

End of Lesson 7

LESSON ASSIGNMENT

LESSON 8	Concentrated Acids and Bases.
TEXT ASSIGNMENT	Paragraphs 8-1 through 8-6. You are still responsible for material in the previous lessons.
LESSON OBJECTIVE	After completing this lesson, your should be able to: Prepare and calculate acid and base solutions.
SUGGESTION	After completing the assignment, complete the exercises of this lesson. These exercises will help you to achieve the lesson objective.

LESSON 8

CONCENTRATED ACIDS AND BASES

8-1. DISCUSSION

As discussed in paragraph 2-31, concentrated liquids are manufactured as weight-per-weight solutions. An example of this type of solution is concentrated hydrochloric acid that contains 37 grams of HCl per 100 grams of total solution (w/w). These figures indicate that concentrated HCl has a percent purity of 37 percent or that 37 percent of the weight of 100 grams of solution is HCl. Since it is impractical to weigh concentrated acids and/or bases, one needs another method to determine how much concentrated liquid to use. The method used includes use of the specific gravity of the solution, which is an expression of density (g/mL), and the percent purity or assay value (%A) that indicates the proportion of the solution by weight that is the desired substance. The specific gravity and percent purity values are found on the label of most reagent-grade acids and bases. Using these values, one can determine the actual amount of the substance in a given volume of the concentrated solution.

8-2. EXAMPLE

The values listed on the label of a bottle of nitric acid are specific gravity (S.G.) 1.42 and assay 70 percent. What so these values mean?

a. These values mean that 1 milliliter of the solution has a mass of 1.42 grams and that 70 percent of this mass is nitric acid (HNO_3).

b. When preparing a solution using this concentrated liquid, we are concerned only with how much HNO_3 is present. To determine this, multiply the density times the percent assay. The answer is the number of grams of HNO_3 per milliliter of solution.

$$\begin{array}{l} 1.42\ \text{g/mL} \\ \underline{\text{X } 0.70}\ \text{percent } HNO_3 \\ 0.994\ \text{grams of } HNO_3 \text{ per mL of solution} \end{array}$$

c. Therefore, density times % A is an expression that yields the number of grams of desired substance per milliliter of concentrated substance.

NOTE: The % A value is equal to the assay value divided by 100. If the assay value is 70 percent then the % A is 0.70 (70/100). The latter value or it's equivalent is to be used in all calculations.

(g/mL)(% A) = grams of desired substance per unit volume

8-3. SPECIFIC GRAVITY AND DENSITY

a. Specific gravity is a method of measuring density. Density is the amount of matter in a given volume, that is, D = g/mL. Specific gravity is a ratio between the mass of a given volume of any substance and the mass of an equal volume of pure water at 4ºC.

$$\text{S.G.} = \frac{\text{density of solid or liquid}}{\text{density of water at 4ºC}}$$

NOTE: When determining specific gravity, the units divide out. Hence, specific gravity has no unit of report.

b. Since 1 milliliter of pure water at 4ºC has a mass of 1 gram, specific gravity is equal to the numerical value of the mass in grams of one milliliter of any substance. Materials less dense than water have a specific gravity of less than one, whereas materials more dense than water have a specific gravity greater than one.

c. Example. The density of 200 mL of HNO_3 solution is 1.42 g/mL. What is the weight of that solution?

Solution. Restating the problem, one milliliter of this solution weighs 1.42 grams; how much would 200 mL of the same solution weigh? This problem can be solved by ratio and proportion.

$$\frac{1.42\text{ g}}{x\text{ g}} = \frac{1\text{ mL}}{200\text{ mL}}$$

$$(x\text{ g})(1\text{ mL}) = (1.42\text{ g})(200\text{ mL})$$

$$x\text{ g} = \frac{(1.42\text{ g})(200\text{ mL})}{1\text{ mL}}$$

$$x\text{ g} = \underline{284\text{ g}}$$

NOTE: The 284 grams represents the total weight of the solution and not just the weight of the HNO_3 present in the solution.

8-4. PREPARING SOLUTIONS FROM CONCENTRATED LIQUIDS

When preparing a solution in the laboratory, one would normally think of taking a chemical in powder or crystal form (such as NaCl) and calculate the amount to be weighed. When preparing such a solution, one would ask, "How many grams of this substance do I need to make the desired concentration?" Using appropriate factors, you caculate the desired weight required. With concentrated liquids, this cannot be done since it is impractical and too dangerous to try to weigh these liquids. We do know, however, that each concentrated liquid contains a certain mass of solute per milliliter of solution and that this can be determined using the information found on the label of the bottle.

Example**.** A bottle of nitric acid (HNO_3) has a S.G. of 1.42 and an assay of 70.0 percent. How many grams of HNO_3 are contained within 1.0 mL of HNO_3?

Solution. The given volume times the density of the concentrated liquid times the % A yields the desired results.

$$1.0 \text{ mL} \times \frac{1.42 \text{ g}}{\text{mL}} \times \frac{70}{100} = \underline{0.994 \text{ g}}$$

8-5. PROBLEM SOLVING

You have seen in earlier lessons that the mass of a solute needed to prepare a solution of a certain concentration may be determined by the use of various expressions, dependent upon the desired concentration unit. These various problem-solving expressions all yielded the mass in <u>grams</u> needed to prepare a solution of known concentration and volume. Similar problem solving techniques may be used to determine the amount of concentrated liquid needed to yield a desired amount of solute. The examples that follow will illustrate several methods needed to solve problems involving concentrated liquids.

a. **Example 1**. How much concentrated HNO_3 with a S.G. of 1.42 and an assay of 70% would you need to prepare 1.0 liter of a 1.0 mol/L HNO_3 solution?

Solution. As before, with problems that involve molarity, calculating the gram molecular weight of the compound involved is a good starting point.

Calculate the GMW of the compound.

HNO_3

$$\begin{array}{lll} \text{H} & 1.0 \times 1 = & 1.0 \\ \text{N} & 14.0 \times 1 = & 14.0 \\ \text{O} & 16.0 \times 3 = & \underline{+\ 48.0} \\ & & 63.0 \text{ g/mol} \end{array}$$

Determine the amount of HNO_3 needed to prepare the solution.

$$1.0\text{ L} \times \frac{1.0\text{ mol}}{\text{L}} \times \frac{63.0\text{ g}}{\text{mol}} = 63\text{ g } HNO_3$$

Calculate the grams of HNO_3 per milliliter of concentrated liquid.

$$\frac{1.42\text{ g}}{\text{mL}} \times \frac{70}{100} = 0.994\text{ g } HNO_3\text{/mL concentrated liquid}$$

Use the grams per milliliter of solute as an appropriate factor times the number of grams of solute needed to determine the required volume of concentrated liquid.

$$63\text{ g} \times \frac{\text{mL}}{0.994\text{ g}} = \underline{63\text{ mL}}$$

b. **Example 2**. How much H_2SO_4 with a S.G. of 1.84 and an assay of 95% is required to make 500 mL of a 4.00 Eq/L H_2SO_4 solution?

Solution. As before, with problems that involve normality, calculating the gram equivalent weight of the compound involved is a good starting point.

Calculate the GEW of the compound.

H_2SO_4

$$\begin{array}{lll} \text{H} & 1.0 \times 2 = & 2.0 \\ \text{S} & 32.1 \times 1 = & 32.1 \\ \text{O} & 16.0 \times 4 = & \underline{+\ 64.0} \\ & & 98.1\text{ g/mol} \end{array}$$

$$\frac{98.1\text{ g}}{\text{mol}} \times \frac{\text{mol}}{2\text{ Eq}} = 49.0\text{ g/Eq}$$

Determine the amount of solute needed in the desired solution.

$$500\text{ mL} \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{4.00\text{ Eq}}{\text{L}} \times \frac{49.0\text{ g}}{\text{Eq}} = 98.0\text{ g } H_2SO_4$$

Calculate the grams of H_2SO_4 per milliliter of concentrated liquid.

$$\frac{1.84\text{ g}}{\text{mL}} \times \frac{95.0}{100} = 1.75\text{ g } H_2SO_4/\text{mL concentrated liquid}$$

Use the grams of solute per milliliter of concentrated liquid as an appropriate factor times the number of grams of solute needed to determine the required volume of concentrated liquid.

$$98.1\text{ g} \times \frac{\text{mL}}{1.75\text{ g}} = \underline{56.1\text{ mL}}$$

c. **Example 3**. How much concentrated HCl with an S.G. of 1.19 and an assay of 38.0% is required to make 250 mL of a 20.0 g/dL HCL solution?

