Teared Pages Within

The Book Only

Problem Book for GENERAL CHEMISTRY

ROYCE H. LeROY
Professor of Chemistry
A. and M. College of Texas

FIRST EDITION

New York Toronto London
McGRAW-HILL BOOK COMPANY, INC.
1951

PREFACE

It is intended that this problem book shall serve at least the following purposes: (1) provide the instructor with an additional source of problem material for use in the classroom and for study and homework assignments; (2) give some extra emphasis to the importance of problems in chemistry by collecting the various types into a volume separate from the textbook selected for the course; (3) provide the earnest student with an approach to the various types of problems that is simple, direct, and perhaps slightly different from that of the textbook; and (4) furnish to the student whose background in arithmetic is weak, an opportunity to review some of the more useful fundamentals of arithmetic along with their application to the problems of chemistry. Most teachers of general and inorganic chemistry believe that it is impossible to develop in the student a real appreciation of the intrinsic and applicational values of the subject unless the descriptive and theoretical material is accompanied by an adequate treatment of the simple numerical and mathematical relations that constitute the framework of the course. Working problems and handling problem material are fundamental in demonstrating the laws of nature as they apply in the transformations of matter.

The introductory chapter is composed of a brief review of some of the fundamental operations of arithmetic—handling of fractions, decimal numbers, and some use of exponential expressions. The inclusion of this material seems desirable because the background of the average freshman is rather inadequate for the mastering of the fairly simple calculations encountered in the general course in college chemistry.

Each chapter introduces only one type of calculation, useful in its own right or characteristic of chemistry in its applications. This makes for simplicity and flexibility. No recommendations are offered as to the sequence of assignments because each instructor will follow a routine that is best suited to the achievement of his class. As the work progresses, the last problems of a given section will include a reference to, or an application combined with, some previous type of problem. Thus it is hoped that the student will realize that the problems of chemistry are general rather than specific and in a practical sense are composed of combinations of types.

While significance of figures has been observed in presenting problem data and in the answer section, no obvious attempt has been made to impress the student with this principle. The author feels that each instructor should develop this principle according to his own desires and the needs and appreciation of his class; and further, he feels that in many cases it may be more important to emphasize the principles of chemistry in the calculations rather than the numerical precision.

No attempt has been made to develop all possible types of problems that could be introduced into the general chemistry course. Only the simplest examples are used in th chapters on Solubility Product and Ionization Constant. The calculation of the degree of ionizatio phenomena has been omitted.

The format of the "Problem Book for General Chemistry' sheets may be torn out and handed in for correction and grading. The d three-ri binder and so may be preserved *in toto* by the student

It is the hope of the author that he has succeede chemistry based on the simple principles of arithm fully through the book from beginning to end.

COLLEGE STATION, TEX
September, 1951

NOTE TO THE STUDENT

In preparing this problem book it has been assumed that you have a reasonable knowledge of ordinary arithmetic. No problem in the general course of chemistry requires any mathematics more complicated or advanced than that encountered in the ordinary problems of everyday life. Addition, subtraction, multiplication, and division, as commonly applied in calculations involving percentage and proportion, constitute what is called chemical calculations, or chemical arithmetic.

The introductory problems are simple applications of the above-mentioned processes, while the tougher problems are nothing more than combinations of two or more of the simple types. For the student whose curriculum requires more chemistry than the general or introductory course, a knowledge of the fundamental problem types and the methods of calculation presented herein is essential. It is hoped that the student whose curriculum requires only the general chemistry course may learn something of orderly thinking and reasoning by solving these problems.

This book is designed to supplement the text, not to replace any part of it. The problems presented here are no different in type or degree of difficulty from those in your text. The method of presenting a type of problem may differ somewhat from that in your text. If so, seeing that problem presented in two different ways may help you master it more quickly.

Following the introductory statement about the specific problem being presented, examples are used to illustrate the method of calculation. These example calculations should be studied carefully. The first few problems that follow are relatively simple and usually are very much like the examples. The problems then tend to become more general in nature and thus more thought-provoking. The last problems in the chapter may involve a combination with some type previously studied.

It is proper to give warning that the useful problem, corresponding to an actual situation in practice, is a general one and may involve combinations of several types. It will be necessary that certain princi-les be kept clearly in mind and that some vision of the significance of the order of consecutive calcula- e developed as the work in the course unfolds. Above all else you must realize that the principles s week are the foundations for the work of the following week, and that a problem type once ned is soon to become a tool in mastering another type.

CONTENTS

1. ARITHMETIC REVIEW

Part 1

It is necessary that the student of general chemistry have a working knowledge of the fundamental operations of arithmetic, and it must be assumed that he has obtained a reasonable skill in making simple calculations before coming to college. Addition, subtraction, multiplication, and division with whole, decimal, and fractional numbers are of course the commonly used operations. Frequently essential to the easy solving of many problems are the methods of reducing, or simplifying, fractions; the conversion of fractional numbers to the decimal equivalents; the understanding and use of percentage calculations; the use of ratios; the devising of proportions and their use; and the manipulation of exponential numbers.

The problems provided in this section are offered as a brief review for those students who show weakness in the application and use of the ordinary operations of arithmetic.

NAME________________________________ DESK No. 242.

PROBLEMS

1. Convert each of the following fractions to the decimal equivalent. (NOTE: Make the indicated division, being careful to place the decimal point in the correct position.)

Example

$$\frac{25}{40} = \frac{25.000}{40} = 0.625$$

(*a*) $\frac{1}{3}$ =	(*b*) $\frac{1}{4}$ =	(*c*) $\frac{3}{8}$ =	(*d*) $\frac{6}{7}$ =
(*e*) $\frac{75}{80}$ =	(*f*) $\frac{55}{90}$ =	(*g*) $\frac{125}{275}$ =	(*h*) $\frac{72}{90}$ =
(*i*) $\frac{5}{8}$ =	(*j*) $\frac{3}{16}$ =	(*k*) $\frac{5}{9}$ =	(*l*) $\frac{11}{12}$ =

2. Simplify, or reduce, each of the following fractions. (NOTE: Find a common divisor for both the numerator and denominator, and then make the division, thus obtaining smaller and simpler numbers for the fractional expression.)

Example

$$\frac{125}{200} = \frac{125 \div 25}{200 \div 25} = \frac{5}{8}$$

(*a*) $\frac{625}{1000}$ =	(*b*) $\frac{325}{625}$ =	(*c*) $\frac{72}{288}$ =	(*d*) $\frac{72}{81}$ =
(*e*) $\frac{60}{96}$ =	(*f*) $\frac{875}{1000}$ =	(*g*) $\frac{56}{196}$ =	(*h*) $\frac{1020}{1400}$ =
(*i*) $\frac{26}{104}$ =	(*j*) $\frac{45}{225}$ =	(*k*) $\frac{273}{364}$ =	(*l*) $\frac{665}{760}$ =

3. Find the sum of each of the following. (NOTE: Addition of fractions requires the use of a common denominator, as shown in the example.)

Example

$$\frac{2}{3} + \frac{7}{11} = \frac{22 + 21}{33} = \frac{43}{33} = 1\tfrac{10}{33}$$

(*a*) $\frac{5}{9} + \frac{3}{4} =$ (*b*) $\frac{3}{7} + \frac{5}{6} =$ (*c*) $\frac{9}{50} + \frac{17}{20} =$ (*d*) $\frac{15}{32} + \frac{9}{16} =$

(*e*) $\frac{13}{25} + \frac{21}{10} =$ (*f*) $3\frac{1}{8} + 2\frac{5}{16} =$ (*g*) $7\frac{3}{4} + 5\frac{3}{16} =$ (*h*) $\frac{83}{144} + 6\frac{3}{18} =$

4. Obtain the difference in each of the following. (NOTE: Subtraction of fractions involves the use of a common denominator, as shown in the example.)

Example

$$\frac{17}{20} - \frac{9}{50} = \frac{85 - 18}{100} = \frac{67}{100} = 0.67$$

(*a*) $\frac{3}{4} - \frac{1}{16} =$ (*b*) $\frac{72}{90} - \frac{15}{36} =$ (*c*) $\frac{81}{144} - \frac{5}{16} =$ (*d*) $\frac{155}{288} - \frac{15}{72} =$

(*e*) $\frac{85}{256} - \frac{5}{32} =$ (*f*) $8\frac{7}{36} - 3\frac{11}{12} =$ (*g*) $11\frac{2}{3} - 7\frac{3}{64} =$ (*h*) $9\frac{1}{6} - 2\frac{5}{8} =$

NAME ______________________________ DESK NO. ______________

5. Obtain the product of each of the following multiplication exercises, and simplify the answer when possible.

Example

$$\frac{2}{3} \times \frac{7}{8} = \frac{2 \times 7}{3 \times 8} = \frac{14}{24} = \frac{7}{12}, \text{ or } 0.583$$

(*a*) $\frac{5}{8} \times \frac{7}{15} =$	(*b*) $\frac{3}{4} \times \frac{5}{16} =$	(*c*) $\frac{9}{21} \times \frac{17}{6} =$	(*d*) $\frac{3}{5} \times \frac{7}{8} =$
(*e*) $\frac{21}{30} \times \frac{45}{25} =$	(*f*) $\frac{72}{90} \times \frac{24}{56} =$	(*g*) $\frac{700}{760} \times \frac{273}{546} =$	(*h*) $\frac{19}{25} \times \frac{45}{60} =$

6. Division of fractions is performed by first inverting the divisor and then multiplying as in the preceding set of multiplication exercises.

Example

$$\frac{19}{36} \div \frac{21}{48} = \frac{19}{36} \times \frac{48}{21} = \frac{19 \times 48}{36 \times 21} = \frac{19 \times 4}{3 \times 21} = \frac{76}{63} = 1.206+$$

(*a*) $\frac{1}{2} \div \frac{3}{4} =$	(*b*) $\frac{3}{8} \div \frac{7}{32} =$	(*c*) $\frac{9}{16} \div \frac{3}{40} =$	(*d*) $\frac{96}{120} \div \frac{7}{15} =$
(*e*) $1\frac{36}{64} \div 2\frac{3}{16} =$	(*f*) $8\frac{9}{11} \div 6\frac{3}{21} =$	(*g*) $12\frac{3}{4} \div 18\frac{5}{6} =$	(*h*) $1\frac{16}{17} \div 3\frac{2}{3} =$

Part 2

In making calculations which involve the handling of either very large or very small numbers, it frequently is convenient to convert such numbers to the exponential form in order to expedite the manipulations. The exponential form of a number is nothing more or less than the expression of that number as a multiple of the appropriate power of 10. For example, 100 becomes 10^2, or 1×10^2, and 0.001 becomes 10^{-3}, or 1×10^{-3}.

Tables 1 and 2 show the relation between several numbers and their expression as multiples of a power of 10. In the left column are given various round numbers along with their exponential equivalents, while in the column at the right are shown examples of the conversion of ordinary numbers to the exponential form.

Immediately following the tables are illustrations of several operations concerned with the handling of exponential numbers. Careful study of these examples and frequent comparison with them as patterns should be made while learning to make calculations with numbers in the exponential form.

Finally, a few problems are presented for drill in learning the techniques of manipulation. Note particularly the relation between the power of 10 in the expression and the placing of the decimal point.

TABLE 1

$10 = 10^1 = 1 \times 10^1$	$70 = 7 \times 10^1$
$100 = 10^2 = 1 \times 10^2$	$730 = 7.3 \times 10^2$
$1000 = 10^3 = 1 \times 10^3$	$1460 = 1.46 \times 10^3$
$10{,}000 = 10^4 = 1 \times 10^4$	$52{,}500 = 5.25 \times 10^4$
$100{,}000 = 10^5 = 1 \times 10^5$	$640{,}000 = 6.4 \times 10^5$
$1{,}000{,}000 = 10^6 = 1 \times 10^6$	$8{,}700{,}000 = 8.7 \times 10^6$
$10{,}000{,}000 = 10^7 = 1 \times 10^7$	$195{,}000{,}000 = 1.95 \times 10^8$

TABLE 2

$0.1 = 10^{-1} = 1 \times 10^{-1}$	$0.16 = 1.6 \times 10^{-1}$
$0.01 = 10^{-2} = 1 \times 10^{-2}$	$0.023 = 2.3 \times 10^{-2}$
$0.001 = 10^{-3} = 1 \times 10^{-3}$	$0.0072 = 7.2 \times 10^{-3}$
$0.0001 = 10^{-4} = 1 \times 10^{-4}$	$0.00054 = 5.4 \times 10^{-4}$
$0.00001 = 10^{-5} = 1 \times 10^{-5}$	$0.000068 = 6.8 \times 10^{-5}$
$0.000001 = 10^{-6} = 1 \times 10^{-6}$	$0.000{,}009{,}6 = 9.6 \times 10^{-6}$
$0.000{,}000{,}1 = 10^{-7} = 1 \times 10^{-7}$	$0.000{,}000{,}31 = 3.1 \times 10^{-7}$

Examples

1. Multiplication of pure exponential numbers involves simply the addition of the exponents.

(*a*) $10^2 \times 10^3 = 10^{2+3} = 10^5$ (*b*) $10^{-3} \times 10^{-4} = 10^{(-3)+(-4)} = 10^{-7}$

(*c*) $10^7 \times 10^{-2} = 10^{7+(-2)} = 10^{7-2} = 10^5$ (*d*) $10^{-25} \times 10^8 = 10^{-17}$

2. Multiplication of the more complex exponentials involves two operations: additions of the exponents and separate multiplication of the multiplicands.

(*a*) $(3 \times 10^3)(5 \times 10^2) = (3 \times 5) \times 10^{3+2} = 15 \times 10^5 = 1.5 \times 10^6$

(*b*) $(9 \times 10^{-2})(6 \times 10^7) = (9 \times 6) \times 10^{(-2)+7} = 54 \times 10^5 = 5.4 \times 10^6$

(*c*) $(2.1 \times 10^{-5})(4 \times 10^{-3}) = (2.1 \times 4) \times 10^{(-5)+(-3)} = 8.4 \times 10^{-8}$

3. Division of pure exponential numbers involves the subtraction of the exponent of the divisor from the exponent of the dividend.

(*a*) $10^7 \div 10^3 = 10^{7-3} = 10^4$ (*b*) $10^{18} \div 10^{12} = 10^{18-12} = 10^6$

(*c*) $10^7 \div 10^{-3} = 10^{7-(-3)} = 10^{10}$ (*d*) $10^{12} \div 10^{18} = 10^{12-18} = 10^{-6}$

4. Division of the more complex exponential numbers involves two operations: subtraction of the exponent of the divisor from the exponent of the dividend and proper division of the multiplicands.

(*a*) $(8 \times 10^4) \div (2 \times 10^2) = \frac{8}{2} \times 10^{4-2} = 4 \times 10^2$

(*b*) $(16 \times 10^{15}) \div (4 \times 10^{12}) = \frac{16}{4} \times 10^{15-12} = 4 \times 10^3$

(*c*) $\dfrac{2.5 \times 10^6}{5 \times 10^{-2}} = 0.5 \times 10^{6-(-2)} = 0.5 \times 10^8 = 5 \times 10^7$

(*d*) $\dfrac{6.4 \times 10^{-12}}{1.6 \times 10^{-10}} = 4.0 \times 10^{-12-(-10)} = 4.0 \times 10^{-2}$

5. Note the simplicity of raising a pure exponential to a higher power of 10, such as squaring, cubing, etc.

(*a*) $(10^2)^2 = 10^2 \times 10^2 = 10^{2+2} = 10^4$

(*b*) $(10^2)^3 = 10^2 \times 10^2 \times 10^2 = 10^6$

(*c*) $(10^5)^2 = 10^{5+5} = 10^{10}$

(*d*) $(10^4)^4 = 10^{4+4+4+4} = 10^{16}$

6. Taking roots of pure exponentials is just as easy as the previously shown operation.

(*a*) $\sqrt[2]{10^{16}} = 10^{16\div 2} = 10^8$

(*b*) $\sqrt[3]{10^{15}} = 10^{15\div 3} = 10^5$

(*c*) $\sqrt[4]{10^{12}} = 10^{12\div 4} = 10^3$

(*d*) $\sqrt[2]{10^{-4}} = 10^{-4\div 2} = 10^{-2}$

(*e*) $\sqrt[5]{10^{-10}} = 10^{-10\div 5} = 10^2$

(*f*) $\sqrt[10]{10^{20}} = 10^{20\div 10} = 10^2$

7. Taking roots of the more complex exponential numbers is shown in the following examples:

(*a*) $\sqrt[2]{4 \times 10^4} = \sqrt[2]{4}\sqrt[2]{10^4} = 2 \times 10^2$

(*b*) $\sqrt[3]{64 \times 10^{12}} = \sqrt[3]{64}\sqrt[3]{10^{12}} = 4 \times 10^4$

(*c*) $\sqrt[2]{1.6 \times 10^5} = \sqrt[2]{16 \times 10^4} = 4 \times 10^2$

(*d*) $\sqrt[3]{2.7 \times 10^{10}} = \sqrt[3]{27 \times 10^9} = 3 \times 10^3$

NAME ______________________________ DESK No. ______________

PROBLEMS

Perform the operations as indicated, and express your answers in the exponential form.

1. $0.0003 \times 0.000076 =$

2. $14{,}000 \times 8100 =$

3. $\dfrac{(2.7 \times 10^{-6})}{(15 \times 10^{10})} =$

4. $175 \div 0.00025 =$

5. $(6.2 \times 10^{7})(3.8 \times 10^{-3}) =$

6. $\sqrt[3]{10^{21}} =$

7. $\sqrt[2]{6.25 \times 10^{6}} =$

8. $(6 \times 10^{3})^{2}(3 \times 10^{2})^{3} =$

9. $\sqrt[3]{1.25 \times 10^{8}} =$

10. $\dfrac{(5 \times 10^{4})^{2}}{(4 \times 10^{6})^{2}} =$

11. A speed of 90 m.p.h. is the same as how many inches per hour?

12. A room measures 20 ft. in length, 12 ft. in width, and the ceiling height is 8 ft. The room contains how many cubic inches?

13. $(10^2)(10^3)(10^4) =$

14. $\dfrac{(10^2)(10^4)(10^6)}{(10^3)(10^5)} =$

15. $\dfrac{(1.8 \times 10^{24})}{(7.2 \times 10^{13})} =$

16. $\dfrac{(1.6 \times 10^{-5})^2}{(9.6 \times 10^{-2})} =$

17. $\dfrac{(2 \times 10^{-32})}{(4.8 \times 10^{-5})^2} =$

18. $\dfrac{(0.00134)^2}{0.09866} =$

19. $\left(\dfrac{0.00251}{233}\right)^2 =$

20. $(\sqrt[2]{1.21 \times 10^{-10}})(170) =$

21. $(2 \times 0.0016)^2(0.0016) =$

22. $\dfrac{(2.3 \times 10^{-10})(137.4)}{(0.1 \times 0.75)} =$

23. $323\sqrt[2]{1.69 \times 10^{-14}} =$

24. $\dfrac{(1.4 \times 10^{-7})(87.6)}{(0.2 \times 0.65)} =$

25. $(3 \times 10^{-5})^3(2 \times 10^{-5})^2 =$

26. $\dfrac{10^{-14}}{0.01 \times 0.0042} =$

2. THE METRIC SYSTEM OF MEASUREMENT

The metric system is used in all branches of science and is the official system of measurement in nearly all the countries of the world except Great Britain and the United States. It is a decimal system and derives its name *metric* from the basic unit of length, the meter. The units of volume and of weight are derived indirectly from the basic unit.

In the *linear* table which follows, note the decimal relation of the units, *i.e.*, the ten-times and the one-tenth relations.

$$1 \text{ meter} = 10 \text{ decimeters} = 100 \text{ centimeters} = 100 \text{ millimeters } 1000$$
$$= \tfrac{1}{10} \text{ decameter} = \tfrac{1}{100} \text{ hectometer} = \tfrac{1}{1000} \text{ kilometer}$$

Note that the prefix attached to the stem, meter, indicates the relation of that unit to the basic unit. Another and more common way of representing this system of units is as follows:

10 millimeters = 1 centimeter

10 centimeters = 1 decimeter

10 decimeters = 1 meter

10 meters = 1 decameter

10 decameters = 1 hectometer

10 hectometers = 1 kilometer

The great advantage that the metric system offers is that conversion of one unit to another becomes merely a matter of properly placing the decimal point to the right or left, according to the desire to express the figure in smaller or larger units, respectively.

The meter is equal to 39.37 in., or roughly $3\frac{1}{2}$ in. longer than the yard. One centimeter, being $\frac{1}{100}$ m., is $\frac{1}{100} \times 39.37$ in., or 0.3937 in. One inch is equal to $1/39.37 \times 100$ cm., or 2.54 cm. (see the Appendix for a table of useful metric-English equivalents).

The unit of *volume* is the liter and may be defined as 1 cubic decimeter, or 1000 cubic centimeters. The units of volume are related in the same way as the units of length.

10 milliliters = 1 centiliter

10 centiliters = 1 deciliter

10 deciliters = 1 liter

10 liters = 1 decaliter

10 decaliters = 1 hectoliter

10 hectoliters = 1 kiloliter

By way of comparison with the English units, 1 liter = 1.056 qt.

The unit of *weight* is the gram and may be defined as the weight of 1 cubic centimeter of water, measured at the temperature at which water is the most dense. The same prefixes are used in relating the various units of the weight system as are used for the linear and volume systems. The pound (avoirdupois) is equal to 453.6 g.

NAME____ DESK No.____________________

PROBLEMS

1. How many centimeters (*a*) in 1 yd.; (*b*) in 1 mile?

2. The standard barometric pressure corresponds to the weight of a column of mercury 1 cm. square and 76 cm. high. Express this height (*a*) in inches; (*b*) in millimeters.

3. From the definition of the liter, calculate the number of liters in 1 cu. m.

4. The distance from New York to San Francisco is about 3000 miles. Express this distance in kilometers.

5. Express the English ton in kilograms.

6. There are 231 cu. in. in 1 gal. Express this volume (*a*) in cubic centimeters; (*b*) in liters.

7. A water tank holds 8000 gal. This is how many (*a*) liters; (*b*) cubic meters?

8. When identical volumes of limestone and water are compared, limestone weighs 2.6 times as much as the water. One cubic foot of water weighs 62.4 lb. One cubic foot of limestone weighs how many (*a*) pounds; (*b*) kilograms?

9. A room measures 20 by 15 by 10 ft. What is its volume in cubic meters?

10. What is the metric equivalent, in liters, of (*a*) a 55-gal. barrel; (*b*) a 30-gal. drum?

11. The Olympic athletes run a 100-m. high-hurdle race, while in the American colleges, a 110-yd. high-hurdle race is run. Which of these two races is the longer, and by how much?

12. The 16-lb. shot used in the field event weighs (*a*) how many grams; (*b*) how many kilograms?

13. One liter of air weighs 1.293 g. What is the weight of the air contained in the room described in Problem 9?

14. A dump truck had a capacity for hauling 3 cu. yd. of dirt. This is how many (*a*) cubic feet; (*b*) liters; (*c*) kiloliters; (*d*) cubic meters?

15. The gasoline tank on a certain automobile has a capacity of 11.5 gal. This is how many liters?

16. Which is the longer, and by what distance, the 6-mile cross-country run, or the 10,000-m. Olympic race?

17. Compare in equivalent units the ton, the long ton, and the metric ton. (The metric ton is 1000 kg.)

18. Evaluate (*a*) the avoirdupois ounce in grams; (*b*) the cubic foot in liters.

19. A giant power shovel can take a bite of 32 cu. yd. of earth. This is how many cubic meters?

20. The speed of 600 m.p.h. is equivalent to how many kilometers per hour? Express this speed in centimeters per hour in exponential form.

21. Make a table of the units of weight in the metric system, similar to those given for the linear and volumetric systems.

3. THE FAHRENHEIT AND CENTIGRADE TEMPERATURE SCALES

To measure the intensity of the heat content of various bodies and systems, temperature scales, or thermometers, have been devised. The ordinary thermometer utilizes the temperature-expansion effect of some highly expansible liquid, such as mercury or alcohol, as a measure of the temperature of the material that is contacted by the thermometer. A graduated scale is engraved upon the glass tube holding the thermometric liquid, and the volume of the liquid is taken as an indication of the temperature of the system. The actual scales are arbitrary, and the zero readings were established according to the ideas of the person who first devised and used the scale.

The most commonly used thermometers are the *Fahrenheit* and the *centigrade*. The first is used in the household and in most industries, and the second finds its greatest use in the scientific laboratory. Comparison of the two scales shows that at the temperature of freezing water the centigrade thermometer registers 0° and the Fahrenheit 32°. And at the temperature of boiling water the centigrade registers 100°, while the Fahrenheit registers 212°. Thus the centigrade scale spans the difference in the two temperatures in 100 degrees, while the Fahrenheit requires 180. Obviously the centigrade degree is nearly twice as large as the Fahrenheit—in fact, $\frac{180}{100}$, or $\frac{9}{5}$ as large.

The difference in the size of the degrees and the fact that the Fahrenheit scale has its zero point 32 degrees below that of the centigrade have been taken into consideration in the development of the following formula which enables us to calculate equivalent readings on the two scales:

$$°\text{C.} = (°\text{F.} - 32) \times \tfrac{5}{9} \qquad (1)$$

where °C. = the centigrade temperature and °F. = the Fahrenheit equivalent. This expression may be solved for °F.

$$°\text{F.} = \tfrac{9}{5}°\text{C.} + 32 \qquad (2)$$

Example. Ordinary room temperature is usually 72°F. What is the corresponding centigrade temperature?

Using the first expression and substituting 72 for °F. gives

$$°\text{C.} = (72 - 32) \times \tfrac{5}{9} = 40 \times \tfrac{5}{9} = 200/9 = 22.2°\text{C.}$$

NAME________________________________ DESK NO.________________

PROBLEMS

1. Convert each of the following Fahrenheit temperatures to the equivalent centigrade values: 15, 40, 70, 98.6, and 180°.

2. Convert the following centigrade temperatures to the Fahrenheit equivalent: −273, −10, 56, 78, and 400°.

3. The melting points of iron, lead, tin, gold, and bismuth are, respectively, 1535, 327.5, 231.8, 1063, and 271°C. What are the Fahrenheit equivalents?

4. Is there any temperature where the Fahrenheit and centigrade thermometers would give the same reading? If so, what is it?

5. At what temperature will the Fahrenheit thermometer read (*a*) exactly twice that of the centigrade; (*b*) three times that of the centigrade?

6. The temperature of the electric arc is often given as 3500°C. What would this temperature be on the Fahrenheit thermometer?

7. The melting points of copper, zinc, silver, and mercury are, respectively, 1981.5, 786.9, and −38.02°F. Calculate the corresponding centigrade values.

8. The critical temperatures of ammonia, carbon dioxide, sulfur dioxide, and freon are, respectively, 132.9, 31.0, 157.12, and 111.7°C. Determine these values on the Fahrenheit temperature scale.

9. An important temperature in the theory of both chemistry and physics is −273°C. What is the Fahrenheit equivalent?

10. A color scale of temperature is used in connection with certain types of furnace operations. The scale is described in the terms dull red, cherry red, orange, yellow, white, and brilliant white. The approximate temperature of the mid-point of each of these ranges of heat is, respectively, 715, 825, 1040, 1160, 1245, and 1400°C. Calculate the corresponding Fahrenheit temperatures.

4. DENSITY AND SPECIFIC GRAVITY

Density

The *density* of a substance is expressed as the weight of a unit volume, and is given in the simple relation,

$$D = \frac{W}{V}$$

where D = density, W = weight, and V = volume of the sample. Various units of measurement can be used to express the density of a substance, but such expressions should be uniform with respect to the system of units used. Thus the density of water is expressed in metric units as 1 g./ml. and in the English system as 62.4 lb./cu. ft., or as 7.33 lb./gal. Table 3 gives the densities of a few substances.

TABLE 3. DENSITIES OF SOME GASES, LIQUIDS, AND SOLIDS

Gases, 0°C., 760 mm.*	*D*, g./liter	*Liquids*	*D*, g./ml.
Nitrogen	1.250	Ethyl alcohol	0.789
Air	1.293	Water	1.0
Oxygen	1.429	Hydrochloric acid (conc.)	1.19
Sulfur dioxide	2.860	Nitric acid (conc.)	1.42
Chlorine	3.214	Sulfuric acid (conc.)	1.84

Solids	*D*, g./cc.	*D*, lb./cu.ft.
Limestone	2.6	162.24
Lead	11.33	707
Iron	7.86	490.5
Sulfur	2.06	

* In the Appendix the densities of several additional gases are given.

Note that in the above illustrations the units are uniform with respect to the system of measurement used, metric units of weight used with metric units of volume, and English units of weight with English units of volume.

Examples

1. An irregular piece of metal weighed 40.79 g. and displaced 3.6 ml. of water when dropped into a graduated cylinder containing a known volume of water. Calculate the density of the metal.

Solution. The volume of the metal is equal to the volume of water which it displaces. That is, the metal caused the meniscus of the water to rise to a higher level when it was dropped into the known volume in the cylinder. The difference in the levels of the meniscus, before and after dropping in the piece of metal, gives the volume of the metal. Then

$$D = \frac{W}{V} = \frac{40.79 \text{ g.}}{3.6 \text{ ml.}} = 11.33 \text{ g./ml.}$$

2. A 25-ml. sample of a liquid weighed 22.345 g. What was the density of the liquid?

Solution. Substituting the given values of the weight and volume in the formula,

$$D = \frac{W}{V} = \frac{22.345 \text{ g.}}{25 \text{ ml.}} = 0.894 \text{ g./ml.}$$

(NOTE: If 25 ml. weighs 22.345 g., would not 1 ml. weigh one twenty-fifth as much?)

Specific Gravity

When the density of a substance such as limestone is compared with that of water,

$$\frac{D_{\text{limestone}}}{D_{\text{water}}} = \frac{2.6 \text{ g./cc.}}{1 \text{ g./cc.}} = 2.6$$

we find that the limestone is 2.6 times as heavy as the water, *per unit volume.* Note that in this indicated calculation the units in which the densities are expressed disappear in the calculation itself, and the answer obtained is a pure number, which can be identified only by the definition of specific gravity. Note also that this ratio, or comparative density, is the same when the comparison is made using the densities expressed in English units.

$$\frac{D_{\text{limestone}}}{D_{\text{water}}} = \frac{162.24 \text{ lb./cu. ft.}}{62.4 \text{ lb./cu. ft.}} = 2.6$$

This ratio of the density of the limestone to the density of the water is called the *specific gravity* and is a very widely and commonly used expression. Since water is the most commonly used reference substance, we understand that *water is the reference substance* unless stated otherwise.

It is important to note, as in the above examples, that comparisons are made only when the densities of the substances being compared are expressed in the same system of units.

Dimensions

Accuracy in the mechanics of a calculation—multiplying, dividing, etc.—is of course necessary in working any problem. But mechanical accuracy alone is not enough. You must learn to use reason both in solving a problem and in testing or checking the solution you have obtained. A simple practice can be of great aid to you in this connection. It has to do with the dimensional units and conversion factors concerned in the calculation.

Consider the operation, 8 × 12 × 24 = 2304. As presented, these figures show nothing of the nature of the units that may be involved and indicate only a multiplication process. But if we show

$$8 \times 12 \times 24 \text{ ft.} = 2304 \text{ cu. ft., or } 2304 \text{ ft.}^3$$

then the product of the multiplication process is immediately identified as a volume. If these figures represent the dimensions of a room, 2304 cu. ft. is its volume.

If the original figures have the dimensions

$$8 \text{ doz.} \times 12 \frac{\text{units}}{\text{doz.}} \times 0.24 \frac{\text{cost}}{\text{unit}} = 23.04$$

then 23.04 becomes the cost of the 8-doz. units. Note that the multiplication process remains the same, and further, that the dimensional units blend into the product and furnish us with a correctly labeled answer.

$$\cancel{\text{Dozen}} \times \frac{\cancel{\text{units}}}{\cancel{\text{dozen}}} \times \frac{\text{cost}}{\cancel{\text{unit}}} = \text{total cost}$$

As another example, consider the calculations and the dimensions involved in determining the weight of a block of limestone which is 2 ft. thick, 3 ft. wide, and 8 ft. long. We have seen already that the specific gravity of limestone is 2.60. Now the volume of the block is

$$2 \times 3 \times 8 \text{ ft.} = 48 \text{ ft.}^3\text{, or } 48 \text{ cu. ft.}$$

From the table of densities (page 21) we know that 1 cu. ft. of water weighs 62.4 lb. Therefore the density of limestone is

$$(2.6 \times 62.4) \text{ lb./cu. ft.}$$

and so we calculate the weight of the block.

$$48 \text{ cu. ft.} \times (2.6 \times 62.4) \text{ lb./cu. ft.} = 7787 \text{ lb.}$$

Note that the dimensions are manipulated in exactly the same way as numerical figures in all calculations.

$$\text{~~Cubic feet~~} \times \frac{\text{pounds}}{\text{~~cubic feet~~}} = \text{pounds}$$

Thus, if you learn to label all the figures involved in a calculation, the dimensions will reduce through the regular operations of arithmetic to the proper dimension of the final answer. Needless to say, just as in learning any new concept, the application and use of *dimensional analysis* can be learned only through diligent practice.