Solution. Read the problem carefully and determine the desired quantity.

Milliliters of concentrated HCl.

Calculate the amount in grams of solute needed to prepare the solution.

$$250\text{ mL} \times \frac{1\text{ dL}}{100\text{ mL}} \times \frac{20.0\text{ g}}{\text{dL}} = 50.0\text{ g of HCl needed}$$

Determine the grams of HCl per milliliter of concentrated liquid.

$$\frac{1.19\text{ g}}{\text{mL}} \times \frac{38.0}{100} = 0.452\text{ g HCl/mL of concentrated liquid}$$

Use the grams of solute per milliliter of concentrated liquid as an appropriate factor times the number of grams of solute needed to determine the required volume of concentrated liquid.

$$50.0\text{ g} \times \frac{\text{mL}}{0.452\text{ g}} = \underline{111\text{ mL}}$$

8-6. DETERMINING THE CONCENTRATION OF CONCENTRATED LIQUIDS

At times, it may be necessary to determine the concentration of a commercial liquid such as the mol/L, Eq/L, or g/dL concentration. To do this consider the definition of each unit of concentration, then multiply the density (g/mL), given on the bottle times the percent assay (%A) times an appropriate conversion factor to yield the desired unit of concentration. Concentrated commercial liquids are based on g/mL of solution.

a. **Conversion from g/mL to g/dL**. Percent concentration is based on grams of solute per 100 mL of solution.

Example. What is the g/dL concentration of a HNO_3 solution that has a specific gravity of 1.42 and an assay of 70.0 percent

Solution. Express the desired unit of concentration (g/dL) using the information given as appropriate factors.

Calculate the grams of HNO_3 per milliter of concentrated liquid.

$$\frac{1.42\text{ g}}{\text{mL}} \text{X} \frac{70.0}{100} = 0.994\text{ g } HNO_3/\text{ mL concentrated liquid}$$

Calculate the g/dL concentration by multiplying the grams of solute per milliliter of concentrated liquid times an appropriate unit volume conversion factor.

$$\frac{0.994\text{ g}}{\text{mL}} \text{X} \frac{100\text{ mL}}{1\text{ dL}} = \underline{99.4\text{ g/dL}}$$

b. **Conversion from g/mL to mol/L**. Molarity is based on moles of solute per liter of solution.

Example. What is the mol/L concentration of a HNO_3 solution with a S.G. of 1.42 and an assay of 70.0 percent.

Solution. Express the desired unit of concentration (mol/L) using the information given as appropriate factors.

Calculate the GMW of the compound.

HNO_3

$$\begin{array}{lll} \text{H} & 1.0 \times 1 = & 1.0 \\ \text{N} & 14.0 \times 1 = & 14.0 \\ \text{O} & 16.0 \times 3 = & \underline{+\ 48.0} \\ & & 63.0\text{ g/mol} \end{array}$$

Determine the grams per liter of HNO_3.

$$\frac{1.42\ \text{g}}{\text{mL}} \text{ X } \frac{70.0}{100} \text{ X } \frac{1000\ \text{mL}}{\text{L}} = 994\ \text{g/L}\ HNO_3$$

Multiply the gram per liter concentration of the acid times the grams per mole, using the factors appropriately.

$$\frac{994\ \text{g}}{\text{L}} \text{ X } \frac{\text{mol}}{63.0\ \text{g}} = \underline{15.8\ \text{mol/L}}$$

c. **Conversion From g/mL to Eq/L**. Normality is based on equivalents of solute per liter of solution.

Example**.** What is the Eq/L concentration of a H_2SO_4 solution with an S.G. of 1.84 and an assay of 95.0%?

Solution. Express the desired unit of concentration (Eq/L) using the information given as appropriate factors.

Calculate the GEW of the compound.

H_2SO_4

$$\begin{array}{lll} \text{H} & 1.0 \text{ X } 2 = & 2.0 \\ \text{S} & 32.1 \text{ X } 1 = & 32.1 \\ \text{O} & 16.0 \text{ X } 4 = & \underline{+\ 64.0} \\ & & 98.1\ \text{g/mol} \end{array}$$

$$\frac{98.1\ \text{g}}{\text{mol}} \text{ X } \frac{\text{mol}}{2\ \text{Eq}} = 49.0\ \text{g/Eq}$$

Determine the grams per liter of H_2SO_4.

$$\frac{1.84\ \text{g}}{\text{mL}} \times \frac{95.0}{100} \times \frac{1000\ \text{mL}}{\text{L}} = 1748\ \text{g/L}$$

Multiply the gram per liter concentration of the acid times the grams per equivalent, using the factors appropriately.

$$\frac{1748\ \text{g}}{\text{L}} \times \frac{\text{Eq}}{49.0\ \text{g}} = \underline{35.7\ \text{Eq/L}}$$

Continue with Exercises

EXERCISES, LESSON 8

INSTRUCTIONS: Answer the following exercises by writing the answer in the space provided at the end of the question.

After you have completed all the exercises, turn to "Solutions to Exercises" at the end of the lesson and check your answers. For each exercise answered incorrectly, reread the material referenced with the solution.

1. How many mL of phosphoric acid (H3PO4), 1.69 S.G., 85.0 would be needed to make 500 mL of a 1.00 Eq/L H3PO4 solution?

2. A commercial HCl solution has a S.G. of 1.15 and 33.0

   a. Find the mol/L concentration of the solution.

   b. How much concentrated solution would be required to make 1.00 L of a 2.00 Eq/L solution?

3. How much NH4OH with a specific gravity of 0.960 and an assay of 26.0 percent is required to make 750 mL of 4.00 Eq/L ammonium hydroxide solution?

4. What is the g/dL concentration of the concentrated solution in exercise 3?

5. What volume of acetic acid (CH3COOH), S.G. 1.05 and assay of 99.7 percent is required to prepare 250 mL of a 0.500 mol/L acetic acid solution?

6. What is the mass of H2SO4 present in 150 mL of sulfuric acid solution with a specific gravity of 1.86 and a purity of 93.0 percent.

7. What is the mol/L concentration of the sulfuric acid solution in exercise 6?

8. How much of the sulfuric acid solution in exercise 6 would be required to prepare 1000 mL of a 2.50 mol/L sulfuric acid solution?

9. A HCl solution of 500 mL was prepared by using 25.0 mL of commercial reagent, S.G. 1.20 and assay of 34.0 percent. What is the mol/L concentration of the solution that was prepared?

10. A H3PO4 solution was prepared by adding 12.5 mL of concentrated acid, S.G. 1.70 and 87.0 percent purity, to a one liter flask and Q.S. to the mark. Assuming that the acid ionizes completely, what is the normal concentration of the phosphoric acid that was prepared?

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 8

1. 11.4 mL (para 8—5)
2. a. 10.4 mol/L (para 8—6b)
   b. 192 mL (para 6—4)
3. 421 mL (para 8—5)
4. 25.0 g/dL (para 8—6a)
5. 7.16 mL (para 8—5)
6. 259 g (para 8—4)
7. 17.6 mol/L (para 8—6b)
8. 142 mL (para 8—5)
9. 0.559 mol/L (para 8—5)
10. 0.565 Eq/L (para 8—5)

End of Lesson 8

LESSON ASSIGNMENT

LESSON 9	pH and Buffers.
TEXT ASSIGNMENT	Paragraphs 9-1 through 9-23. (You are still responsible for materials in the previous lessons.)
LESSON OBJECTIVES	After completing this lesson, you should be able to: 9-1. Calculate the pH and pOH of a molar acid or base concentration. 9-2. Calculate buffers problems. After studying the assignment, complete the exercises of this lesson. These exercises will help you to achieve the lesson objectives.