NAME____ DESK NO.____________

PROBLEMS

1. What is the density of a liquid, 65 ml. of which weighed 58.5 g.?

2. A sample of sulfuric acid weighed 190 g. and measured 112 ml. in volume. What is its density?

3. A block of limestone measured $2\frac{1}{4}$ by $3\frac{1}{2}$ by $8\frac{1}{2}$ ft. Using the figure given for the density of limestone in Table 3, calculate the weight of the block.

4. The directions for a laboratory experiment required 100 g. of concentrated nitric acid, the density of which is given in Table 3. What is the volume of this amount of the acid?

5. What volume does 1750 g. of oxygen occupy?

6. In determining the density of an irregularly shaped solid, the following data were collected:

Weight of the sample	19.5 g.
Volume of water in the graduated cylinder	31.6 ml.
Volume of water plus that of the sample	34.1 ml.

What was its density?

7. A sample of gas was accurately measured and weighed. Its volume was 330 ml., and its weight was 1.046 g. Express the density of the gas in grams per liter.

8. Using the density given in Table 3, calculate the volume of 500 g. of concentrated hydrochloric acid.

9. One cubic foot of a certain wood weighed 45 lb. What is the weight of 1 cc. of this wood?

10. Cork has a specific gravity of 0.2404. What is the weight of (*a*) 1 cc. of cork in grams; (*b*) 1 cu. ft. of cork in pounds?

11. Refer to the table of densities, and calculate the specific gravity of (*a*) oxygen, (*b*) chlorine, and (*c*) sulfur dioxide, as compared with air.

12. Using the English units, compute the specific gravity of (*a*) lead; (*b*) iron.

13. What is the volume of 500 g. of concentrated sulfuric acid?

14. The specific gravity of sulfur is 2.06. How much does 1 cu. ft. of sulfur weigh? What volume has 1 ton of sulfur?

15. A certain wood has a density of 48 lb./cu. ft. A block of the wood 4 in. by 8 in. by 12 ft. would have what weight?

5. APPLICATION OF PERCENTAGE CALCULATIONS

It is often convenient to recalculate a ratio or a fraction to some standard basis for comparison purposes (usually the base is 100 parts of the whole), rather than leave such ratios or fractions in the form in which they are first obtained. Laboratory samples for analysis may be of various weights, and the fractional parts that are indicated by the analysis may be difficult to compare and check when left in the form expressed by the actual data.

Example. Three separate samples of cupric oxide were analyzed for the copper content, and the results were:

1. 4.240 g. of cupric oxide gave 3.386 g. of copper.
2. 5.565 g. of cupric oxide gave 4.445 g. of copper.
3. 2.675 g. of cupric oxide gave 2.137 g. of copper.

Comparison of these results does not reveal any significant relation. However, recalculation of these figures on the basis of 100 parts of cupric oxide shows

(1) $$3.386:4.240 = X:100$$

or $$4.240 \times X = 100 \times 3.386$$

and $$X = \frac{338.6}{4.240} = 79.89 \text{ parts copper in 100 parts copper oxide}$$

(2) $$4.445:5.565 = X:100$$

or $$X = \frac{444.5}{5.565} = 79.89 \text{ parts copper in 100 parts copper oxide}$$

(3) $$2.137:2.675 = X:100$$

or $$X = \frac{213.7}{2.675} = 79.89 \text{ parts copper in 100 parts copper oxide}$$

From these figures it can be seen that the separate analyses agree, *viz.*, that 79.89 parts of copper are present in every 100 parts of cupric oxide. In other words, when calculations are made on the basis of a standard number of parts of the whole, namely, 100, accurate comparisons are possible and previously obscure relations usually become apparent.

The term *per cent* is derived from the Latin expression, per centum, which literally means by the hundred, or per 100. Thus 1 part per 100 parts ($\frac{1}{100}$, or 0.01) is expressed as 1 per cent, or 1%. Again, 12 parts per 100 parts ($\frac{12}{100}$, or 0.12) is 12 per cent, or 12%. The results obtained above, then, show that cupric oxide is actually 79.89% copper.

The form of the calculation may be varied somewhat. Thus 3.386 parts of copper:4.240 parts of cupric oxide is commonly expressed

$$\frac{3.386}{4.240} = 0.7989$$

and performing the indicated division converts the fraction into its decimal equivalent, 0.7989, which is taken as 79.89%.

NAME____ DESK No.____________________

PROBLEMS

1. A man bought a home for $7250 and some time later sold it for $8000. What per cent of his investment did he realize as profit in these transactions?

2. A man bought a home for $7250, and exactly two years later he sold it for $8250. During the two years he owned the place he paid a total of $192 taxes, painted the house at a cost of $275, and redecorated the interior at a cost of $317. What was the percentage profit or loss in his transactions?

3. Samples were taken from a 28,560-lb. carload of wheat, and analysis showed the moisture content of the wheat to be 4.63 per cent. What was the weight of the moisture contained in the wheat?

4. A sample of coal taken for analysis weighed 4.866 g. After heating and igniting, the ash remaining weighed 0.149 g. What per cent of the coal was ash?

5. An empty crucible weighed 13.383 g. With a sample of coal in it, the weight was 17.383 g. After heating and igniting, the crucible with the ash in it weighed 13.608 g. What was the per cent of ash that the coal contained?

6. A man has an income of $140 per year from an amount of money which he has invested at $3\frac{1}{2}$ per cent per annum. What is his capital investment?

7. To heat his home through the winter a man bought $5\frac{1}{2}$ tons of coal. When he shut down his furnace in the spring, he found that he had 300 lb. of coal left and an accumulation of 540 lb. of ashes. What per cent of his coal was ash?

8. The compound mercuric oxide is 7.34 per cent oxygen and can be completely decomposed into mercury and oxygen by heating to a sufficiently high temperature. What weight of this compound would contain 140 g. of combined oxygen?

9. The compound potassium chlorate is 39.34 per cent oxygen by weight. All of the oxygen can be released from this compound by proper heating. What weight of potassium chlorate should be heated to release 7.868 g. of oxygen?

10. What weight of oxygen can be prepared by heating 33.24 g. of potassium chlorate? (Use the data of the previous problem.)

NAME________________________________ DESK No.______________

11. After 8.400 g. of calcium carbonate (limestone) was heated to 1000°C., it was found that the residue weighed 4.704 g. What per cent of the compound was volatile (had evaporated)?

12. From the following data, calculate the per cent of oxygen which was released by heating a sample of barium peroxide.

Weight of empty crucible	8.716 g.
Weight of crucible + barium peroxide	12.302
Weight of crucible + residue, after heating	11.968

13. What weight of pure hydrogen peroxide does 952 g. of a 3 per cent solution of hydrogen peroxide actually contain?

14. A sample of potassium nitrate was cautiously heated until no further evolution of oxygen was noted. From the tabulated data, determine the percentage loss in weight of the sample due to the liberation of oxygen by the heating process.

Weight of empty test tube	32.185 g
Weight of tube + potassium nitrate	35.263
Weight of tube + residue, after heating	34.776

15. If the oxygen released in the experiment described in Problem 14 was only one-third of the total oxygen actually contained in the compound, (*a*) what per cent of the original potassium nitrate is really oxygen; (*b*) what per cent of the residue in Problem 14 would be oxygen?

16. It is desired to prepare 450 g. of a 5 per cent solution of bromine in water. What are the weights of bromine and water to be used?

17. The compound barium dioxide is 18.93 per cent oxygen. When heated, only one-half of the oxygen is liberated. What weight of oxygen will be released when 10 g. of the barium dioxide is heated to decomposition?

18. Concentrated sulfuric acid has a density of 1.84 g./ml. and is actually 96 per cent pure acid by weight. What weight of acid does 1 liter of the concentrated acid solution contain?

19. A certain mineral assays 60 per cent hematite. Hematite is an oxide of iron, which is actually 70 per cent iron. What weight of the mineral must be taken to obtain 1 ton of iron?

20. Sodium iodate occurs in the mineral Chile saltpeter as an impurity to the extent of 0.2 per cent. Sodium iodate is 64 per cent iodine by weight. How much iodine is actually present in 1 ton of the saltpeter?

6. CALCULATION OF THE MOLECULAR WEIGHT OF A COMPOUND FROM ITS FORMULA

The formula of a compound shows the number of atoms of each element combined in the molecule of the compound.

Consider the formula for potassium chlorate, $KClO_3$. This formula tells you that *each molecule* of this substance contains *one atom* of potassium, *one atom* of chlorine, and *three atoms* of oxygen. Now the chemist has assigned a definite weight to the atom of each element, and these atomic weights are given in your text and in a table in the Appendix. Referring to this table, we find for one atom of potassium 39.096, for one atom of chlorine 35.457, and for one atom of oxygen 16.000. It follows that the weight of any molecule is the sum of the weights of all the atoms of which it is composed. This last is simply a specific statement of a general truth: the weight of the whole is the sum of the weights of the component parts.

Examples

1. The weight of the molecule of potassium chlorate is (the weight of one potassium atom) + (the weight of one chlorine atom) + (the weight of three oxygen atoms) = 39.096 + 35.457 + 48 = 122.553.

2. The formula for the molecule of sodium chloride is NaCl. What is the molecular weight (weight of the molecule) of the compound?

Solution

$$\begin{aligned} \text{1 atomic weight of sodium, Na} &= 22.997 \\ \text{1 atomic weight of chlorine, Cl} &= 35.457 \\ \hline \text{Molecular weight (the sum)} &= 58.454 \end{aligned}$$

3. Given the formula of calcium nitrate, $Ca(NO_3)_2$, calculate the molecular weight.

Solution

$$\begin{aligned} 1Ca &= 1 \times 40.08 = 40.08 \\ 2N &= 2 \times 14.008 = 28.016 \\ 6O &= 6 \times 16.000 = 96.000 \\ \hline \text{Molecular weight} &= 164.024 \end{aligned}$$

4. The compound manganese sulfate has the formula $MnSO_4$. What is its molecular weight?

Solution

$$\begin{aligned} 1Mn &= 1 \times 54.93 = 54.93 \\ 1S &= 1 \times 32.06 = 32.06 \\ 4O &= 4 \times 16.00 = 64.00 \\ \hline \text{Molecular weight} &= 150.99 \end{aligned}$$

5. What is the molecular weight of the crystalline substance sodium sulfate decahydrate, $Na_2SO_4 \cdot 10H_2O$?

Solution. The formula shows the presence of 2 atoms of sodium; 1 atom of sulfur; 4 atoms of oxygen, O_4; and 10 molecules of water of crystallization containing a total of 20 atoms of hydrogen and 10 atoms of oxygen, $10H_2O$. To obtain the total weight of all these atoms,

$$\begin{aligned}
2Na &= 2 \times 22.997 = 45.994 \\
1S &= 1 \times 32.06 = 32.060 \\
4O &= 4 \times 16.000 = 64.000 \\
20H &= 20 \times 1.008 = 20.016 \\
10O &= 10 \times 16.000 = 160.000 \\
\text{Molecular weight} &= 322.164
\end{aligned}$$

NAME ____________________ DESK No. ____________

PROBLEMS

Calculate the molecular weight of each of the following:

1. Mercuric oxide, HgO

2. Ferrous sulfide, FeS

3. Water, H_2O

4. Sodium hydroxide, NaOH

5. Magnesium sulfate, $MgSO_4$

6. Magnesium chloride, $MgCl_2$

7. Sodium sulfate, Na_2SO_4

8. Potassium chlorate, $KClO_3$

9. Potassium bromate, $KBrO_3$

   167.012

10. Barium nitrate, $Ba(NO_3)_2$

11. Sucrose (cane sugar), $C_{12}H_{22}O_{11}$

12. Aluminum nitrate, $Al(NO_3)_3$

13. Aluminum sulfate, $Al_2(SO_4)_3$

14. Cupric sulfate, $CuSO_4$

15. Blue vitriol (copper sulfate pentahydrate), $CuSO_4 \cdot 5H_2O$

16. Ammonium sulfate, $(NH_4)_2SO_4$

17. Alum, $K_2Al_2(SO_4)_4 \cdot 24H_2O$

18. Sal soda (sodium carbonate decahydrate), $Na_2CO_3 \cdot 10H_2O$

19. Epsom salts, $MgSO_4 \cdot 7H_2O$

20. Trisodium phosphate nonahydrate, $Na_3PO_4 \cdot 9H_2O$

7. THE MOLE

The *gram-molecular weight* of any given substance is simply the molecular weight of that substance in grams. Thus the gram-molecular weight of sodium chloride is 58.454 g., and the gram-molecular weight of manganese sulfate is 150.99 g. (See Examples 2 and 4, Chap. 6, for calculation of these molecular weights.)

Now the gram-*mole*cular weight is a very important unit in chemistry, and soon you will see its wide use. It is used so much that the simple word, *mole*, has been coined as a contraction of the longer name. The chemist speaks of 1 mole of sodium chloride, meaning 1 gram-molecular weight, or 58.454 g. of the salt; 0.1 mole of calcium nitrate signifies one-tenth of a gram-molecular weight, or (0.1 × 164.024) g. of the substance; 4.5 moles of manganese sulfate indicates (4.5 × 150.99) g. of this material.

Examples illustrating the use of the *mole* follow.

Examples

1. What is the weight in grams of 2.5 moles of potassium chlorate, $KClO_3$?

Solution. In the preliminary discussion, page 33, potassium chlorate was found to have a molecular weight of 122.553. Therefore,

$$2.5 \text{ moles} = (2.5 \times 122.553) \text{ g.} = 306.382 \text{ g.}$$

2. How many moles is 1 lb. of barium chloride dihydrate, $BaCl_2 \cdot 2H_2O$?

Solution. Calculate the molecular weight of the substance.

$$\begin{array}{rcl} \text{1 atomic weight of barium} = 1 \times 137.36 & = & 137.36 \\ \text{2 atomic weights of chlorine} = 2 \times 35.457 & = & 70.914 \\ \text{4 atomic weights of hydrogen} = 4 \times 1.008 & = & 4.032 \\ \text{2 atomic weights of oxygen} = 2 \times 16.000 & = & 32.000 \\ \hline \text{Molecular weight} & = & 244.306 \end{array}$$

Now 1 lb. = 453.6 g. So the number of moles of barium chloride dihydrate, which equals 1 lb., or 453.6 g., will be

$$\frac{453.6 \text{ g./lb.}}{244.3 \text{ g./mole}} = 1.856 \text{ moles}$$

3. In a certain reaction, 0.4 mole of sodium bicarbonate (baking soda), $NaHCO_3$, was required. Express this quantity in grams.

Solution. Find the molecular weight of $NaHCO_3$.

$$\begin{array}{rcl} 1Na = 1 \times 22.997 & = & 22.997 \\ 1H = 1 \times 1.008 & = & 1.008 \\ 1C = 1 \times 12.010 & = & 12.010 \\ 3O = 3 \times 16.000 & = & 48.000 \\ \hline \text{Molecular weight} & = & 84.015 \end{array}$$

Then

$$0.4 \text{ mole} = (0.4 \times 84.015) \text{ g.} = 33.606 \text{ g.}$$

NAME____ DESK No.________________

PROBLEMS

1. How many grams does 10 moles of calcium carbonate, $CaCO_3$, contain?

Sol: Molecular wt: of $CaCO_3$ = 40+12+48 = 100
1 mole of $CaCO_3$ weighs 100 grams.
10 moles of " " 1000 grams = 1 kg. —— Ans.

2. Convert 0.75 mole of glucose, $C_6H_{12}O_6$, to grams.

Sol: ↳ 3/4 moles. Molecular wt of $C_6H_{12}O_6$ = 12(6) + 12(1) + 6(16)
= 72+12+96 180 gr
1 mole = 180 gr
3/4 mole = 3/4 × 180 gr = 135 grams. Ans.

3. How many moles is 100 g. of potassium chloride, KCl?

1: molecular wt of KCl - 39+35.5 = 74.5 g ⇒ 1 mole ≡ 74.5 g
x mole ≡ 100 gr
→ x = 100/74.5 moles. (Ans)

4. Convert 1 kg. of sodium nitrate, $NaNO_3$, to moles.

Sol:- Mol. wt = 23+14+48 = 85 grams ≡ 1 mole
1000 g ≡ 1000/85 moles = 200/17 = 11.76 moles — Ans

5. One-sixth mole of sodium carbonate, Na_2CO_3, equals how many grams?

Sol:- 1 mole = 2(23)+12+48 = 106 gram
1/6 mole = 17.66 grams —— (Ans)

6. One-sixth mole of sodium carbonate decahydrate, $Na_2CO_3 \cdot 10H_2O$, equals how many grams?

Sol:- 1/6 mole = 47.66 grams. Molecular wt = 2(23)+12+48 +10(18)
⇒ 46+60+180
= 286

7. Sucrose, commonly known as cane or beet sugar, is sold in 5-lb. bags at the grocery. How many moles does 5 lb. of sucrose correspond to? (Sucrose = $C_{12}H_{22}O_{11}$.)

Sol:

8. It is desired to prepare 1 liter of a solution that contains 3 moles of dissolved potassium hydroxide, KOH. How many grams of the potassium hydroxide is required? Find molarity = 3/1 = 3 moles/li

Sol: 168 grams.

9. (*a*) One liter of water, measured at 4°C., weighs 1000 g. (density = 1 g./ml.). This amount of water, H_2O, is how many moles? (*b*) At 25°C., the density of water is 0.99707 g./ml. How many moles does 1 liter of water at 25°C. correspond to?

No ob moles present = $\frac{1000}{18}$ moles. ✓

mass = volume × density = 997.07 grams

No. of moles = $\frac{997.07}{18}$ moles

10. A solution has a density of 1.2256 g./ml. and contains dissolved sodium nitrate, $NaNO_3$, 30 per cent by weight. How many moles of dissolved sodium nitrate does 1 liter of this solution contain?

mass of the solution = 1000 × 1.2256 g/ml × ml = 1225.6 g.

mass of $NaNO_3$ present 0.3 × 1225.6 g ≃ 368 grams

molecular wt = 23 + 14 + 48 = 85 grams

No of moles = $\frac{368}{85}$ moles

11. A phosphoric acid solution has a density of 1.526 g./ml. and contains 70 per cent H_3PO_4 by weight. How many moles of dissolved phosphoric acid does 1 liter of this solution contain?

mass of solution = 1526 grams mass of H_3PO_4 present = $\frac{7}{10}$ × 1526 grams.

molecular wt of H_3PO_4 =

No. of moles = $\frac{\text{mass}}{\text{mole wt.}}$

12. A hydrochloric acid solution contains 1.69 moles of dissolved HCl in 1 liter of solution. The density of the solution is 1.0279 g./ml. (*a*) How many grams of HCl is contained in 1 liter of this solution? (*b*) What is the per cent of HCl in the solution?

36.5 × 1.69 grams

percentage = $\frac{36.5 \times 1.69}{1027.9}$ × 100

13. A phosphate rock was found to be 84 per cent $Ca_3(PO_4)_2$ by weight. How many moles of the calcium phosphate does 1 kg. of the rock contain?

1000 g of Rock contain

14. A water-alcohol solution is 50 per cent alcohol by weight. The formula for alcohol is C_2H_6O. What is the ratio of moles of alcohol to moles of water in this solution?

8. THE PERCENTAGE COMPOSITION OF A COMPOUND FROM ITS FORMULA

The percentage composition of a compound can be calculated from its formula by means of the atomic weights.

In Chap. 6 the molecule of table salt, sodium chloride, was used as an example in illustrating the method of calculating the molecular weight of a compound. The figure obtained was 58.454. Of this total weight the sodium atom contributed 22.997 units. From this it is seen that the fraction of the total weight of the molecule which is due to sodium alone is

$$\frac{22.997}{58.454} = 0.3934, \text{ or } 39.34\%$$

The per cent of chlorine in the molecule is either (100 − 39.34) = 60.66 per cent, or

$$\frac{35.457}{58.454} = 0.6066, \text{ or } 60.66\%$$

As a second example, consider the compound potassium sulfate, K_2SO_4. The molecular weight (see Chap. 6) is 174.252, of which potassium contributed 78.129 parts of the total weight. Therefore the fraction of the whole weight which is due to potassium is

$$\frac{78.129}{174.252} = 0.4483, \text{ or } 44.83\%$$

Similarly for the sulfur,

$$\frac{32.06}{174.252} = 0.1839, \text{ or } 18.39\%$$

For the oxygen,

$$\frac{64.000}{174.252} = 0.3677, \text{ or } 36.77\%$$

As a final example consider the compound sodium sulfate decahydrate, $Na_2SO_4 \cdot 10H_2O$, whose molecular weight was calculated in Example 5, Chap. 6, and found to be 322.164. The water of crystallization is what per cent of the compound by weight?

The 10 molecules of water contribute ten times the molecular weight of water to the total weight of the crystalline molecule

$$\begin{aligned} 2H &= 2 \times 1.008 = 2.016 \\ 1O &= 1 \times 16.000 = 16.000 \\ \hline \text{Molecular weight} &= 18.016 \end{aligned}$$

and $10 \times 18.016 = 180.16$ (total weight of the 10 molecules of water of crystallization).

Now the fraction of the whole crystalline molecule that is water is given by

$$\frac{180.16}{322.164} = 0.5592, \text{ or } 55.92\%$$

NAME________________________ DESK NO.____________

PROBLEMS

1. Calculate the per cent of oxygen in each of the following compounds:

(*a*) Mercuric oxide, HgO

(*b*) Barium peroxide, BaO_2

(*c*) Potassium chlorate, $KClO_3$

(*d*) Manganese dioxide, MnO_2

(*e*) Lead dioxide, PbO_2

(*f*) Silicon dioxide, SiO_2

2. Calculate the per cent of hydrogen in each of the following compounds:

(*a*) Hydrogen chloride, HCl

(*b*) Sulfuric acid, H_2SO_4

(*c*) Sucrose, $C_{12}H_{22}O_{11}$

(*d*) Ethyl alcohol, C_2H_6O

(*e*) Hydrogen peroxide, H_2O_2

(*f*) Sodium hydroxide, NaOH

3. The following substances are minerals of iron. What is the per cent of iron in each?

(*a*) Pyrites, FeS_2

(*b*) Hematite, Fe_2O_3

(*c*) Magnetite, Fe_3O_4

(*d*) Siderite, $FeCO_3$

(*e*) Arsenopyrite, FeAsS

(*f*) Limonite, $2Fe_2O_3 \cdot 3H_2O$

4. What is the per cent of water in each of the following compounds? (These substances are called *hydrates*.)

(*a*) Barium chloride dihydrate, $BaCl_2 \cdot 2H_2O$

(*b*) Copper sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$

(*c*) Sodium carbonate decahydrate, $Na_2CO_3 \cdot 10H_2O$

(*d*) Ferrous sulfate heptahydrate, $FeSO_4 \cdot 7H_2O$

(*e*) Potassium alum, $K_2Al_2(SO_4)_4 \cdot 24H_2O$

NAME________________________________ DESK No.____________

5. Calculate the per cent of the element whose symbol is underlined in each of the following formulas:

(*a*) Potassium permanganate, $K\underline{Mn}O_4$

(*b*) Magnesium perchlorate, $Mg(\underline{Cl}O_4)_2$

(*c*) Sodium sulfate, $Na_2\underline{S}O_4$

(*d*) Cupric nitrate, $\underline{Cu}(NO_3)_2$

(*e*) Calcium sulfate dihydrate, $\underline{Ca}SO_4 \cdot 2H_2O$

(*f*) Lead acetate trihydrate, $\underline{Pb}(C_2H_3O_3)_2 \cdot 3H_2O$

6. The element sulfur, S, is burned in the contact catalytic process to produce sulfuric acid, H_2SO_4. If this process were 97 per cent efficient in converting sulfur to sulfuric acid, what weight of H_2SO_4 should be obtained from 1000 lb. of sulfur?

7. The mineral arsenopyrite has the formula FeAsS, and the operations in recovering the arsenic are only 85 per cent efficient. What weight of arsenic would be obtained from 1200 lb. of this mineral?

8. What weight of phosphorus, P, can be obtained from 1800 lb. of the mineral phosphorite, $Ca_3(PO_4)_2$, if the chemical recovery of the phosphorus is 92 per cent efficient?

9. The metal zinc is obtained by processing the mineral sphalerite, ZnS. (*a*) What weight of zinc should be obtainable from 1 ton of pure sphalerite? (*b*) If the sphalerite ore were only 83 per cent pure ZnS, how much zinc would be obtained from 2400 lb. of the ore?

10. Iron is mined as hematite, corresponding to the formula Fe_2O_3. How much iron should be obtained from 800 lb. of ore which is 78 per cent Fe_2O_3, if the recovery process is only 91 per cent efficient?

11. Gypsum, $CaSO_4 \cdot 2H_2O$, is heated to change it into plaster of paris and loses three-fourths of its water of crystallization in the process. (*a*) What is the percentage loss of weight in this process due to the removal of water? (*b*) What weight of gypsum must be heated to make 100 lb. of plaster of paris?

12. When the compound barium dioxide, BaO_2, is heated it loses exactly one-half its oxygen. (*a*) What is the percentage weight loss to be expected? (*b*) What weight of barium dioxide should be so heated to obtain 25 g. of oxygen?

9. CALCULATION OF A FORMULA FROM THE PERCENTAGE COMPOSITION OF A COMPOUND

We have seen that the percentage composition of a compound can be calculated from the formula when it is known and from the composition as determined by actual analysis. The reverse of this calculation is possible and has real significance. It is possible to calculate the ratio by which the atoms of the various elements are combined to form the molecule of the compound from the percentage figures with the aid of the accepted atomic weights. The following example illustrates the solution.

Examples

1. A compound of iron was analyzed and found to be 63.528 per cent iron and 36.472 per cent sulfur. What is the formula of the compound?

Solution. Assuming a 100-g. sample of the compound, there would be 63.528 g. of iron and 36.472 g. of sulfur combined.

STEP 1. Since the atomic weight of iron is 55.84, the number of gram-atomic weights of iron in the sample would be given by the expression, 63.528/55.84; similarly for the sulfur, the number of gram-atomic weights present in the sample would be given by 36.472/32.06. Using these quotients, we may write an elementary formula,

$$Fe_{\left(\frac{63.528}{55.84}\right)}S_{\left(\frac{36.472}{32.06}\right)}$$

(Remember that a formula indicates the ratio by which the component elements combine to form the molecule and should be expressed as simply as possible.)

STEP 2. Performing the indicated division gives

$$Fe_{1.137}S_{1.137}$$

STEP 3. Inspection of this last expression shows that the ratio between the iron and sulfur atoms in the molecule is simply 1:1, and so we write the formula

$$FeS$$

A second example will illustrate the need for accuracy in the calculations, and the value of carrying out these calculations to three and even four significant figures.

2. Analysis showed that 7.2026 g. of manganese was combined with 2.7974 g. of oxygen. Determine the formula of the compound.

Solution. STEP 1. Using the atomic weights of manganese and oxygen as 54.93 and 16, respectively, and proceeding as before, we write

$$Mn_{\left(\frac{7.2026}{54.93}\right)}O_{\left(\frac{2.7974}{16}\right)}$$

STEP 2. Performing the indicated division gives

$$Mn_{0.1311}O_{0.1748}$$

Clearly this is not a 1:1 ratio as in the preceding example. Nor can this be interpreted as a 1:2 ratio A ratio such as this can frequently be reduced to a more simple form by the use of a common divisor. These common divisors are not always easy to discover; so recourse to the decimal equivalent is usually the simplest solution

Dividing both of the ratio numbers by the larger of the two gives

$$Mn_{\left(\frac{0.1311}{0.1748}\right)}O_{\left(\frac{0.1748}{0.1748}\right)} = Mn_{0.75}O_{1.00}$$

STEP 3. Now it is obvious that the ratio between the manganese and oxygen in the molecule is 3:4, or the formula is

$$Mn_3O_4$$

(NOTE: The ratio 0.1311:0.1748 could have been simplified by dividing both of the figures by the smaller of the two instead of the larger. The reduced figures would then have been 1:1.333, and inspection shows that this ratio is the same as the one previously obtained, *viz.*, 3:4.)

NAME______________________________ DESK NO.______________

PROBLEMS

1. Analysis of an oxide of barium showed that it was 18.89 per cent oxygen (the rest was barium). What is the formula of the oxide?

2. Iron combines with sulfur to form a compound which is 46.55 per cent iron and 53.44 per cent sulfur. Calculate the formula.

3. One of the oxides of manganese is 63.19 per cent manganese by weight. What is its formula?

4. A compound was analyzed and found to be 52.17 per cent carbon, 13.03 per cent hydrogen, and 34.79 per cent oxygen. Determine the formula.

5. A compound was found to contain 29.08 per cent sodium, 40.56 per cent sulfur, and 30.35 per cent oxygen. What is its formula?

6. A sample of magnesium weighed 0.9728 g. It was heated slowly in an atmosphere of oxygen until all reaction had ceased. The resulting ash weighed 1.6128 g. Determine the formula of the oxide formed.

7. Determine the formula of a compound whose analysis showed it to contain 20.70 per cent sodium, 28.86 per cent sulfur, and 50.42 per cent oxygen, by weight.

8. Iron forms three oxides and their percentage compositions are (*a*) 77.727 per cent iron, 22.272 per cent oxygen; (*b*) 72.359 per cent iron, 27.641 per cent oxygen; (*c*) 70.00 per cent iron, 30.00 per cent oxygen. What is the formula of each?

9. The molecular weights of benzene and acetylene are 78.078 and 26.026, respectively. Each of these compounds is actually 92.25 per cent carbon and 7.74 per cent hydrogen by weight. What is the formula of each?

10. A certain acid was found to have its molecular weight between 75 and 100. Analysis showed that its composition was 2.23 per cent hydrogen, 26.67 per cent carbon, and 71.09 per cent oxygen. What is the correct formula of the acid?

10. THE EQUIVALENT WEIGHT OF AN ELEMENT AND VALENCE

The equivalent weight of an element has been defined as that amount which will combine with or displace 1.008 g. of hydrogen.

This definition immediately gives us a standard for measuring the combining and reacting capacity of any element. If the gram-atomic weight of an element reacts with or displaces exactly 1.008 g. of hydrogen, we say that its reacting capacity is the same as that of the standard, the gram-atomic weight of hydrogen, and that its valence number is 1. If the gram-atomic weight of an element reacts or combines exactly (2 × 1.008) g. of hydrogen, then its reacting capacity is twice that of the standard and its valence number is 2. Similarly when the gram-atomic weight of an element is found to be equivalent to (3 × 1.008) g. of hydrogen, then its valence number is 3. The *gram-atomic weight* of any element then is related to its *gram-equivalent weight* as

$$\frac{\text{Gram-atomic weight}}{\text{Gram-equivalent weight}} = \text{valence number}$$

Examples

1. Careful analysis of water shows that it is 11.1886 per cent hydrogen, the rest of the compound being oxygen. What is the equivalent weight of oxygen? What is the valence number of oxygen?

Solution. For a 100-g. sample of water, 11.1886 g. of hydrogen would be combined with (100.0000 − 11.1886) g. of oxygen. Therefore, 1.008 g. of hydrogen would be combined with a corresponding amount, or the gram-equivalent weight of oxygen is

$$\frac{1.008}{11.1886} \times 88.8114 = 8.000 \text{ g. oxygen}$$

Then
$$\frac{\text{Gram-atomic weight}}{\text{Gram-equivalent weight}} = \frac{16.000}{8.000} = 2 = \text{valence number of oxygen}$$

2. It was found that 9.7235 g. of chlorine was combined with 0.2765 g. of hydrogen. Calculate the gram-equivalent weight of chlorine.