LESSON 9

pH and BUFFERS

Section I. INTRODUCTION

9-1. DISCUSSION

a. The interaction of charged particles in which the total number of protons do not equal the number of electrons is responsible for many chemical reactions. These charged particles are called ions. A cation is an ion in which the protons are more numerous than the electrons (+ charge), and anions are ions in which the electrons outnumber the protons (- charge). These oppositely charged ions are attracted to each other and form bonds called ionic bonds. When ionically bound compounds are dissolved in a solvent, the ions separate. This separation is known as dissociation or ionization. The only way in which most ions can undergo chemical reaction is to be in this dissociated state.

b. Three general types of ionic compounds exist: acids, bases, and salts. Simply and incompletely stated, acids are compounds that contribute hydrogen ions to a solution, bases are compounds that contribute hydroxide ions to the solution, and salts yield neither hydrogen nor hydroxide ions to the solution. The concentration of the hydrogen and hydroxide ions will determine the degree of acidity or alkalinity of a solution, and thus have an effect on the kinds of reactions and the speed of the reactions that will occur. It is important, therefore, to know the relative concentrations of the hydrogen and hydroxide ions in a solution.

c. Before discussing hydrogen and hydroxide ion concentration further, it will be important to consider in greater detail the definition of an acid and a base. There are several definitions of the terms acid and base. We will examine two of the definitions that will aid you in further course work.

9-2. ARRHENIUS CONCEPT

An acid is a substance that will produce hydrogen ions (H^+) in an aqueous solution, and a base is a substance that will produce hydroxide ion (OH^-) in an aqueous solution. It is important to note that hydrogen ions do not exist as such in an aqueous solution. Each hydrogen ion is associated with one molecule of water to produce a hydronium ion, H_3O^+.

$$H^+ + H_2O \Longleftrightarrow H_3O^+$$

When hydrochloric acid (HCl), which exists in its pure state as a gas, is dissolved in water, the following reaction takes place:

$HCl + H_2O <====> H_3O^+ + Cl^-$

The hydronium ion can easily react with other groups as a proton donor. The hydrogen ion itself is made up of one proton.

NOTE: The terms hydronium ion, hydrogen ion, and proton donor may be used interchangeably. However, we will simply refer to the hydrogen ion in future course work, but you should understand that this actually implies the presence of the hydronium ion.

One of the primary drawbacks of the Arrhenius concept is that it defines only aqueous solutions as acids and bases. Other theories allow for nonaqueous solutions.

a. **Examples of Arrhenius Acids.**

(1) HCl -- hydrochloric acid.

(2) H_2SO_4 -- sulfuric acid.

(3) H_3PO_4 -- phosphoric acid.

b. **Examples of Arrhenius Bases.**

(1) NaOH -- sodium hydroxide.

(2) $Ca(OH)_2$ -- calcium hydroxide.

(3) KOH -- potassium hydroxide.

9-3. BRONSTED-LOWRY THEORY

The Bronsted-Lowry theory presents a broader definition of the terms acid and base. An acid is a proton donor and a base is a proton acceptor. Hydrochloric acid in solution is a Bronsted-Lowry use it forms hydrogen ions but because it is able to give up that hydrogen ion to another substance, such as to a molecule of water to form hydronium ions. For HCl, the difference between the two definitions seems minor. However, it becomes important when the solvent is not water. For example, when HCl reacts with NH_3 (ammonia), the proton is given up by the HCl and accepted by the NH_3, making the HCl an acid and NH_3 a base.

$HCl + NH_3 <====> NH_4^+ + Cl^-$

acid base acid base

a. **Examples of Bronsted-Lowry Acids.**

(1) NH_4^+ -- ammonium ion.

(2) $HC_2H_3O_2$ -- acetic acid.

(3) H_2CO_3 -- carbonic acid.

b. **Examples of Bronsted-Lowry Bases.**

(1) Cl^- -- chloride ion.

(2) NH_3 -- ammonia.

(3) $Fe(OH)_2$ -- iron (II) hydroxide.

9-4. WEAK ACIDS/BASES VERSUS STRONG ACIDS/BASES

As discussed previously, we stated that acids, bases, and salts are ionic compounds. Substances that break up into ions in solution are termed electrolytes because the solution has the ability to conduct electrical current. A strong electrolyte (acid or base) is a substance that exhibits a high degree of ionization, and a weak electrolyte is a substance that only partially ionizes.

a. Hydrochloric (HCl), sulfuric (H_2SO_4), and nitric (HNO_3) acids are strong acids because they exhibit a high degree of ionization in aqueous solutions. Hydrobromic (HBr), hydroiodic (HI), and perchloric ($HClO_4$) acids are also considered strong acids. All others are weak.

b. The alkaline metal (Group IA, i.e., lithium, sodium, potassium) hydroxides and the alkaline earth (Group IIA, that is, magnesium, calcium, strontium,…) hydroxides are strong bases. Other bases are weak ones.

c. Hydrogen and hydroxide only affect pH or pOH when in the ionized form.

Section II. DYNAMIC EQUILIBRIUM

9-5. DISCUSSION

When an acid ionizes, an equilibrium is established between the unionized acid and its ions. This is indicated by double arrows or a doubled-headed arrow in an equation, showing that the two reactions occur simultaneously. HA is used to represent any acid made up of hydrogen and some anion.

$HA <====> H^+ + A^-$

At the same time, the acid molecule is dissociating, a certain percentage of the ions are reassociating to reform the acid. When the acid is first placed in water, a high degree of dissociation occurs, and the rate of the forward reaction is greater. Gradually, as more ions are formed, the rate of the reverse reaction increases. Eventually, a state of equilibrium is reached in which molecules are being ionized and reassociated at a given rate. Equilibrium exists when the rates, not the number of molecules or concentrations, of the opposing reactions are equal. This is a dynamic, constant process and continues until a force is added to change this equilibrium. For example, if the temperature is increased, the rates of the reaction are increased, and a new equilibrium state is reached. Removing one of the ions or tying them up with another reaction will also shift the equilibrium.

9-6. LAW OF MASS ACTION

a. **Law.** The law of mass action states that the rate of reaction is proportional to the product of the molar concentrations of the reactants.

$$A + B \Longleftrightarrow C + D$$

For this reaction, A and B are the reactants, and C and D are the products. The rate at which C and D are formed is proportional to the concentration of A and B.

$$\text{rate} = K_1 [A] [B]$$

where K_1 is a proportionality constant, and the brackets denote mol/L concentration of the enclosed substance.

The rate for the reverse reaction is:

$$\text{rate} = K_2 [C] [D]$$

b. **Equilibrium Constant (K_{eq}).**

The constants K_1 and K_2 are different. The ratio of the constants is known as the equilibrium constant (K_{eq}), dissociation constant, or ionization constant.

$$K_{eq} = \frac{K_1}{K_2}$$

(1) At equilibrium, the rates of the forward and reverse reactions are equal.

$rate_f = rate_r$

$rate_f = K_1 [A] [B]$

$rate_r = K_2 [C] [D]$

$K_1 [A] [B] = K_2 [C] [D]$

$$\frac{K_1}{K_2} = \frac{[C] [D]}{[A] [B]}$$

$$\frac{K_1}{K_2} = K_{eg}$$

(2) Simply, the ionization (equilibrium) constant is equal to the product of the mol/L concentrations of the products formed in the reaction divided by the product of the concentrations of the reactants.