Solution. If 0.2765 g. of hydrogen combines with 9.7235 g. of chlorine, 1.008 g. of hydrogen will combine with a corresponding quantity.

$$\frac{1.008}{0.2765} \times 9.7235 = 35.457 \text{ g., the gram-equivalent weight of chlorine}$$

Since 35.457 is also the gram-atomic weight of chlorine, the valence number of chlorine is 1.

3. In a reaction 1.0000 g. of the metal sodium displaced exactly 0.04373 g. of hydrogen. What is the gram-equivalent weight of sodium?

Solution. Since 0.04383 g. of hydrogen = 1.0000 g. of sodium, 1.008 g. of hydrogen will be equivalent to

$$\frac{1.008}{0.04383} \times 1.000 = 22.99 \text{ g., the gram-atomic weight of sodium}$$

4. Analysis of silver chloride shows it to be 75.263 per cent silver and 24.736 per cent chlorine. Using the gram-equivalent weight of chlorine obtained in Example 2, calculate the gram-equivalent weight of the silver.

Solution. Since chlorine has been standardized in terms of the *primary* standard, 1.008 g. of hydrogen, we now can use it as a *secondary* standard. (Things equal to the same thing are equal to each other.) So for a 100-g. sample of the silver chloride we would have

$$24.736 \text{ g. chlorine} : 75.263 \text{ g. silver}$$

and 35.457 g. chlorine: X g. silver

or
$$\frac{35.457 \times 75.263}{24.736} = 107.88 \text{ g. (gram-eauivalent weight of silver)}$$

Example 4 illustrates the fact that it may not always be convenient to measure or determine the equivalent weight of a given element by direct comparison with hydrogen. In such cases a study is made of the ratio by which the element displaces or combines with some other element whose equivalent weight already has been determined with great accuracy. Thus, the gram-equivalent weights of oxygen 8.000 g., chlorine 35.457 g., and silver 107.880 g., are commonly used to measure the gram-equivalent weights of other elements, and some of the following problems illustrate such use.

NAME____ DESK No.______________

PROBLEMS

1. A sample of magnesium weighing 2.1352 g. was allowed to react with hydrochloric acid, and 0.1770 g. of hydrogen was displaced. (*a*) Calculate the equivalent weight of the magnesium. (*b*) Compare the equivalent weight with the atomic weight of magnesium, and find the valence of the metal.

2. The analysis of the carefully prepared chloride of a newly discovered element showed it to be 24.737 per cent chlorine (the rest being the element). What was the equivalent weight of the element?

3. When 2.2677 g. of aluminum was heated in an atmosphere of pure oxygen, the resulting aluminum oxide weighed 4.2857 g. (*a*) Calculate the equivalent weight; (*b*) find the valence of aluminum.

4. A sample of mercurous oxide weighing 3.4766 g. was heated to decomposition, and 3.3433 g. of mercury (metal) was obtained. (*a*) Assume the loss in weight to be due entirely to oxygen, and find the equivalent weight of the mercury. (*b*) What per cent of mercurous oxide is oxygen?

5. A compound was found to be 82.247 per cent nitrogen and 17.752 per cent hydrogen. What is the equivalent weight of nitrogen in this compound?

6. A sample of a compound composed of silver and element X weighed 1.1737 g. Analysis showed that this sample contained 0.6743 g. of silver. What is the equivalent weight of X?

7. An oxide of manganese was found to be 77.44 per cent manganese and 22.56 per cent oxygen. (*a*) Calculate the equivalent weight of manganese; (*b*) find its valence number.

8. A 1.000-g. sample of a certain element was heated carefully in air until no further action was noted, and the resulting oxide was found to weigh 1.268 g. Assuming that the increase in weight was entirely due to combined oxygen, calculate the equivalent weight of the element.

9. What is the equivalent weight of iron in each of its oxides listed in Problem 8, Chap. 9?

10. A sample of copper weighing 1.6027 g. was dissolved in nitric acid to form copper nitrate. This copper nitrate then was heated in a crucible until it was converted to copper oxide. The resulting copper oxide weighed 2.0061 g. What is the equivalent weight of copper?

NAME____ DESK No.______________

11. The following data were obtained from a series of experiments in which an oxide of tin was prepared by heating the samples of tin with concentrated nitric acid. Calculate the average equivalent weight of tin.

	A	*B*	*C*	*D*
Weight of tin, g........	1.1250	1.5632	3.4283	2.7696
Weight of tin oxide, g...	1.4283	1.9846	4.3525	3.5175

12. A sample of lead weighed 3.525 g. It was heated in air at 700°C. and gradually changed to a yellow powder which weighed 3.797 g. Then the yellow powder was heated at 400° in a stream of oxygen and changed to a red powder which weighed 3.888 g. Finally the red oxide was treated with nitric acid, washed, and carefully dried. The resulting chocolate-brown oxide weighed 4.069 g. (*a*) Calculate the equivalent weight of the lead in each of the oxides, the yellow, the red, and the brown. (*b*) Calculate the percentage composition of each oxide. (*c*) Determine the formula of each oxide.

11. CALCULATIONS BASED UPON THE BALANCED CHEMICAL EQUATION

In Example 1, Chap. 9, the formula for iron sulfide was shown to be FeS. The equation for the reaction between iron and sulfur is written

$$Fe + S \rightarrow FeS$$

If we place the atomic weights and the molecular weight under the proper components, a significant relation is at once apparent.

$$\underset{55.84\ \text{g.}}{Fe} \quad + \quad \underset{32.06\ \text{g.}}{S} \quad \rightarrow \quad \underset{87.90\ \text{g.}}{FeS}$$

The equation now lends itself to the following interpretation: 55.84 parts of iron react with 32.06 parts of sulfur to form 87.90 parts of iron sulfide (parts by weight throughout). According to the law of definite composition, the ratio by weight between the iron and sulfur, which react to form the compound iron sulfide, will always be the same. This calls to our attention the fact that any excess of either of the reacting substances which may be present in the beginning will be left unchanged when the reaction has been concluded.

The ratios that are obtained from the *balanced equation* by the use of the proper atomic- and molecular-weight quantities are exact. Therefore it is possible to calculate the quantities of reacting substances which are required for a given reaction or the quantities of products resulting from that reaction.

Examples

1. Using the relations already indicated above, what quantity of sulfur is required to react with 2.000 g. of iron to form iron sulfide?

Solution. Since 55.84 g. of iron requires 32.06 g. of sulfur, then 2.000 g. of iron will require a corresponding quantity of sulfur, *i.e.*,

$$55.84:32.06 = 2.000:X$$

or

$$X = \frac{32.06 \times 2.000}{55.84} = 1.148 \text{ g. of sulfur required}$$

2. From the balanced equation for the decomposition of mercuric oxide, calculate the weight of the oxide which is required to furnish 10.00 g. of oxygen.

Solution:

$$2HgO \rightarrow 2Hg + O_2$$

$$2(200.6 + 16.00)\ \text{g.} \qquad 2 \times 16.00\ \text{g.}$$

or

$$433.2\ \text{g.} \qquad 32.00\ \text{g.}$$

therefore,

$$X\ \text{g.} \qquad 10.00\ \text{g.}$$

from which we write

$$433.2:32 = X:10$$

and solving,

$$X = \frac{433.2 \times 10}{32} = 135.3 \text{ g. mercuric oxide required}$$

Since the atomic weights, and hence the molecular weights of compounds, are pure weight-ratio numbers, they may be expressed in any system of weights. The gram-molecular weight of a compound is the molecular weight of that compound expressed in grams, a unit characteristic to the metric system. That molecular weight could be expressed in kilograms, pounds, ounces, or even in tons. Only one rule is necessary: the same unit of weight must be used throughout any one given problem or calculation. This means that the unit used in a calculation must be consistent and not mixed.

NAME________________________________ DESK No.____________

PROBLEMS

1. Calculate the weight of iron oxide which should be obtained by burning 12.5 g. of iron, according to the equation

$$3Fe + 2O_2 \rightarrow Fe_3O_4$$

(1) Fe_3O_4 produced
(2) O_2 required.
(3) we don't know molecular of Fe when x grams Fe used y grams of Fe formed then M = 64x/y.

(1) x amount of Fe
y amount of O_2 is given
required to find produced Fe_3O_4
which compound remain unreacted.
modification

2. Given the equation for the decomposition of potassium chlorate,

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

what weight of oxygen will be evolved when 10 g. of the potassium chlorate is heated to decomposition?

(a) 12. (b) 9

when you encounter objective question time is less and to do arithmetic

3. What weight of water must be decomposed by electrolysis to yield 100 g. of hydrogen?

$$2H_2O + \text{electric current} \rightarrow 2H_2 + O_2$$

How much O_2 is produced
we can also find current required
using Faraday laws of elect
1st law: $W = Z\phi$
2nd law: $\frac{W_1}{W_2} = \frac{E_1}{E_2}$

4. Calculate the weight of magnesium required to liberate 6.35 g. of hydrogen from hydrochloric acid, according to the equation

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

we can also find HCl required
$MgCl_2$ produced.

5. (*a*) What weight of sulfur is needed to react with 3.25 lb. of copper in the preparation of copper sulfide ($Cu + S \rightarrow CuS$)? (*b*) What weight of copper sulfide should be produced?

6. When barium peroxide is heated to 1000°C., oxygen is liberated ($2BaO_2 \rightarrow 2BaO + O_2$). What volume of oxygen can be produced from 2 kg. of barium peroxide? (See the Appendix for the density of oxygen.)

volume = mass / density → 1st equation, 1st calculate BaO_2 mass of O_2.

7. When phosphorus is burned in oxygen, phosphorus pentoxide is produced ($4P + 5O_2 \rightarrow 2P_2O_5$). (*a*) What weight of phosphorus is required for the preparation of 11.5 g. of the pentoxide? (*b*) What volume of oxygen (density = 1.429 g./liter) is required?

8. Calcium oxide reacts with water to form calcium hydroxide, $Ca(OH)_2$. What volume of water (density = 1 g./ml.) will be required to react with 8.5 g. of the oxide? What weight of the hydroxide will result? Balanced chemical equation.

$CaO + H_2O \longrightarrow Ca(OH)_2$

9. (*a*) Calculate the weight of zinc sulfate formed by the reaction of 13.2 g. of zinc with dilute sulfuric acid ($Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$). (*b*) If hydrogen gas weighs 0.0898 g./liter, what is the volume of the resulting hydrogen?

10. Hydrogen peroxide is prepared in the laboratory by the reaction $BaO_2 + H_2SO_4 \rightarrow BaSO_4 + H_2O_2$. (*a*) What weight of barium peroxide must be used in order to obtain 12 g. of the hydrogen peroxide? (*b*) What volume of a 3 per cent solution of hydrogen peroxide (density = 1 g./ml.) could be prepared from the 12 g. of H_2O_2?

NAME________________________ DESK No.________________

11. When heated, limestone decomposes according to the equation $CaCO_3 + \text{heat} \rightarrow CaO + CO_2$. (*a*) What weight must be heated to obtain 140 g. of lime, CaO? (*b*) What volume of carbon dioxide should be produced at the same time? (NOTE: See the Appendix for density of CO_2.)

12. Heating cupric oxide in a stream of hydrogen causes the reaction $CuO + H_2 \rightarrow Cu + H_2O$. (*a*) Calculate the weight of copper obtainable from 159 g. of cupric oxide. (*b*) What volume of liquid water should be obtained at the same time?

13. Phosphorus pentoxide reacts with water to form phosphoric acid, $P_2O_5 + 3H_2O \rightarrow 2H_3PO_4$. What weight of the P_2O_5 is required to make 500 g. of a 20 per cent solution of H_3PO_4?

14. Air contains 21 per cent oxygen by weight. What weight of air is required to burn 12 g. of ethane, C_2H_6, according to the equation $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$? What volume of carbon dioxide will be produced at the same time?

15. A sample of rock containing 80 per cent iron pyrites, FeS_2, was burned to furnish sulfur dioxide, SO_2. What weight of sulfur dioxide could be obtained for each kilogram of the rock burned?

$$2FeS_2 + 5O_2 \rightarrow 2FeO + 4SO_2?$$

16. Calculate (*a*) the weight of and (*b*) the volume of ammonia required to produce ammonium chloride, NH_4Cl, from 146 g. of hydrogen chloride, according to the equation

$$NH_3 + HCl \rightarrow NH_4Cl$$

17. According to the data of Problem 11, how many moles (*a*) of limestone was required; (*b*) how many moles of carbon dioxide was produced?

18. According to the data of Problem 13, (*a*) how many moles of H_3PO_4 was needed to make the solution; (*b*) how many moles of P_2O_5 should have been used in making the H_3PO_4?

19. (*a*) How many moles of FeS_2 was contained in 1 kg. of the rock in Problem 15; (*b*) how many moles of SO_2 was produced?

20. From the equation given in Problem 14, what is the number of moles of each reactant and resultant involved when 1000 g. of ethane is burned?

12. HEAT RELATIONS

The unit of heat energy in the metric system is known as the calorie and may be defined simply as that quantity of heat which is required to raise the temperature of 1 gram of water through 1°C. at 15°C.

Specific Heat

The specific heat of a substance may be defined as the quantity of heat required to raise the temperature of 1 gram of the substance through a temperature change of 1°C. Thus, the amount of heat required to raise the temperature of a given weight of the substance through a definite temperature change can be calculated if the specific heat of the substance is known. Conversely, the amount of heat released by a given weight of the substance in cooling through a definite change of temperature can be calculated. (*It should be noted that the definition of the calorie defines also the specific heat of water, viz.*, 1 cal./g./°C.

Examples

1. What quantity of heat is required to raise the temperature of 100 g. of mercury from 25 to 100°C.?

Solution. The specific heat of mercury is 0.0332 cal./g./deg. (see Table 4). Then to raise the temperature of 1 g. of mercury through the desired temperature change,

$$1 \text{ g.} \times (100 - 25) \text{ deg.} \times 0.0332 \text{ cal./g./deg.} = 2.49 \text{ cal.}$$

Or for 100 g. through the same temperature change,

$$100 \text{ g.} \times (100 - 25) \text{ deg.} \times 0.0332 \text{ cal./g./deg.} = 249 \text{ cal.}$$

The calculation just completed can be formulated: $W \times dT \times c$ = heat required, where W = the weight of the substance, dT = the temperature change, and c = the specific heat of the substance.

2. How many calories of heat must be supplied to raise the temperature of 1 kg. of iron from room temperature, 25°C., to 1000°C.?

Solution. The specific heat of iron (Table 4) is 0.1135 cal./g./deg. Then substitution of data in the formula given above shows

$$1000 \text{ g. iron} \times (1000 - 25) \text{ deg.} \times 0.1135 \text{ cal./g./deg.} = (1000 \times 975 \times 0.1135) \text{ cal.} = 127.962.5 \text{ cal.}$$

or, in significant figures, 128,000 cal.

NAME________________________________ DESK No.________________

PROBLEMS

1. What amount of heat is needed to raise the temperature of 36 g. of water from 1 to 99°C.?

2. It is desired to raise the temperature of 1 lb. of benzene from 27 to 65°C. What quantity of heat will be required?

3. Calculate the number of calories required to heat 500 g. of aluminum from 20 to 300°C.

4. When 1 gal. of water (8.33 lb.) is cooled from 95 to 15°C., how much heat is given to the surroundings?

5. Forty grams of lead, temperature 100°C., was dropped into 30 ml. of water at 20°C. The final and equilibrium temperature of the system was 23.1°. Neglecting the specific heat of the container, (*a*) how many calories of heat was transferred to the water; (*b*) calculate the specific heat of the lead.

6. The heat required to raise the temperature of 100 g. of water from 0° to 100°C. is sufficient to raise the temperature of how much iron from 0 to 200°C.?

Heat of Fusion

When a solid substance is heated to its melting temperature, it is observed that continued application of heat does not result in a further increase in temperature *until* the substance has completely melted and changed to the liquid form. The process of melting involves the absorption of a definite amount of heat energy which enables the molecules of the substance to escape or overcome the bonds which hold them in the fixed and rigid structure which is characteristic of the solid state. The quantity of heat energy necessary to change 1 gram of the solid to the liquid form at the melting point and with no change in temperature is characteristic for each substance and is defined as the *heat of fusion* of the substance.

From Table 4, the heat of fusion of benzene at its melting point is seen to be 30.1 cal./g. To melt 50 g. of benzene at 5.4°C., 50 × 30.1 = 1505 cal. must be supplied. When 50 g. of liquid benzene freezes at this same temperature, 5.4°C., 1505 cal. of heat will be given up to the surroundings.

Heat of Vaporization

When a liquid substance is vaporized at its boiling temperature and changed to the gaseous form (vapor), an absorption of heat energy occurs, which is entirely analogous to that of the fusion phenomena just described. The heat energy required to vaporize 1 gram of a liquid substance at its boiling point with no change in temperature is defined as the *heat of vaporization* of the substance.

Similarly, to evaporate 50 g. of benzene at its boiling point, 80.1°C., 50 g. × 94.3 cal./g. = 4715 cal. of heat must be supplied. Conversely, when 50 g. of benzene vapor is condensed (at the boiling point) to liquid benzene, this same amount of heat will be given up to the surroundings.

Example. How many calories of heat is required to change 100 g. of ice at −10°C. to steam at 110°C.?

Solution. Preliminary consideration shows this problem to consist of five distinct processes: (1) heating the ice from −10 to 0°C. (the melting point); (2) melting the ice at 0°C.; (3) heating the water from 0 to 100°C. (the boiling point); (4) evaporating the water at 100°C.; (5) heating the vapor (steam) from 100 to 110°C. So

(1) $$100 \text{ g. ice} \times 0 - (-10) \text{ deg.} \times 0.49 \text{ cal./g./deg.} = (100 \times 10 \times 0.49) \text{ cal.} = 490 \text{ cal.}$$

(2) $$100 \text{ g. ice} \times 79.67 \text{ cal./g.} = 7967 \text{ cal.}$$

(3) $$100 \text{ g. water} \times (100 - 0) \text{ deg.} \times 1.00 \text{ cal./g./deg.} = (100 \times 100 \times 1) \text{ cal.} = 10{,}000 \text{ cal.}$$

(4) $$100 \text{ g. water} \times 539.6 \text{ cal./g.} = 53{,}960 \text{ cal.}$$

(5) $$100 \text{ g. steam} \times (110 - 100) \text{ deg.} \times 0.48 \text{ cal./g./deg.} = (100 \times 10 \times 0.48) \text{ cal.} = 480 \text{ cal.}$$

Thus the total heat required = sum of these steps.

$$(490 + 7967 + 10{,}000 + 53{,}960 + 480) \text{ cal.} = 72{,}897 \text{ cal.}$$

TABLE 4. PROPERTIES OF VARIOUS SUBSTANCES

Substance	Melting point, °C.	Boiling point, °C.	Specific heat, cal./g./deg.			Heat of fusion, cal./g.	Heat of vaporization, cal./g.
			Solid	Liquid	Gas		
Water	0	100	0.49	1.0000	0.48	79.67	539.6
Aluminum	660	1800	0.2143	0.2484	.....	76.8	1994
Mercury	−39	357	0.0335	0.0332	.....	2.77	70.8
Benzene	5.4	80.1	0.305	0.42	0.325	30.1	94.3
Iron	1535	3000	0.1135	......	.....	47.9	1626

Law of Dulong and Petit

In 1819, Dulong and Petit announced the discovery of a general relation between the specific heat of a metallic element and its gram-atomic weight, which can be expressed

$$(\text{Specific heat of the element})(\text{the gram-atomic weight}) = \text{a constant}$$

Examples

1. Specific heat of gold = 0.0324 cal./g.; gram-atomic weight of gold = 197.2 g. Therefore

$$0.324 \text{ cal./g.} \times 197.2 \text{ g.-at. wt.} = 6.40 \text{ cal./g.-at. wt.}$$

2. Specific heat of iron = 0.112 cal./g.; gram-atomic weight of iron = 55.85 g. Therefore

$$0.112 \text{ cal./g.} \times 55.85 \text{ g.-at. wt.} = 6.26 \text{ cal./g.-at. wt.}$$

3. Specific heat of mercury = 0.0333 cal./g.; gram-atomic weight of mercury = 200.61 g. Therefore

$$0.0333 \text{ cal./g.} \times 200.61 \text{ g.-at. wt.} = 6.66 \text{ cal./g.-at. wt.}$$

The figure 6.4 cal./g.-at. wt. is usually accepted as an average value and is known as the Dulong and Petit constant. (Note that the expression represents the *gram-atomic heat* of the element.)

This relation, now known as the law of Dulong and Petit, has two direct applications:

(1) From the known atomic weight of a solid element and the Dulong and Petit constant, 6.4 cal./g.-at. wt., it is possible to obtain the approximate specific heat of that element.

(2) When the specific heat of a newly discovered element has been carefully determined, the approximate value for the atomic weight of that element can be calculated.

It must be kept in mind that 6.4 cal./g.-at. wt. (or 6.4 cal./g.-atom) is only an average value and therefore any calculations made with it in this relation will give only approximate values.

In a preceding chapter it was learned that the equivalent weight of an element can be determined with very great accuracy and that the equivalent weight of an element is related to the atomic weight as a simple factor. That is, the equivalent weight multiplied by a simple integer, *e.g.*, 1, 2, 3, etc., equals the atomic weight. The integer of this relation is identified as the valence number of the element (see Chap. 10). Thus we write

$$\text{Equivalent weight of the element} \times \text{valence number} = \text{gram-atomic weight}$$

Rearranging gives for any element

$$\text{Valence number} = \frac{\text{gram-atomic weight}}{\text{equivalent weight}}$$

Summarizing, we have in the Dulong and Petit relation a method of determining the approximate atomic weight of a newly discovered solid element, and then by analysis we can determine exactly the equivalent weight of the element. Substitution of these data into the relation shown above allows us to estimate the valence, and multiplication of the exact equivalent weight by the valence number gives the exact atomic weight. The valence number obtained by this method will be close to a small whole number, *e.g.*, 2.85 or 3.18. Such a number should be taken as the nearest whole number, and in either of these cases, 3 would be chosen.

4. A previously undiscovered metal, *M*, was found to have a specific heat of 0.0568 cal./g., and accurate analysis showed that 0.10970 g. of the element combined with oxygen to form 0.13264 g. of the oxide. Calculate (*a*) the approximate atomic weight of the element; (*b*) the equivalent weight of the element; (*c*) the exact atomic weight.

Solution. (*a*) By the law of Dulong and Petit, the approximate atomic weight of *M* is

$$\frac{6.4}{0.0568} = 112.6$$

(b) Then 0.13264 − 0.10970 = 0.02294 g. of combined oxygen. Now we can write

$$0.10970 \text{ g. } M : 0.02294 \text{ g. oxygen} = X \text{ g. } M : 8.00000 \text{ g. oxygen}$$

and

$$X = \frac{0.10970 \times 8.00000}{0.02295} = 38.26 \text{ g.}$$

which is the gram-equivalent weight of M.

The valence number of M is

$$\frac{112.6 \text{ g.}}{38.26 \text{ g.}} = 2.91$$

which is to be taken as the nearest whole number 3. Finally, the exact gram-atomic weight of M is

$$3 \times 38.26 \text{ g.} = 115.8 \text{ g.}$$

Thus the atomic weight has been determined as 115.8.

NAME_____________________________________ ____________________ DESK No.____________________

PROBLEMS

7. What quantity of heat, measured in calories, is required to melt 1 lb. of (*a*) ice; (*b*) mercury; (*c*) aluminum? (How many grams in 1 lb.?)

8. Calculate the number of calories of heat necessary to vaporize 1 gal. of water at its boiling point. (Density of water = 8.33 lb./gal.)

9. The specific gravity of mercury is 13.6. How many calories of heat is needed to vaporize 1 gal. of mercury at its boiling point?

10. If 750 g. of water is to be heated to boiling and then evaporated (at the boiling temperature) and if the initial temperature of the water is 27°C., what amount of heat will be required? (HINT: Note that two different processes are involved here.)

11. How many calories of heat is required to raise the temperature of 137 g. of mercury from −50 to 100°C.?

12. How many calories of heat is required to change 22 g. of ice at −10°C. to steam at 100°C.? (How many processes here?)

13. It has been the custom in the past to place a tub of water in the fruit cellar in the wintertime to protect the canned food from freezing. How much heat is given to the surroundings when 14 gal. of water at 10° is changed to ice at −5°C.?

14. The heat required to change 1 kg. of water at 25° to steam at 105°C. could convert what amount of benzene through the same temperature change?

15. Using the atomic weights of silver and of copper (as in the Appendix) calculate their respective specific heats.

NAME____________________ DESK No.__________

16. The specific heats of tin and zinc are, respectively, 0.0515 and 0.0825 cal./g./deg. What are their approximate atomic weights according to the Dulong and Petit law? (Compare with the atomic weights in the Appendix.)

17. The specific heat of a metallic element, *M*, is 0.032 cal./g./deg. (*a*) What is its approximate atomic weight? (*b*) If 25.8 g. of this element combined with exactly 2.000 g. of oxygen to form the oxide, what is the valence of *M*? (*c*) What is the exact atomic weight?

18. The specific heat of a solid element was found to be 0.13 cal./g./deg. (*a*) What is its approximate atomic weight? (*b*) If 1.30 g. of this element displaced exactly 0.0504 g. of hydrogen, what is the exact atomic weight of the element?

19. A 50-g. sample of a metallic element was cooled from 300 to 50°C., and the heat given up was sufficient to raise the temperature of a 25-g. sample of water from 25 to 50°C. What was the approximate atomic weight of the metal?

20. A sample of monovalent metal weighing 17.59 g. displaced exactly 0.4536 g. of hydrogen. (*a*) What is the atomic weight of the metal? (*b*) What is the approximate specific heat of the element?

13. PROPERTIES OF GASES: 1. BOYLE'S LAW

It was in 1662 that Boyle announced the discovery of the relation of the volume of a sample of gas to the pressure put upon it. This relation is stated as follows: **If the temperature is kept constant, the volume of a given sample of gas will vary inversely with the pressure put upon it.**

This important generalization is expressed in mathematical fashion by the inverse proportion,

$$V_1:V_2 = P_2:P_1 \tag{1}$$

Or by the first step in solving the inverse proportion,

$$P_1 \times V_1 = P_2 \times V_2 \tag{2}$$

Knowing the volume of a sample of gas at a definite temperature and pressure, it is possible to calculate the volume that the gas will occupy at a new pressure and at the *same* temperature.

The pressure exerted by a gas may be expressed in several ways. The practical unit is pounds per square inch. The metric unit which corresponds to this is grams (or kilograms) per square centimeter. However, it is the custom in all laboratories and research work to express the pressure on the gas volumes used in terms of the height of a mercury column which that pressure will support. The length of this mercury column, which is conveniently used as an index to, or a direct measure of, the actual pressure, is given in inches in weather reports and in centimeters in the research and chemistry-laboratory data. Finally, the standard atmospheric pressure, which is approximately 15 p.s.i. (more exactly, 14.7 p.s.i.) is referred to as 1 atm. of pressure. Thus 30 lb. pressure would be 2 atm., and 5 lb. pressure would be $\frac{1}{3}$ atm., etc. So we have as equivalent pressure expressions

$$14.7 \text{ p.s.i.} = 1 \text{ atm.} = 760 \text{ mm.} = 76 \text{ cm.} = 29.95 \text{ in.}$$

Again, be reminded that it is customary to express all units in the same system, either all metric or all English.

Examples

1. A sample of gas had a volume of 120 ml. when the pressure upon it was 760 mm. What volume would the gas occupy at 900 mm.?

Solution. Tabulation of the data shows

$P_1 = 760$ mm. $\qquad P_2 = 900$ mm.

$V_1 = 120$ ml. $\qquad V_2 = ?$

Substituting in form (2) above,

$$P_1 \times V_1 = P_2 \times V_2$$

gives

$$760 \times 120 = 900 \times V_2$$

From which

$$V_2 = \frac{760 \times 120}{900} = \frac{91{,}200}{900} = 101.3 \text{ ml.}$$

2. It is desired to pump 100 liters of air, which was measured at 700 mm. pressure, into a 10-liter container. What pressure will the air exert in its new and smaller container?

Solution. Tabulation of the data shows

$P_1 = 700$ mm. $\qquad P_2 = ?$

$V_1 = 100$ liters $\qquad V_2 = 10$ liters

Substitution in

$$P_1 \times V_1 = P_2 \times V_2$$

gives

$$700 \times 100 = P_2 \times 10$$

and

$$P_2 = \frac{700 \times 100}{10} = 7000 \text{ mm.}$$

NAME ______________________________ DESK No. _______________

PROBLEMS

1. A sample of nitrogen occupied a volume of 90 ml. at 740 mm. pressure. What would be the volume of the sample at 700 mm.?

2. Six hundred and fifty cubic feet of air was held under a pressure of 3000 p.s.i. What will be the volume if the compressed air is allowed to expand into the atmosphere? (Atmospheric pressure = 14.7 p.s.i.)

3. Forty milliliters of carbon dioxide was collected at a pressure of 700 mm. What pressure must be applied to make the gas occupy a volume of 15 ml.?

4. An automobile tire has a volume of 1 cu. ft. What volume of air at atmospheric pressure must be pumped into the tire to give a gauge pressure of 55 p.s.i.? (The tire pressure gauge reads 0, at atmospheric pressure, 14.7 p.s.i.)

5. A tank of 2.5 liters capacity was filled with helium at 40 atm. pressure. What volume would the helium occupy at atmospheric pressure? (1 atm. pressure = 760 mm. = 14.7 p.s.i.)

6. A balloon is filled with 100 liters of hydrogen at atmospheric pressure, 760 mm. If the balloon should rise to an altitude of 50,000 ft., where the pressure upon it would be 87 mm., what volume will the bag have to have in order to allow for the resultant expansion?

7. A compressed-air tank has a capacity of 12.5 cu. ft. and is filled with air at a gauge pressure of 175 p.s.i. What volume will the air occupy when it is released into the atmosphere?

8. What pressure must be applied to a given sample of gas in order to compress it to (*a*) one-third; (*b*) one-half; (*c*) two-fifths of its original volume?

9. If 3.56 liters of oxygen were measured at 600 mm. pressure, what volume would the gas occupy at 760 mm.?

10. Natural gas is usually delivered into the home for domestic use at a pressure of 4 oz./sq. in. In the mains (pipe lines) through which the gas is distributed throughout the city, the pressure is maintained at 25 p.s.i. If the meter of a householder shows a monthly consumption of 2000 cu. ft. (at the 4 oz. pressure), what was the volume of this gas in the main?

14. PROPERTIES OF GASES: 2. THE ABSOLUTE SCALE OF TEMPERATURE AND CHARLES'S LAW

Let us define a new temperature scale that differs from the centigrade by exactly 273°. This new scale, which we shall call the absolute, starts just 273° below the zero reading on the centigrade scale. That is, −273°C. = 0° abs., and 0°C. = 273° abs. Then 100°C. = 373° abs., and to convert any centigrade temperature to its equivalent on the absolute scale, it is necessary to add only 273°.

In 1787 Charles described the variation of the volume of a gas with respect to the changes in the temperature of its environment. In view of the new temperature scale which we have just defined, we may now state the law of Charles as follows: **Provided that the pressure is kept constant, the volume of a given sample of gas will vary directly with the absolute temperature.** This relation is expressed in the formula

$$V_1:T_1 = V_2:T_2$$

or

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

where T_1 is the initial and T_2 is the final temperature, expressed on the absolute scale.

Examples

1. When measured at 27°C., the volume of a sample of gas was 600 ml. What should be the volume of the gas at 100°C.? (Pressure is constant unless stated otherwise.)

Solution. Tabulation of the data shows

$V_1 = 600$ ml. $\qquad T_1 = 27 + 273 = 300°$ abs.

$V_2 = ?$ $\qquad T_2 = 100 + 273 = 373°$ abs.

Substituting in the equation above,

$$\frac{600}{300} = \frac{V_2}{373} \qquad V_2 = \frac{600 \times 373}{300} = 746 \text{ ml.}$$

2. The volume of a sample of gas, when measured at 17°C., was 4.5 cu. ft. To what temperature should the gas be heated in order that its volume shall be doubled, keeping the pressure constant all the while?

Solution. Tabulation of the data shows

$V_1 = 4.5$ cu. ft. $\qquad T_1 = 17°\text{C.} + 273 = 290°$ abs.