(3) The equilibrium constant for an acid is called the K_a and for a base the K_b.

(4) K_a and K_b are determined for weak acids and bases only. If the K_a or K_b of a strong acid or base, respectively, were determined it would be an infinitely large number.

(5) Values of equilibrium constants are interpreted as follows:

(a) The constant indicates the strength of a weak electrolyte.

(b) The smaller the value of the K_a or K_b, the weaker the acid or base respectively.

9-7. EXAMPLE

What is the $[H^+]$ of a 0.100 mol/L solution of acetic acid if the K_a for $HC_2H_3O_2$ is 1.75×10^{-5}?

Solution. Determine the equation for the dissociation of the weak acid.

$HC_2H_3O_2$ <====> $H^+ + C_2H_3O_2^-$

Write the equilibrium expression.

$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

Substitute the given information, and solve for the unknown quantity.

$$1.75 \text{ X } 10^{-5} \text{ mol/L} = \frac{[H^+]\ [C_2H_3O_2]}{0.100 \text{ mol/L}}$$

NOTE: Although the actual concentration of the acetic acid is less than 0.100 mol/L after ionization, it is a negligible amount (less than 5 your future course work, the amount of ionized acid or base in this type of problem should be considered to be a negligible amount.

Substitute the variable x for H^+ and $C_2H_3O_2^-$ since for every dissociated molecule of $HC_2H_3O_2$, one hydrogen ion and one acetate radical are produced.

$x^2 = (1.75 \text{ X } 10^{-5} \text{ mol/L})\ (0.100 \text{ mol/L}) = 1.75 \text{ X } 10^{-6} \text{ mol}^2/\text{L}^2$

$x = [H^+] = \underline{1.32 \text{ X } 10^{-3} \text{ mol/L}}$

NOTE: By convention, the equilibrium constants of weak electrolytes are written without units. However, they do have units.

Section III. DISSOCIATION OF WATER

9-8. DISCUSSION

Although water is not an electrolyte, its molecules do have a limited tendency to dissociate into H^+ and OH^- ions.

$HOH <====> H^+ + OH^-$

It can be shown that the tendency of water to dissociate is given by:

$$K = \frac{[H^+]\ [OH^-]}{[HOH]}$$

At 25ºC, the value of K is $1.8 \text{ X } 10^{-16}$.

9-9. ION PRODUCT OF WATER

a. Since undissociated water is present in great excess, its concentration is virtually constant at 55.6 mol/L.

NOTE: The gram molecular weight of water is 18.0; therefore, in 1.00 L (approximately 1000 grams), there are:

$$\frac{\frac{(1000\text{ g})}{(18.0\text{ g/mol})}}{1\text{ L}} = 55.6\text{ MOL/L}$$

b. This constant value for the concentration of water can be incorporated into the dissociation constant to give a new constant, the ion product of water, or K_w.

$$1.8 \text{ X } 10^{-16} = \frac{[H^+]\,[OH^-]}{[55.6]}$$

$$[H^+]\,[OH^-] = 1.0 \text{ X } 10^{-14}$$

Therefore, in pure water,

$$[H^+] = 1.0 \text{ X } 10^{-7}\text{ mol/L}$$

$$[HO^-] = 1.0 \text{ X } 10^{-7}\text{ mol/L}$$

$$K_w = [H^+]\,[HO^-] = 1.0 \text{ X } 10^{-14}$$

c. It is important to realize that the ion product of water, $(H^+]\,[OH^-]$, is constant for all aqueous solutions, even those that contain dissolved acids or bases.

(1) If a large number of H^+ ions are added to pure water, the concentration of OH^- ions must decrease in order that the product of $[H^+]\ [OH^-]$ remains the same.

(2) Conversely, if a large number of hydroxyl ions are added, the $[H^+]$ will have to decrease.

Section IV. pH and pOH

9-10. CONCEPT OF pH

Pure water contains an equal concentration of hydrogen and hydroxide ion. Thus, it is a neutral substance, neither acidic nor basic. If a solution has an excess of hydrogen ions, the solution will be acidic. If a solution has an excess of hydroxide ions, the solution will be alkaline. It is customary to use the hydrogen ion concentration (pH) as a measure of the acidity or alkalinity of a solution. pH is a unit of concentration that allows the technician to work with very dilute concentrations of hydrogen and hydroxide ion in a convenient form. This scale permits the representation of the enormous range of the possible $[H^+]$ concentrations from a 1.0 mol/L $[H^+]$ to a 1.0 X 10^{-14} mol/L $[H^+]$.

9-11. APPLICATION OF CONCEPT OF pH

The notation p denotes "negative log of." Logarithms are discussed in lesson 1, and there is a four-place logarithm table in Appendix B. pH is mathematically defined as:

$-\log [H^+]$ or $\log 1/[H^+]$

a. **Example 1.** What is the pH of a 0.133 mol/L HCl solution?

Solution. Select the expression that allows you to solve for the unknown quantity.

$\text{pH} = -\log [H^+]$

Substitute the given information, and solve for the unknown quantity.

$\text{pH} = -\log 0.133 = -\log [1.33 \times 10^{-1}] = -(0.1239 - 1)$

$\text{pH} = -(-0.876)$

pH = <u>0.876</u>

b. **Example 2**. What is the pH of a 0.020 mol/L acid solution that is ionized 1.8 percent?

Solution. From previous discussion, hydrogen or hydroxide does not contribute to the acidity or alkalinity of a solution unless ionized. So in this problem consider only that hydrogen ion that is ionized.

$\text{pH} = -\log [H^+]$

$\text{pH} = -\log [(0.020 \text{ mol/L}) (0.018)] = -\log [3.6 \times 10^{-4}] = -(.5563 - 4)$

pH = <u>3.4</u>

c. **Example 3.** Determine the pH of a 0.0100 mol/L $HC_2H_3O_2$ solution. The K_a of the acid is 1.75 X $10^{—5}$.

Solution. In determining the pH of a solution of weak electrolyte, you must first calculate the hydrogen ion concentration. Acetic acid ionizes as follows:

$HC_2H_3O_2$ <====> $H^+ + C_2H_3O_2^-$

Write the equilibrium expression:

$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

Substitute the given information, and solve for the hydrogen ion concentration.

$$1.75 \text{ X } 10^{-5} \text{ mol/L} = \frac{[H^+]\,[C_2H_3O_2^-]}{0.100 \text{ mol/L}}$$

Substitute the variable x for H^+ and $C_2H_3O_2^-$, since one hydrogen ion and one acetate ion are produced for every dissociated molecule of $HC_2H_3O_2$.

$[H^+]\,[C_2H_3O_2^-] = (1.75 \text{ X } 10^{-5})\ (0.0100 \text{ mol/L})$

$x^2 = (1.75 \text{ X } 10^{-5} \text{ mol/L})\ (0.0100 \text{ mol/L})$

$x^2 = 1.75 \text{ X } 10^{-7} \text{ mol}^2/\text{L}^2$

$x = [H^+] = 4.18 \text{ X } 10^{-4} \text{ mol/L}$

Using the calculated value for the hydrogen ion concentration, determine the pH of the solution.

$pH = -\log [H^+]$

$pH = -\log [4.18 \text{ X } 10^{-4} \text{ mol/L}] = —(.6212 — 4)$

$pH = -(-3.38)$

pH = <u>3.38</u>

9-12. pOH

pOH is mathematically defined as:

-log $[OH^-]$ or log $1/[OH^-]$

NOTE: pH and pOH have no unit of report.

9-13. pH SCALE

a. pH range: 0-14

b. Neutral: pH 7

c. Acidic: less than pH 7

d Basic: greater than pH 7

pH below 7	**pH = 7**	**pH above 7**
0		**14**
More Acidic	**Neutral**	**More Basic**

In examining the definition of pH, we find that there is an inverse relationship between hydrogen ion concentration and pH. As $[H^+]$ increases, pH decreases and vice versa. Also, notice that each increase of one unit on the pH scale corresponds to a tenfold decrease in $[H^+]$, and each decrease of one unit on the pH scale corresponds to a tenfold increase in $[H^+]$. If this is true, then as the $[H^+]$ increases, $[HO^-]$ decreases, and as $[H^+]$ decreases, $[OH^-]$ increases.