$V_2 = (2 \times 4.5)$ cu. ft. $= 9$ cu. ft. $\qquad T_2 = ?(°$ abs.)

Substituting in the equation as given,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad \frac{4.5}{9.0} = \frac{290}{X}$$

Solving,

$$X = \frac{9 \times 290}{4.5} = 580° \text{ abs.}$$

and

$$580 - 273 = 307°\text{C.}$$

NAME________________________________ DESK No.________________

PROBLEMS

1. Convert the following centigrade temperatures to the absolute equivalent: (*a*) 73°; (*b*) 137°; (*c*) 441°; (*d*) 7°; (*e*) 27°; (*f*) −27°; (*g*) −118°; (*h*) 53°.

2. Convert the following absolute temperatures to the centigrade equivalent: (*a*) 310°; (*b*) 260°; (*c*) 411°; (*d*) 150°; (*e*) 1243°; (*f*) 291°; (*g*) 376°; (*h*) 87°.

3. A sample of oxygen has a volume of 890 ml. at 233°C. What should be its volume at 0°C.? (NOTE: When no mention is made of the pressure in a problem of this type, it is considered as constant.)

4. One hundred forty-six milliliters of hydrogen was measured at 23°C. If the gas were heated to 100°C., what should be its new volume?

5. A gas oven has a volume of 1.8 cu. ft. After being heated to 240°C. the oven was opened and allowed to cool to room temperature, 25°C. What volume did the hot enclosed air assume on cooling?

6. A room measures 24 by 14 by 9 ft. As it is warmed from 10 to 27°C., what volume of the warmed air must escape through the openings around the doors and windows in order that the pressure inside the room shall remain constant?

7. The temperature of the oxyhydrogen blowtorch flame is about 2500°C. What volume would 1 liter of water vapor formed at that temperature occupy at 100°C.?

8. Given 300 ml. of a gas at 18°C.: what temperature must be attained in order (*a*) to double the volume of the gas; (*b*) to triple the volume?

9. During a reaction at 1200°C. carbon dioxide was evolved. When cooled to room temperature, 25°C., what fraction of the original volume did the gas occupy?

10. If 1 liter of oxygen at 0°C. (and 1 atm. pressure) weighed 1.429 g., what volume will it occupy at 100°C. and what will be the weight of 1 liter of it at the new temperature?

11. A sample of a gas measured 600 ml. at 27°C. What should be its volume at 327°C.?

12. A room has the dimensions, 50 by 90 by 20 ft. It is desired to change the air every 5 min. and to maintain the temperature at 72°F. How many cubic feet of air must be pumped per minute through the heating system (from the outside) if the outside temperature is 39°F.? (You are concerned only with the volume of the air involved as it changes from the outside temperature to the inside temperature.)

15. PROPERTIES OF GASES: 3. BOYLE'S AND CHARLES'S LAWS COMBINED: STANDARD CONDITIONS

Suppose a given quantity of gas was subjected to simultaneous changes in both temperature and pressure. What would be the proper method of calculating the resultant volume?

It is necessary to consider only the initial and final conditions and in effect make the calculations required by Boyle's and Charles's laws in one simultaneous operation. There are two methods of making this calculation.

In the first method to be described, the expressions for the laws of Boyle and Charles have been combined into one equation,

$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

where the subscript 1 signifies the initial set of conditions, and the subscript 2 denotes the final set of conditions. The data given should be properly tabulated, and then substitution into the equation above affords a direct calculation for the answer sought.

In the second method, multipliers in the form of fractions may be devised from the tabulated data so that when applied to the initial volume the effect of the changes in the conditions will be reflected in the answer obtained, according to the requirements of Boyle's and Charles's laws.

Example. What volume will 500 ml. of nitrogen at 710 mm. and 17°C. occupy under a pressure of 550 mm. and a temperature of −13°C.?

Solution. Tabulation of the data shows

$P_1 = 710$ mm. $\quad V_1 = 500$ ml. $\quad T_1 = 17 + 273 = 290°$ abs.

$P_2 = 550$ mm. $\quad V_2 = ?$ $\quad T_2 = -13 + 273 = 260°$ abs.

The change in the *temperature* will cause a *decrease* in the volume. Thus the two temperatures are to be arranged in the form of a fraction so that it will act as a *decreasing* multiplier,

$$\frac{260}{290}$$

The change in the *pressure* is such that the volume of the gas should expand, or *increase*. Therefore, the pressure values are arranged in a fractional form to make an *increasing* multiplier when applied to the original volume. Or

$$\frac{710}{550}$$

Now combining these into a single expression,

$$500 \times \tfrac{260}{290} \times \tfrac{710}{550} = 578.6 \text{ ml.}$$

The student should solve this same problem by substitution of the tabulated data into the formula given in the first method described, in order to prove to himself that the two methods are really identical. Ordinarily the student will not be asked to apply both of these methods on gas-law problems; he is urged rather to study them, to choose the one which he better understands, and then to use it alone from this time on.

Standard Conditions

When in the text and in the statement of problems, reference is made to *standard conditions*, the conditions of temperature and pressure, 0°C. and 760 mm. (1 atm.), are implied. These are the conditions that have been chosen as the *standard* reference conditions under which gas densities are compared and listed in tables. Standard conditions are often indicated by the abbreviation S.T.P.

NAME________________________________ ______________________ DESK No.____________________

PROBLEMS

1. If 236 ml. of gas is measured at 720 mm. pressure at 27°C., what will be the volume of the gas at standard conditions?

2. At 600 mm. pressure and 10°C. a sample of nitrogen had a volume of 124 ml. What should the volume be at 732 mm. and 37°C.?

3. What pressure must be applied at 137°C. to make a gas occupy a volume 360 ml. if its volume at 735 mm. and 23°C. is 420 ml.?

4. An automobile tire whose volume was 0.8 cu. ft. was filled with air in the early morning. The tire pressure gauge showed a pressure of 36 lb., and the temperature was 22°C. Later in the day the temperature of the tire, running on the highway, reached 70°C., and the tire itself had expanded to a volume of 0.82 cu. ft. What was the pressure in the tire? (Remember that the gauge registers 0 pressure at atmospheric pressure.)

5. A sample of nitrous oxide occupied a volume of 120 ml. at 17°C. and 720 mm. The pressure was increased to 1520 mm. To what temperature should the gas be heated so that it may have a volume of 100 ml. at the new pressure?

6. What volume of oxygen, measured at 5°C. and 0.9 atm., must be pumped into a cylinder of 1.5 cu. ft. capacity to furnish a pressure of 3000 p.s.i. at 28°C.?

7. Calculate the volume that a sample of nitrogen would occupy at 640 mm. and 33°C. if it has a volume of 320 ml. at 736 mm. and 25°C.

8. At 745 mm. pressure and 40°F. a sample of gas had a volume of 450 ml. What should be the volume of the gas at 550 mm. and −20°F.?

9. What would be the volume, at 718 mm. and 29°C., of the oxygen liberated by the decomposition of 140 g. of mercuric oxide? (The density of oxygen at standard temperature and pressure is 1.429 g./liter.)

10. What volume of hydrogen, measured at standard conditions of temperature, should be displaced by treating 165 g. of zinc with excess dilute sulfuric acid? What should be the volume of this hydrogen at 31°C. and 580 mm.? (The density of hydrogen at standard temperature and pressure is 0.0898 g./liter.)

16. PROPERTIES OF GASES: 4. DALTON'S LAW

If two gases occupy the same container (*i.e.*, the same volume), *each gas acts as if it alone occupied that volume* and exerts its own individual pressure: The *total pressure* exerted by the mixture is equal to *the sum of the individual pressures* exerted by the various gases. For example, in a mixture of oxygen and nitrogen contained in a given vessel, the oxygen exerts its own pressure, as also does the nitrogen, and the total pressure in the vessel is the sum of the pressures exerted by the individual gases. Actually, atmospheric pressure is equal to the sum of the individual pressures exerted by all of the various gases that compose the atmosphere.

Dalton was the first to recognize these relations. They are summarized in the expression

$$P_{\text{total}} = P_1 + P_2 + P_3 + \cdots$$

Examples

1. If 1 liter of oxygen, 1 liter of nitrogen, and 2 liters of carbon dioxide, each measured at standard conditions of temperature and pressure, were pumped into a single container whose volume was 1 liter, what was the total pressure exerted by the three gases, the temperature remaining at 0°C.?

Solution. Calculate the pressure exerted by each gas as if it alone occupies the container. $P_{\text{total}} = P_1 + P_2 + P_3$.

For the oxygen, the final conditions of volume and temperature are identical with the initial conditions. Therefore the pressure due to the oxygen will be 1 atm., or 760 mm.

The same applies to the nitrogen. Therefore the pressure due to the nitrogen will be 1 atm., or 760 mm.

Pumping the carbon dioxide into the 1-liter container at the same temperature means that the final volume it will occupy is only half of the original volume. According to Boyle's law this means that its pressure will be doubled. (Make the calculation.)

$$P_{\text{total}} = P_{\text{oxygen}} + P_{\text{nitrogen}} + P_{\text{carbon dioxide}}$$

$$P_{\text{total}} = 1 \text{ atm.} + 1 \text{ atm.} + 2 \text{ atm.} = 4 \text{ atm.}$$

or $$P_{\text{total}} = 760 \text{ mm.} + 760 \text{ mm.} + 1520 \text{ mm.} = 3040 \text{ mm.}$$

The most useful laboratory application of Dalton's law has to do with the collection of a sample of gas over water. Two gases will occupy the measured volume, the one desired, along with water vapor. Since water vapor is a gas, and as such exerts a pressure similar to oxygen or any other gas, a suitable correction must be made for its presence if an accurate measurement of collected gas is desired. Fortunately, the pressure due to water vapor in the presence of water has a definite value for each temperature. These values have been carefully measured and recorded in tables which the student may use (see the Appendix). The following example illustrates how the correction for the water vapor pressure (often called aqueous tension) is made.

2. A sample of oxygen collected over water occupied a volume of 325 ml. at 25°C. and a measured pressure of 720 mm. (*a*) What is the correct pressure of the oxygen? (*b*) What should be the volume of the oxygen (dry) at standard temperature and pressure?

Solution. (*a*) Reference to the Table of Vapor Pressures (see the Appendix) shows that the pressure due to water vapor at 25°C. is 23.6 mm.

Now $$P_{\text{measured}} = P_{\text{oxygen}} + P_{\text{water vapor}}$$

or $$720 \text{ mm.} = P_{\text{oxygen}} + 23.6 \text{ mm.}$$

and $$P_{\text{oxygen}} = (720 - 23.6) \text{ mm.} = 696.4 \text{ mm.}$$

(*b*) Tabulation of the data now shows

$V_1 = 325$ ml.	$T_1 = 25°C. = 298°$ abs.	$P_1 = 696.4$ mm.
$V_2 = ?$	$T_2 = 0°C. = 273°$ abs.	$P_2 = 760.0$ mm.

[Note that (1) the temperature decrease, 298 to 273, causes the volume to decrease, and that (2) the pressure increase also causes the volume to decrease.]

Therefore,
$$V_2 = 325 \times \frac{273}{298} \times \frac{696.4}{760.0} = 272.8 \text{ ml.}$$

3. A 5.000-g. sample of potassium chlorate was heated to decomposition in preparing oxygen which was collected over water (pneumatic trough) at 18°C. and an atmospheric pressure of 742 mm. What volume should the wet gas measure?

Solution. Reference to the balanced equation shows

$$2KClO_3 = 2KCl + 3O_2$$

and	2(122.6) g.	3(32) g.
So	5.000 g.	X g.

Then
$$X = \frac{5 \times 3 \times 32}{2 \times 122.6} = 1.549 \text{ g. oxygen which will be released}$$

The density of oxygen (see the Appendix) is 1.429 g./liter standard temperature and pressure; so

$$\frac{1.549 \text{ g.}}{1.429 \text{ g./liter}} = 1.084 \text{ liter} = 1084 \text{ ml. at S.T.P.}$$

Now note that the 742 mm. pressure was the measured total pressure of the volume of oxygen plus water vapor. The vapor pressure of water at 18°C. is 15.5 mm. (see the Appendix). Subtraction of this figure from the total pressure of the two gases gives the pressure due to the oxygen alone.

$$742 - 15.5 = 726.5 \text{ mm.}$$

Tabulation of the data now shows

$V_1 = 1084$ ml.	$T_1 = 0°C. = 273°C.$	$P_1 = 760$ mm.
$V_2 = ?$	$T_2 = 18°C. = 291°$ abs.	$P_2 = 726.5$ mm.

Therefore,
$$V_2 = 1084 \times \frac{291}{273} \times \frac{760}{726.5} = 1208 \text{ ml.}$$

NAME________________________________ DESK NO.____________

PROBLEMS

1. If 500 ml. of nitrogen and 500 ml. of oxygen were measured separately at 760 mm. pressure and then were forced into a 1-liter flask, what was the total pressure in the flask? (The temperature remains constant.)

2. If 500 ml. of nitrogen at 1 atm. pressure and 500 ml. of oxygen at 2 atm. pressure were forced into a 1-liter flask, what was the resulting pressure in the flask?

3. Hydrogen was collected over water at 17°C. and 710 mm. pressure. What was the true pressure of the hydrogen?

4. If 1 liter of nitrogen at 500 mm. pressure was forced into a 500-ml. container which already contained oxygen at 1000 mm. pressure, what was the total pressure exerted by the mixture?

5. If 265 ml. of oxygen was collected over water at 31°C. and 733 mm., what volume would the oxygen occupy, dry, at 760 mm. and 0°C.? (Correct the measured pressure for the presence of water vapor, and calculate as in the previous chapter.)

6. A sample of nitrogen had a volume of 143 ml. when measured over water at 21°C. and 726 mm. What should be the volume of the dry gas at standard temperature and pressure?

7. A flask of 1-liter volume contained three gases at a total pressure of 1500 mm. The ratios of the partial pressures of the gases were 1:1.5:0.5. What volume would each gas occupy separately at a pressure of 760 mm.?

8. When collected over water at 18°C. and 745.4 mm., a sample of a gas had a volume of 333 ml. What volume should the dry gas occupy at standard temperature and pressure?

9. If 300 ml. of hydrogen at 720 mm., 800 ml. of nitrogen at 600 mm., and 450 ml. of oxygen at 750 mm. were forced into a 1-liter container, what was the total pressure in the container? (Consider the temperature as constant throughout.)

10. A 2.5-g. sample of magnesium was treated with excess dilute sulfuric acid, and the resulting hydrogen was collected over water at 718 mm. and 16°C. What was the measured volume of the hydrogen collected? (NOTE: The measured pressure = pressure of hydrogen + pressure of water vapor.)

NAME____ DESK No.________________

11. Suppose dry air is composed of nitrogen 78.03 per cent, oxygen 20.99 per cent, argon 0.94 per cent, and carbon dioxide 0.04 per cent by volume. (*a*) What is the partial pressure of each when the barometric pressure is 740 mm.? (*b*) What is the composition of air expressed in per cent by weight?

12. A sample of zinc weighing 0.6 g. was treated with dilute sulfuric acid in order to prepare hydrogen, $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$. What volume should the released hydrogen occupy if collected over water at 26°C. and 754.3 mm. atmospheric pressure?

13. When 0.2837 g. of a certain metal reacted by displacement with hydrochloric acid, 298 ml. of hydrogen was collected over water at a measured temperature of 27°C. and a barometric pressure of 757.7 mm. What is the equivalent weight of the metal?

14. A volume of 312 ml. of hydrogen was collected over water at 22°C. and 729.8 mm. barometric pressure when 0.2158 g. of a metal reacted by displacement with dilute sulfuric acid. (*a*) Calculate the equivalent weight of the metal. (*b*) If the valence of this metal is 3, what is its atomic weight? (*c*) What is its specific heat?

17. THE MOLAR VOLUME OF GASES

Under standard conditions of temperature and pressure (0°C., and 760 mm.) 1 liter of oxygen weighs 1.429 g. The gram-molecular weight of oxygen is 32 g. By dividing the weight of 1 liter of oxygen into the gram-molecular weight of the substance, we find the volume occupied by 32.000 g. of oxygen at standard temperature and pressure.

$$\frac{32.000 \text{ g.}}{1.429 \text{ g./liter}} = 22.4 \text{ liters}$$

The volume of the gram-molecular weight of any gas can be calculated in a similar manner if the density of the gas (weight of 1 liter in grams) and the molecular weight are known.