9-14. pOH SCALE

a. pOH range: 0 -14

b. Neutral: pOH 7

c. Acidic: greater than pOH 7

d. Basic: less than pOH 7

pOH below 7	**pOH = 7**	**pOH above 7**
0		**14**
More Basic	**Neutral**	**More Acidic**

9-15. RELATIONSHIP BETWEEN pH AND pOH

$K_w = [H^+] [OH^-] = 1.0 \times 10^{-14}$

$pK_w = pH + pOH$

$14 = pH + pOH$

a. **Example 1**. What is the pOH of a H_2SO_4 solution that has a pH of 4.3?

Solution. Select the expression that allows you to solve for the unknown quantity.

$14 = pH + pOH$

Substitute the given information, and solve for the unknown quantity.

$14 = 4.3 + pOH$

$pOH = 14 - 4.3$

$pOH = \underline{9.7}$

b. **Example 2**. What is the pH of a 0.010 mol/L NaOH solution?

Solution. In solving this problem, it is necessary to first determine the pOH of the solution before solving for the pH.

$$pOH = \log \frac{1}{0.010} = 2.00$$

Now, utilize the expression that relates pH to pOH by substituting the given information and solving for the unknown quantity.

$14 = pH + pOH$

$14 = pH + 2.00$

$pH = 14 - 2.00$

$pH = \underline{12.00}$

Section V. HYDROLYSIS OF SALTS

9-16. DISCUSSION

If you were to place a salt in water and then perform a pH determination, you would be surprised to find that some salts would have an acidic pH, some would be basic, and others would be neutral. Most salts are formed by the reaction of acids with bases. The anion (-) of the acid and the cation (+) of the base are the component parts of a salt molecule. The term hydrolysis refers to a reaction with water. A salt when placed in water will ionize and depending on the type of salt, will undergo hydrolysis (remember water itself ionizes to a small degree). There are three types of salts we will consider: salts that are composed of the ions of a weak acid and a strong base, salts of a strong acid and a weak base, and salts of a weak acid and a weak base. Anions that form a strong acid (NO_3^-, Cl^-, $C10_4^-$, Br^- and I^-) and cations that form strong bases (Group IA and IIA) do not undergo hydrolysis but simply dissociate into ions: (i.e., NaCl does not hydrolyze in water). This information will help in understanding the action of buffers.

9-17. HYDROLYSIS OF THE SALT OF A WEAK ACID AND A STRONG BASE

This salt will hydrolyze to form a basic solution. The salt sodium acetate will ionize into sodium ions and acetate ions. The sodium ion does not undergo hydrolysis but the acetate ion will react with free hydrogen ions forming acetic acid.

$Na^+ + C_2H_3O_2^- + HOH <====> HC_2H_3O_2 + Na^+ + OH^-$

This reaction causes a shift in the $[H^+]/[OH^-]$ by removing $[H^+]$. This salt is classified as a basic salt because the acetate ion, by definition, is a weak Bronsted-Lowry base (proton acceptor).

9-18. HYDROLYSIS OF THE SALT OF A STRONG ACID AND A WEAK BASE

The ions of this type of salt will hydrolyze to form an acidic solution. The salt ammonium chloride reacts as follows:

$NH_4^+ + Cl^- + HOH <====> NH_3 + H_3O^+ + Cl^-$

In this case, the ammonium cation is a Bronsted-Lowry acid (proton donor). This salt is classified as an acidic salt.

9-19. HYDROLYSIS OF THE SALT OF A WEAK ACID AND A WEAK BASE

The ions of this type of salt will hydrolyze to form a solution that may be acidic, basic, or neutral, depending on the strength (ionization constant) of the weak acid and weak base that formed the salt. If the ionization constant of the weak acid is greater than the ionization constant of the weak base, the solution will be acidic. If the constant is greater for the weak base, the solution will be basic. If the constants are equal, the resulting solution will be neutral.

Section VI. BUFFERS

9-20. DISCUSSION

Buffer systems are commonly used in the laboratory to help maintain a constant pH in a reaction mixture. A buffer solution acts to resist a change in pH. A buffer is composed of either a weak acid and its salt or a weak base and its salt. Buffers made up of a weak acid and it's salt are called acidic buffers and function from pH 0 to 7. Basic buffers consist of a weak base and its salt and function from pH 7 to 14.

9-21. MECHANISM OF ACTION

a. **Addition of a Strong Acid.** A common buffer used in the laboratory is the acetic acid/acetate buffer. As acid is added $[H^+]$ into the solution, it will react with the acetate anion from the salt (proton acceptor) forming acetic acid. Since acid is added, a shift in the equilibrium of the solution will occur, but the amount of free hydrogen ions in solution will be controlled by the ionization constant of the acetic acid formed and only a slight change in $[H^+]$ takes place.

Buffer system:

$HC_2H_3O_2/NaC_2H_3O_2$ (Acetic acid/sodium acetate)

Buffer system with HCl added:

$H^+ + Cl^- + Na^+ + C_2H_3O_2^- \text{<====>} HC_2H_3O_2 + Cl^- + Na^+$

b. **Addition of a Strong Base.** As base is added to this buffer, acetic acid reacts with the base to form salt (acetate ions) and water. Once again a shift in the equilibrium of the solution occurs but with a minimal change in pH.

NaOH added:

$Na^+ + OH^- + HC_2H_3O_2 \text{<====>} Na^+ + C_2H_3O_2^{\ -} + H_2O$

c. **Notes.**

(1) A buffer's ability to resist changes in pH is limited to the concentration of either the weak acid of base and its salt. The addition of an excessive amount of either acid or base will "exhaust" the buffering capacity of the buffer.

(2) The closer the pH of the buffer to the pK_a or pK_b of the weak acid or base respectively the greater the buffering capacity.

9-22. THE HENDERSON-HASSELBALCH EQUATION

The preparation of laboratory buffers can be accomplished by using the Henderson-Hasselbalch equation. The equation is derived from the ionization constant of weak acids and bases.

$$K = \frac{[H^+]\ [A^-]}{[HA]}$$

The hydrogen ion concentration can be calculated by rearranging the equation.

$$[H^+] = \frac{K_a \text{ X } [HA]}{[A^-]}$$

Because $[H^+]$ is usually expressed as pH, it is usually more convenient to express all concentration in this equation in logarithmic form.

$$\log [H^+] = \log K_a + \log \frac{[HA]}{[A^-]}$$

Since pH = $-\log [H^+]$ multiply both sides of the equation by -1.

$$-\log [H^+] = -\log K_a - \log \frac{[HA]}{[A^-]}$$

Restated:

$$-\log [H^+] = -\log K_a + \log \frac{[A^-]}{[HA]}$$

$$pH = pK_a + \log \frac{[salt]}{[acid]}$$

The expression relating pOH to the components of a buffer may be developed similarly.

$$pOH = pK_b + \log \frac{[salt]}{[base]}$$

9-23 SOLVING BUFFER PROBLEMS

The most important consideration when preparing a buffer is determining whether an acidic or basic buffer is being prepared. Examine the components of the buffer. If a weak acid and its salt are the components, use the expression for acidic buffers. If a weak base and its salt are used, select the expression for basic buffers.

a. **Example 1**. Calculate the pH of a buffer solution that contains 0.010 mol/L acetic acid and 0.020 mol/L sodium acetate. The K_a for acetic acid is 1.75×10^{-5}.

Solution. Read the problem carefully, and select the expression that allows you to solve for the unknown quantity.

$$pH = pK_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

Make any necessary conversions.

$$pK_a = -\log K_a$$

$$pK_a = -\log (1.75 \times 10^{-5})$$

$$pK_a = 4.76$$

Substitute the given information, and solve for the unknown quantity.

$$pH = 4.76 + \log \frac{0.020 \text{ mol/L}}{0.010 \text{ mol/L}}$$

$$pH = 4.76 + 0.301$$

$$pH = \underline{5.06}$$

b. **Example 2**. What is the pH of a buffer solution that contains 1.50×10^{-3} mol/L NH_4Cl and 2.00×10^{-4} mol/L NH_4OH? The pK_b for ammonium hydroxide is 4.75.

Solution Read the problem carefully, and select the expression that allows you to solve for the unknown quantity. Note that you are dealing with a basic buffer from which you must determine pH.

$$pOH = pK_b + \log \frac{[\text{salt}]}{[\text{base}]}$$

Substitute the given information, and solve for the unknown quantity.

$$\text{pOH} = 4.75 + \log \frac{1.50 \text{ X } 10^{-3} \text{ mol/L}}{2.00 \text{ X } 10^{-4} \text{ mol/L}}$$

pOH = 4.75 + 0.8751

pOH = 5.63

Now that you have determined the pOH, using the appropriate expression, determine the pH of the buffer.

14 = pH + pOH

pH = 14 - pOH

pH = 14 - 5.63

pH = <u>8.37</u>

c. **Example 3**. What amount of salt/acid is required to prepare a 0.100 mol/L acetate buffer at a pH of 5.50? The pK_a of acetic acid is 4.76.

Solution. At times, the pK of the acid/base of a buffer and the total concentration of the buffer is known, and the amount of salt and acid/base required to prepare a buffer must be calculated.

After reading the problem carefully, select the expression that allows you to solve for the unknown quantity.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

Substitute the given information, and solve for the unknown quantity.

$$5.50 = 4.76 + \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$5.50 - 4.76 = \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$0.74 = \log \frac{[\text{salt}]}{\text{acid}]}$$

$$\text{antilog } 0.74 = \frac{[\text{salt}]}{[\text{acid}]}$$

$$5.5 = \frac{[\text{salt}]}{[\text{acid}]} = \frac{5.5 \text{ mol/L}}{1 \text{ mol/L}}$$

You have now established the proper ratio of salt and acid that will yield the desired pH. However, the problem states that a molar concentration of 0.100 mol/L for the buffer is desired. This mol/L concentration is the sum of the mol/L concentrations of the salt and the acid. The mol/L concentration of the buffer, based on your calculations so far, is 6.5 mol/L. The following method is one of several that you may employ to determine the concentrations necessary to yield the proper mol/L concentration of the buffer while maintaining the proper ratio.

Express the desired mol/L concentration of the buffer as an equality.

$$[\text{salt}] + [\text{acid}] = 0.100 \text{ mol/L}$$

Solve the equality for either [salt] or [acid].

$$[\text{salt}] = 0.100 \text{ mol/L} - [\text{acid}]$$

Now based on the substitution property, you are able to express the [salt] in terms of [acid], and in effect, derive an equation in one variable instead of the two, in the original expression.

$$5.5 = \frac{[\text{salt}]}{[\text{acid}]}$$

$$5.5 = \frac{0.100 \text{ mol/L} - [\text{acid}]}{[\text{acid}]}$$

Simplify the expression by clearing the fraction.

$5.5\ [\text{acid}] = 0.100\ \text{mol/L} - [\text{acid}]$

Solve the expression for [acid].

$5.5\ [\text{acid}] + [\text{acid}] = 0.100\ \text{mol/L}$

$6.5\ [\text{acid}] = 0.100\ \text{mol/L}$

$$[\text{acid}] = \frac{0.100\ \text{mol/L}}{6.5}$$

$[\text{acid}] = 0.0154\ \text{mol/L}$

or equivalently, <u>$1.54 \text{ X } 10^{-2}$ mol/L</u>

To determine the amount of salt required, substitute the calculated value of the acid concentration in the expression relating the salt and acid concentration to the desired mol/L concentration of the buffer.

$[\text{salt}] = 0.100\ \text{mol/L} - [\text{acid}]$

$[\text{salt}] = 0.100\ \text{mol/L} - 0.0154\ \text{mol/L}$

$[\text{salt}] = 0.0846\ \text{mol/L}$

or equivalently, <u>8.46×10^{-2} mol/L</u>

d. **Example 4**. What amount of salt/base is required to prepare a 0.100 mol/L ammonia buffer at a pH of 9.8? The pK_b for ammonium hydroxide is 4.75.

Solution. After reading the problem carefully, note that you are to prepare a basic buffer of a certain pH. Select an expression that will allow you to solve for the unknown quantity.

$$\text{pOH} = pK_b + \log \frac{[\text{salt}]}{[\text{base}]}$$

Convert pH to pOH.

$14 = \text{pH} + \text{pOH}$

$\text{pOH} = 14 - \text{pH}$

$\text{pOH} = 14 - 9.8$

$\text{pOH} = 4.2$

Substitute the given values, and solve for the unknown quantity.

$$4.2 = 4.75 + \log \frac{[\text{salt}]}{[\text{base}]}$$

$$4.2 - 4.75 = \log \frac{[\text{salt}]}{[\text{base}]}$$

$$-0.6 = \log \frac{[\text{salt}]}{[\text{base}]}$$

$$\text{antilog}\ (-0.6) = \frac{[\text{salt}]}{[\text{base}]}$$

$$0.3 = \frac{[\text{salt}]}{[\text{base}]}$$

Express the desired mol/L concentration of the buffer as an equality.

$[\text{salt}] + [\text{base}] = 0.100\ \text{mol/L}$

Solve the equality for either [salt] or [base].

$[\text{salt}] = 0.100\ \text{mol/L} - [\text{base}]$

Now, based on the substitution property, you are able to express the [salt] in terms of [base], and in effect derive an equation in one variable instead of the two in the original expression.

$$0.3 = \frac{[\text{salt}]}{[\text{base}]}$$

$$0.3 = \frac{0.100\ \text{mol/L} - [\text{base}]}{[\text{base}]}$$

Simplify the expression by clearing the fraction.

$0.3\ [\text{base}] = 0.100\ \text{mol/L} - [\text{base}]$

Solve the expression for [base].

0.3 [base] + [base] = 0.100 mol/L

1.3 [base] = 0.100 mol/L

$$[\text{base}] = \frac{0.100 \text{ mol/L}}{1.3}$$

[base] = 0.077 mol/L

or equivalently, $\underline{7.7 \times 10^{-2} \text{ mol/L}}$

To determine the amount of salt required, substitute the calculated value of the base concentration in the expression relating the salt and base concentration to the desired mol/L concentration of the buffer.

[salt] = 0.100 mol/L - [base]

[salt] = 0.100 mol/L - 0.077 mol/L

[salt] = $\underline{0.023 \text{ mol/L}}$

or equivalently, $\underline{2.3 \times 10^{-2} \text{ mol/L}}$

Continue with Exercises

EXERCISES, LESSON 9

INSTRUCTIONS: Answer the following exercises by writing the answer in the space provided at the end of the question.

After you have completed all the exercises, turn to "Solutions to Exercises" at the end of the lesson and check your answers. For each exercise answered incorrectly, reread the material referenced with the solution.

1. What is the pH of a 4.00 X 10^{-4} mol/L HCl solution?

2. What is the pH of a 0.100 mol/L $HC_2H_3O_2$ that ionizes 1.3

3. What is the pOH of a 0.250 mol/L H_2SO_4 solution? (**NOTE:** Use the given concentration for pH and pOH calculations that involve polyprotic acids, that is, acids with more than one hydrogen.)

4. What is the pH of a 1.40 X 10^{-2} mol/L $Ca(OH)_2$ solution? (**NOTE:** To determine the actual hydroxide concentration of this and like compounds, multiply the given concentration by the number of hydroxide ions per molecule.)

5. What is the pH of a buffer prepared by adding 1.20×10^{-3} mol/L $NaC_2H_3O_2$ and 1.20×10^{-2} mol/L $HC_2H_3O_2$? The K_a for acetic acid is 1.75×10^{-5}.

6. What is the pOH of a buffer prepared by adding 3.50×10^{-5} mol/L NH_4OH and 2.50×10^{-4} mol/L NH_4Cl? The pK_b for NH_4OH is 4.75.

7. What is the pH of the buffer in exercise 6?

8. An acetate buffer, pH 5.20, was prepared using 0.100 mol/L acetic acid. The pK_a for acetic acid is 4.76. What concentration of salt is needed?

9. A 0.050 mol/L bicarbonate buffer with a pH of 5.60 is to be prepared. The pK_a of carbonic acid is 6.12. What concentration of sodium bicarbonate and carbonic acid are needed to prepare the buffer?

10. A 0.0670 mol/L buffer, pH 7.10 is to be prepared. The pK_a of KH_2PO_4 is 4.71. A volume of 1.00 liter is needed.

    a. What concentration of KH_2PO_4 and Na_2HPO_4 is required?

    b. What weight of each constituent is needed to prepare the buffer?

11. An ammonia buffer pH 7.60 with a concentration of 1.20×10^{-2} mol/L is required. The K_b for NH_4OH is 1.79×10^{-5}. The desired volume is 500 mL.

    a. What concentration of NH_4OH and NH_4Cl are required?

    b. What weight of NH_4Cl is needed?

    c. How many milliliters of NH_4OH with a S.G. 0.950 and percent purity of 28.0

Check Your Answers on Next Page

SOLUTIONS TO EXERCISES, LESSON 9

1. pH = 3.40 (para 9-11)

2. pH = 2.9 (para 9-11)

3. pOH = 13.7 (para 9-12)

4. pH = 12.4 (para 9-15)

5. pH = 3.76 (para 9-23)

6. pOH = 5.60 (para 9-23)

7. pH = 8.40 (para 9-23)

8. [salt] = 0.275 mol/L (para 9-23)

9. [salt] = 0.012 mol/L
   [acid] = 0.038 mol/L (para 9-23)

10. a. [acid] = 0.000272 mol/L
       [salt] = 0.0667 mol/L (para 9—23)

    b. Acid = 0.0370 g
       Salt = 9.47 g (lesson 3)

11. a. [base] = 0.000261 mol/L
       [salt] = 0.0117 mol/L (para 9—23)

    b. 0.313 g (lesson 3)

    c. 0.0172 mL (para 9-23 and lesson 8)

End of Lesson 9

APPENDIX A

REVIEW OF DIMENSIONAL ANALYSIS

1. Read the problem carefully. What is the problem asking for? Be sure the entire problem has been read and understood. This may require you to read the problem two or three times. <u>YOU CANNOT ANSWER THE PROBLEM IF YOU DO NOT KNOW WHAT IT IS ASKING</u>!!!!

2 Determine exactly what results are to be produced by the calculations.

3. Determine what principles and relationships are involved.

4. Think about possible methods to use in solving the problem.

5 Use the sample problems to help you set up and solve the problem.

6. Based on definition, determine the appropriate factors that allow you to solve for the unknown quantity.

7. Once you have selected the appropriate factors for that specific problem type, write them down on your paper.

8. Units are treated the same as numbers in any mathematical calculations.

9. Write the intermediate stages of the calculations clearly. Avoid writing one number on top of another as a method of correction. Make each digit legible. This will allow you to go back and check your work later.

10. Mentally estimate an answer before working the problem.

11. Do the mathematics involved and check your work. Do not round off any intermediate calculations. Be extremely careful in positioning the decimal point and make certain the final answer has the appropriate number of significant figures.

12. Cancel units. The units you have left should be an appropriate unit for what the problem asked. Example: If the problem asked for "How many grams," your final answer should be in grams. If it is not, go back and check your work. Often, all that is required is a simple metric conversion.

13. Compare the calculated result with your estimated answer. If the two figures disagree drastically, determine which result is wrong.

14. Finally, go back and read the problem again. Did you answer the question correctly and does your answer make sense?

15. Ratio and proportion is consistent with and is the basis of dimensional analysis.

16. Example problems will serve as a reference to the various problem solving techniques.

End of Appendix A

APPENDIX B

TABLE OF FOUR-PLACE LOGARITHMS

	0	1	2	3	4	5	6	7	8	9
10	0000	0043	0086	0128	0170	0212	0253	0294	0334	0374
11	0414	0453	0492	0531	0569	0607	0645	0682	0719	0756
12	0792	0828	0864	0899	0934	0969	1004	1038	1072	1106
13	1139	1173	1206	1239	1271	1303	1335	1367	1399	1430
14	1461	1492	1523	1553	1584	1614	1644	1673	1703	1732
15	1761	1790	1818	1847	1875	1903	1931	1959	1987	2014
16	2041	2068	2095	2122	2148	2175	2201	2227	2253	2279
17	2304	2330	2355	2380	2405	2430	2455	2480	2504	2529
18	2553	2577	2601	2625	2648	2672	2695	2718	2742	2765
19	2788	2810	2833	2856	2878	2900	2923	2945	2967	2989
20	3010	3032	3054	3075	3096	3118	3139	3160	3181	3201
21	3222	3243	3263	3284	3304	3324	3345	3365	3385	3404
22	3424	3444	3464	3483	3502	3522	3541	3560	3579	3598
23	3617	3636	3655	3674	3692	3711	3729	3747	3766	3684
24	3802	3820	3838	3856	3874	3892	3909	3927	3945	3962
25	3979	3997	4014	4031	4048	4065	4082	4099	4116	4133
26	4150	4166	4183	4200	4216	4232	4249	4265	4281	4298
27	4314	4330	4346	4362	4378	4393	4409	4425	4440	4456
28	4472	4487	4502	4518	4533	4548	4564	4579	4594	4609
29	4624	4639	4654	4669	4683	4698	4713	4728	4742	4757
30	4771	4786	4800	4814	4829	4843	4857	4871	4886	4900
31	4914	4928	4942	4955	4969	4983	4997	5011	5024	5038
32	5051	5065	5079	5092	5105	5119	5132	5145	5159	5172
33	5185	5198	5211	5224	5237	5250	5263	5276	5289	5302
34	5315	5328	5340	5353	5366	5378	5391	5403	5416	5428
35	5441	5453	5465	5478	5490	5502	5514	5527	5539	5551
36	5563	5575	5587	5599	5611	5623	5635	5647	5658	5670
37	5682	5694	5705	5717	5729	5740	5752	5763	5775	5786
38	5798	5809	5821	5832	5843	5855	5866	5877	5888	5899
39	5911	5922	5933	5944	5955	5966	5977	5988	5999	6010
40	6021	6031	6042	6053	6064	6075	6085	6096	6107	6117
41	6128	6138	6149	6160	6170	6180	6191	6201	6212	6222
42	6232	6243	6253	6263	6274	6284	6294	6304	6314	6325
43	6335	6345	6355	6365	6375	6385	6395	6405	6415	6425
44	6435	6444	6454	6464	6474	6484	6493	6503	6513	6522

TABLE OF FOUR-PLACE LOGARITHMS (Continued)

	0	1	2	3	4	5	6	7	8	9
45	6532	6542	6551	6561	6571	6580	6590	6599	6609	6618
46	6628	6637	6646	6656	6665	6675	6684	6693	6702	6712
47	6721	6730	6739	6749	6758	6767	6776	6785	6794	6803
48	6812	6821	6830	6839	6848	6857	6866	6875	6884	6893
49	6902	6911	6920	6928	6937	6946	6955	6964	6972	6981
50	6990	6998	7007	7016	7024	7033	7042	7050	7059	7067
51	7076	7084	7093	7101	7110	7118	7126	7135	7143	7152
52	7160	7168	7177	7185	7193	7202	7210	7218	7226	7235
53	7243	7251	7259	7267	7275	7284	7292	7300	7308	7316
54	7324	7332	7340	7348	7356	7364	7372	7380	7388	7396
55	7404	7412	7419	7427	7435	7443	7451	7459	7466	7474
56	7482	7490	7497	7505	7513	7520	7528	7536	7543	7551
57	7559	7566	7574	7582	7589	7597	7604	7612	7619	7627
58	7634	7642	7649	7657	7664	7672	7679	7686	7694	7701
59	7709	7716	7723	7731	7738	7745	7752	7760	7767	7774
60	7782	7789	7796	7803	7810	7818	7825	7832	7839	7846
61	7853	7860	7868	7875	7882	7889	7896	7903	7910	7917
62	7924	7931	7938	7945	7952	7959	7966	7973	7980	7987
63	7993	8000	8007	8014	8021	8028	8035	8041	8048	8055
64	8062	8069	8075	8082	8089	8096	8102	8109	8116	8122
65	8129	8136	8142	8149	8156	8162	8169	8176	8182	8189
66	8195	8202	8209	8215	8222	8228	8235	8241	8248	8254
67	8261	8267	8274	8280	8287	8293	8299	8306	8312	8319
68	8325	8331	8338	8344	8351	8357	8363	8370	8376	8382
69	8388	8395	8401	8407	8414	8420	8426	8432	8439	8445
70	8451	8457	8463	8470	8476	8482	8488	8494	8500	8506
71	8513	8519	8525	8531	8537	8543	8549	8555	8561	8567
72	8573	8579	8585	8591	8597	8603	8609	8615	8621	8627
73	8633	8639	8645	8651	8657	8663	8669	8675	8681	8686
74	8692	8698	8704	8710	8716	8722	8727	8733	8739	8745
75	8751	8756	8762	8768	8774	8779	8785	8791	8797	8802
76	8808	8814	8820	8825	8831	8837	8842	8848	8854	8859
77	8865	8871	8876	8882	8887	8893	8899	8904	8910	8915
78	8921	8927	8932	8938	8943	8949	8954	8960	8965	8971
79	8976	8982	8987	8993	8998	9004	9009	9015	9020	9026

TABLE OF FOUR-PLACE LOGARITHMS (Concluded)

	0	1	2	3	4	5	6	7	8	9
80	9031	9036	9042	9047	9053	9058	9063	9069	9074	9079
81	9085	9090	9096	9101	9106	9112	9117	9122	9128	9133
82	9138	9143	9149	9154	9159	9165	9170	9175	9180	9186
83	9191	9196	9201	9206	9212	9217	9222	9227	9232	9238
84	9243	9248	9253	9258	9263	9269	9274	9279	9284	9289
85	9294	9299	9304	9309	9315	9320	9325	9330	9335	9340
86	9345	9350	9355	9360	9365	9370	9375	9380	9385	9390
87	9395	9400	9405	9410	9415	9420	9425	9430	9435	9440
88	9445	9450	9455	9460	9465	9469	9474	9479	9484	9489
89	9494	9499	9504	9509	9513	9518	9523	9528	9533	9538
90	9542	9547	9552	9557	9562	9566	9571	9576	9581	9586
91	9590	9595	9600	9605	9609	9614	9619	9624	9628	9633
92	9638	9643	9647	9652	9657	9661	9666	9671	9675	9680
93	9685	9689	9694	9699	9703	9708	9713	9417	9722	9727
94	9731	9736	9741	9745	9750	9754	9759	9763	9768	9773
95	9777	9782	9786	9791	9795	9800	9805	9809	9814	9818
96	9823	9827	9832	9836	9841	9845	9850	9854	9859	9863
97	9868	9872	9877	9881	9886	9890	9894	9899	9903	9908
98	9912	9917	9921	9926	9930	9934	9939	9943	9948	9952
99	9956	9961	9965	9969	9974	9978	9983	9987	9991	9996

End of Appendix B

APPENDIX C

TABLE OF COMMON ELEMENTS AND RADICALS

NAME	SYMBOL	APPROX. ATOMIC WEIGHT	COMMON VALENCE
Aluminum	Al	27.0	+3
Antimony	Sb	121.8	+3
Arsenic	As	74.9	+3
Barium	Ba	137.3	+2
Bismuth	Bi	209.0	+3
Bromine	Br	79.9	-1
Calcium	Ca	40.1	+2
Carbon	C	12.0	+4
Chlorine	Cl	35.5	-1
Chromium	Cr	52.0	+3, +6
Copper	Cu	63.5	+1, +2
Fluorine	F	19.0	-1
Gold	Au	197.0	+3
Hydrogen	H	1.0	+1
Iodine	I	126.9	-1
Iron	Fe	55.8	+2, +3
Lead	Pb	207.2	+2, +4
Lithium	Li	6.9	+1
Magnesium	Mg	24.3	+2
Manganese	Mn	54.9	+2
Molybdenum	Mo	95.9	+6
Mercury	Hg	200.6	+1, +2
Nickel	Ni	58.7	+2, +3
Nitrogen	N	14.0	+3, -3
Oxygen	O	16.0	-2
Phosphorus	P	31.0	+5
Potassium	K	39.1	+1
Silicon	Si	28.1	+4
Silver	Ag	107.9	+1

NAME	SYMBOL	APPROX. ATOMIC WEIGHT	COMMON VALENCE
Sodium	Na	23.0	+1
Strontium	S	87.6	+2
Sulfur	S	32.1	-2, +6
Tin	Sn	118.7	+2, +4
Tungsten	W	183.9	+6
Zinc	Zn	65.4	+2
Ammonium	NH_4		+1
Acetate	$C_2H_3O_2$		-1
Bicarbonate	HCO_3		-1
Bisulfate	HSO_4		-1
Carbonate	CO_3		-2
Chlorate	ClO_3		-1
Dichromate	Cr_2O_7		-2
Hydroxide	OH		-1
Hypochlorite	ClO		-1
Nitrate	NO_3		-1
Nitrite	NO_2		-1
Oxalate	C_2O_4		-2
Permanganate	MnO_4		-1
Phosphate	PO_4		-3
Monohydrogen phosphate	HPO_4		-2
Dihydrogen phosphate	H_2PO_4		-1
Sulfate	SO_4		-2
Sulfite	SO_3		-2
Tungstate	WO_4		-2

APPENDIX C (concluded)

APPROXIMATE CONCENTRATIONS OF COMMERCIAL ACIDS AND BASES

ACID OR BASE	FORMULA	PERCENT (w/v)	S.G.	MOLARITY	NORMALITY
Acetic Acid	$HC_2H_3O_2$	99.6	1.05	17	17
Hydrochloric Acid	HCl	36	1.19	12	12
Lactic Acid	$HC_3H_5O_3$	85	1.20	11	11
Nitric Acid	HNO_3	70	1.42	16	16
Phosphoric Acid	H_3PO_4	85	1.71	15	45
Sulfuric Acid	H_2SO_4	97	1.84	18	36
Ammonium Hydroxide	NH_4OH	28 (NH_3)	0.90	15	15

End of Appendix C

APPENDIX D

METRIC PREFIXES

PREFIX	ABBREVIATION	SCIENTIFIC NOTATION	DECIMAL VALUE
Giga—	G	10^9	1,000,000,000
Mega—	M	10^6	1,000,000
Kilo—	k	10^3	1,000
Hecto—	h	10^2	100
Deka—	dk	10^1	10
(Base unit)	**(m, L, or g)**	$\mathbf{10^0}$	**1**
Deci—	d	10^{-1}	0.1
Centi—	c	10^{-2}	0.01
Milli—	m	10^{-3}	0.001
Micro—	u	10^{-6}	0.000001
Nano—	n	10^{-9}	0.000000001
Pico—	f	10^{-15}	0.000000000000001

End of Appendix D