Exercise. The densities of several gases measured at standard temperature and pressure are given below. As in the example above, determine the volumes occupied by the gram-molecular weights of the various gases.

Nitrogen, N_2	1.250 g./liter	Carbon dioxide, CO_2	1.966 g./liter
Chlorine, Cl_2	3.214	Sulfur dioxide, SO_2	2.927
Hydrogen, H_2	0.0898	Hydrogen chloride, HCl	1.640

That the gram-molecular weight of any gas will occupy a definite and constant volume at standard conditions, namely, 22.4 liters, is an important principle in making chemical calculations and has several applications: (1) the determination of the approximate molecular weight of a gas; (2) the calculation from the balanced equation of the volumes of gases involved in reactions; (3) the calculation of the density of a gas at standard conditions when its molecular weight is known.

Determination of the Approximate Molecular Weight of a Gas

A sample of a gas may be weighed and its volume measured under known conditions of temperature and pressure. The measured volume is corrected to standard conditions by applying Boyle's and Charles's laws, and then by using simple proportion, the weight of 22.4 liters of the gas may be calculated.

Example. A sample of gas weighed 0.2683 g. and occupied a volume of 225 ml. at 18°C. and 720 mm. pressure. What is its approximate molecular weight?

Solution. Tabulation of data shows

$V_1 = 225$ ml. $\quad P_1 = 720$ mm. $\quad T_1 = 18°\text{C.} + 273 = 291°$ abs.

$V_2 = ?$ $\quad P_2 = 760$ mm. $\quad T_2 = 0°\text{C.} + 273 = 273°$ abs.

Then

$$V_2 = 225 \times \frac{720}{760} \times \frac{273}{291} = 200 \text{ ml.} = \text{the volume occupied by the sample of gas at standard conditions}$$

Now if 200 ml. of the gas weighs 0.2683 g., 22,400 ml. will weigh a corresponding amount.

$$200:0.2683 = 22{,}400:X$$

or

$$X = \frac{0.2683 \times 22{,}400}{200} = 30.05 \text{ g.} = \text{the approximate molecular weight of the gas}$$

Volumes of Gases Involved in Reactions

Since the gram-molecular weight of a gas is now known to occupy 22.4 liters under standard conditions, the calculation of the volumes of gases taking part in, or being evolved during, a chemical reaction can be simplified by substituting the figure for the gram-molecular volume of the gas (22.4 liters) for its molecular weight in the calculation from the balanced equation.

Example. What volume of hydrogen, at standard temperature and pressure, is produced by the action of 10 g. of zinc on dilute hydrochloric acid?

Solution. The balanced equation gives the relation

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

or

$$65.37 \text{ g.} \qquad\qquad 2.016 \text{ g. } (22{,}400 \text{ ml.})$$

Thus 65.37 g. of zinc will release 22,400 ml. of hydrogen (standard temperature and pressure); therefore 10 g. of zinc will release a corresponding volume X

or

$$65.37:22{,}400 = 10:X$$

and

$$X = \frac{22{,}400 \times 10}{65.37} = 3426 \text{ ml.}$$

$$= 3.426 \text{ liters of hydrogen released by the action of 10 g. zinc on dilute hydrochloric acid}$$

Note that in this calculation it is assumed that the gas is released under standard conditions of temperature and pressure. In the actual experiment the gas will immediately assume the conditions of its environment, and its volume will change accordingly. The usual application of Boyle's and Charles's laws then may be made in order to determine the volume it should occupy under laboratory conditions.

Calculation of the Density of a Gas at Standard Temperature and Pressure

This is the reverse of the calculation required in the first application described. It is obvious that if it is known that the gram-molecular weight of a gas occupies 22.4 liters under standard conditions, then the weight of 1 liter of the gas can be estimated by dividing the molecular weight by 22.4.

Example. What is the density of hydrogen sulfide gas, H_2S, under standard conditions?

Solution. Calculate the molecular weight of the gas from the formula and the atomic weights.

$$(2 \times 1.008) + 32.06 = 34.076 = \text{molecular weight}$$

Since 22.4 liters of the gas weighs 34.076 g., then 1 liter will weigh

$$\frac{34.076}{22.4} = 1.521 \text{ g.}$$

NAME______________________________ DESK No.______________

PROBLEMS

1. Calculate the densities at standard temperature and pressure (weight per liter) of each of the following gases: (*a*) nitric oxide, NO; (*b*) carbonyl chloride (phosgene), $COCl_2$; (*c*) ammonia, NH_3.

2. What volume of oxygen at standard conditions will be released by the heat decomposition of 1 kg. of potassium chlorate, $KClO_3$? What volume will this oxygen occupy at 720 mm. pressure and 22°C.?

3. What volume of hydrogen (standard conditions) should be released by the action of 17.5 g. of magnesium on dilute sulfuric acid?

4. When calcium carbonate is heated to 1000°C., it decomposes with the formation of carbon dioxide and calcium oxide. What volume of carbon dioxide should be released by the heat decomposition of 1 lb. of calcium carbonate (standard temperature and pressure)? What was the volume of the gas at 1000°C. and 700 mm. pressure?

5. What volume does 7.06 g. of oxygen occupy at standard conditions? What would be the volume of this oxygen at 710 mm. and 35°C.?

6. It is desired to prepare a solution which contains hydrogen chloride dissolved in water at the ratio of 400 volumes of the gas to 1 volume of water. What weight of salt, NaCl, is required to furnish by the action of concentrated sulfuric acid the hydrogen chloride to saturate 600 ml. of water at the above rate?

7. Given the formula for hydrogen bromide gas, HBr, calculate the weight of 1 liter of the gas at standard conditions. What would be the weight of 1 liter of the gas at 25°C. and 750 mm.?

8. If 333 cc. of a gas was collected over water at 18°C. and 745.4 mm. atmospheric pressure and the sample of the gas weighed 0.805 g., what was the molecular weight of the gas?

9. The equation for the electrolysis of brine solution (Nelson and Vorce cells) is $2NaCl + 2H_2O = H_2 + Cl_2 + 2NaOH$. (*a*) What volume of hydrogen is liberated at standard conditions with the consumption of 1 kg. of sodium chloride? (*b*) What weight of sodium chloride is required for the production of enough hydrogen to fill a cylinder of $2\frac{1}{2}$ cu. ft. capacity at 25°C. and 3000 lb. pressure? (Atmospheric pressure = 14.7 lb.)

10. A sample of gas weighed 1.066 g., and its volume was 264 cc. when measured at 100°C. and 735 mm. What is the molecular weight of the gas?

NAME____ DESK No.________________

11. What weight of barium peroxide is required for the liberation of 100 liters of oxygen to be measured at 15°C. and 710 mm. pressure?

12. What is the molecular weight of a gas whose density is just twice that of oxygen?

13. When 500 cc. of a certain gas was measured at standard conditions, it was found to be 0.625 g. (*a*) What was the molecular weight of the gas? (*b*) Analysis showed the gas to be 85.71 per cent carbon and 14.29 per cent hydrogen. What is the correct formula for the gas?

14. The molecular weight of a certain gas is 70.914. What volume will 1 g. of this gas occupy at 600 mm. and 17°C.?

15. Acetylene gas, C_2H_2, is formed when water is dropped on calcium carbide, $CaC_2 + H_2O \rightarrow CaO + C_2H_2$. What weight of calcium carbide is needed to produce 1000 cu. ft. of acetylene, measured at 21°C. and 740 mm.?

16. What volume will 11.000 g. of carbon dioxide, CO_2, occupy in liters at 77°C. and 570 mm. pressure?

17. (*a*) Calculate the density of chlorine gas, Cl_2, at standard temperature and pressure. (*b*) What is its density at 37°C. and 600 mm. pressure?

18. Calculate the density of hydrogen sulfide gas, H_2S, at 127°C. and 950 mm. pressure.

19. A 617-ml. sample of a hydrocarbon gas weighing 0.823 g., at 0.9 atm. and 27°C., was found to be 79.878 per cent carbon and 20.122 per cent hydrogen. (*a*) What is the molecular weight of the gas; (*b*) what is the formula?

20. What volume of chlorine gas, measured at 18°C. and 720 mm. pressure, is required to displace the bromine from 100 g. of sodium bromide? ($Cl_2 + 2NaBr \rightarrow Br_2 + 2NaCl$.)

18. SOLUTIONS—CONCENTRATION EXPRESSIONS AND DEFINITIONS

The *concentration* of a solution, or the relation of the *solute* to the *solvent*, can be expressed in a number of ways. Each method of expression has some certain advantage in that emphasis may be given to a certain relation. The expression chosen in a given instance may be one of convenience, either in preparing the solution or in using the solution after it has been prepared.

Most tables of solubility in handbooks and manuals give solubility data in the form

(a) *Grams* of *solute* in 100 g. of *solvent*

Such data can be transformed quickly and simply by a simple calculation into

(b) *Per cent* of solute (by weight)

And if the formula of the solute is known (and therefore the molecular weight), an expression can be obtained directly in terms of

(c) Number of *moles* of *solute* per 1000 g. of *solvent*

This last is defined as the *molality* of the solution. (Consult your text for a formal definition of the ***molal*** solution.)

Another very important and widely used expression of concentration is in terms of

(d) Number of *moles* of *solute* per *liter of solution*

This is known as the *molarity* of the solution. Note carefully the difference in the expressions of (c) and (d). The molarity of a solution is not interconvertible with the preceding expressions *unless* additional information is available, namely, the density of the solution.

Example. When 12 g. of sucrose, $C_{12}H_{22}O_{11}$, was dissolved in 80 g. of water, the resulting solution had a density of 1.0508 g./ml. Express the concentration of this solution in terms of (a) grams of sucrose per 100 g. of water; (b) percentage sucrose by weight; (c) molality; (d) molarity.

Solution.

(a) $$12 \text{ g. sucrose} : 80 \text{ g. water} = X \text{ g. sucrose} : 100 \text{ g. water}$$

or $$X = 15 \text{ g. sucrose per } 100 \text{ g. water}$$

(b) $$12 \text{ g. sucrose} + 80 \text{ g. water} = 92 \text{ g. solution}$$

The fraction of the whole solution due to sucrose alone is

$$\tfrac{12}{92} = 0.1304 = 13.04\% \text{ sucrose}$$

(c) $$12 \text{ g. sucrose} : 80 \text{ g. water} = X \text{ g. sucrose} : 1000 \text{ water}$$

or $$X = 150 \text{ g. sucrose per } 1000 \text{ g. water}$$

Now the gram-molecular weight of sucrose, $C_{12}H_{22}O_{11}$, is 342. Therefore the number of moles of sucrose is

$$\frac{150 \text{ g.}}{342 \text{ g./mole}} = 0.438 \text{ moles sucrose per } 1000 \text{ g. water}$$

or the concentration of this solution is 0.438 molal, or 0.438 *m*.

(*d*) One liter (1000 ml.) of this solution will weigh

$$1000 \text{ ml.} \times 1.0508 \text{ g./ml.} = 1050.8 \text{ g.}$$

And the amount of sucrose dissolved in this volume of the solution will be

$$13.04\% \text{ of the weight} = (0.1304 \times 1050.8) \text{ g.} = 137.02 \text{ g. sucrose}$$

Now the number of moles of dissolved sucrose is

$$\frac{137.02 \text{ g.}}{342 \text{ g./mole}} = 0.400 \text{ moles sucrose per liter of solution}$$

and the solution is said to be 0.400 molar, or 0.400 *M*.

(NOTE: Calculations involving equivalent weights and normal solutions are reserved for a later chapter.)

NAME________________________________ DESK NO.____________

PROBLEMS

1. A solution was made by dissolving 12 g. of glucose, $C_6H_{12}O_6$, in 60 g. of water. Express the concentration of the solution in (*a*) grams of solute per 100 g. of solvent; (*b*) per cent of solute by weight.

2. From the data in the preceding problem, express the concentration of the solution in terms of molal concentration.

3. When 64 g. of magnesium chloride, $MgCl_2$, was dissolved in 1 liter of solution, the density was found to be 1.0517 g./ml. Express the concentration of the solution as (*a*) molar; (*b*) grams of solute per 100 g. of solvent; (*c*) molal; (*d*) per cent of solute by weight.

4. It is desired to prepare a 10 per cent solution of potassium iodide in water. (*a*) What are the respective weights of solute and solvent to be used in preparing a total of 600 g. of solution? (*b*) Express the figures in molal concentrations. (*c*) What is the weight of solute per 100 g. of solvent?

5. (*a*) If 25 g. of calcium chloride, $CaCl_2$, was dissolved in 150 ml. of solution, what was the molar concentration? (*b*) If 25 g. of calcium chloride, $CaCl_2$, was dissolved in 150 g. of water, what was the molal concentration? (*c*) Are these two solutions identical? Explain your answer.

6. In preparing a 0.7 molal solution of sodium sulfate, Na_2SO_4, (*a*) what weight of the salt should be taken per 100 g. of solvent? (*b*) Express the concentration in terms of per cent of solute in the solution.

7. A solution of acetic acid, $HC_2H_3O_2$, has a density of 1.0151 g./ml., and is 2*M* in concentration. Express the concentration in (*a*) per cent of solute; (*b*) grams of solute per 100 g. of solvent; (*c*) molal terms.

8. Concentrated nitric acid has a density of 1.4176 g./ml. and is 71 per cent HNO_3 by weight. What is the concentration of this acid in (*a*) grams of HNO_3 per 100 g. of water; (*b*) moles of HNO_3 per 1000 g. of water (molal); (*c*) moles of HNO_3 per liter of solution (molar).

9. The density of liquid bromine is 3.119 g./ml. When 5 ml. of bromine is added to 300 g. of water (density = 1 g./ml.), what is the concentration of the solution in terms of per cent of solute?

10. The acetic acid solution in Problem 7 is to be diluted with water at the rate of 500 ml. of the solution: 1000 ml. of water (density of water = 1 g./ml.). Calculate (*a*) the molality and (*b*) the percentage composition of the dilute solution.

19. PROPERTIES OF SOLUTIONS

The physical properties of a solvent are directly affected by the introduction of solute molecules into it. The variations in the properties depend upon the number of molecules of solute dissolved in a given quantity of solvent. That is, the change in a property is directly proportional to the concentration of the solute molecules in the solution.

Freezing- and Boiling-point Phenomena

The following table shows the effect upon the freezing and boiling points of some representative solvents when 1 *mole of solute* is dissolved in 1000 g. *of the solvent.* (NOTE: This is a 1 *molal* solution.) The effects or variations are listed in the columns headed K_f and K_b, and it should be noted that the variation is different and characteristic for each solvent.

TABLE 5

Solvent	Freezing point of solvent, °C.	Freezing point of 1*m* solution, °C.	K_f, °C.	Boiling point of solvent, °C.	Boiling point of 1*m* solution, °C.	K_b, °C.
Benzene	5.4	0.28	5.12	80.1	82.67	2.57
Camphor	178.4	140.7	37.7	209.0	215.1	6.1
Carbon tetrachloride	−22.6	−25.58	−2.98	76.8	81.68	4.88
Naphthalene	80.2	73.3	6.9	218.0	223.8	5.8
Nitrobenzene	5.7	−2.4	8.1	210.9	216.7	5.27
Water	0.00	−1.86	1.86	100.0	100.52	0.52

It is important to remember that the values recorded above for K_f and K_b are for 1000 g. of solvent. This must be noted in all calculations which are based on these phenomena.

The effect caused by dissolving 2 moles of solute in 1000 g. of the solvent would be twice that caused by dissolving 1 mole. Similarly the effect of dissolving a fractional part of a mole of solute would correspond to the fraction used.

When the molecular weight of the solute, the respective weights of the solute and solvent, and the value of K_f for the solvent are known, it is possible to calculate the freezing point of the given solution. Similarly when K_b for the solvent is known, the boiling point of the given solution can be calculated.

Examples

1. The formula for acetone is C_2H_6O. If 10 g. of acetone is dissolved in 120 g. of water, what is the freezing point of the solution?

Solution. Summarize the data:

(1) K_f for water = 1.86° per mole of solute per 1000 g. of solvent.
(2) The given solution = 10 g. of acetone in 120 g. of water.
(3) The molecular weight of acetone C_3H_6O = 58.

Since 58 g. of acetone (1 mole) in 1000 g. of water lowers the freezing point 1.86°, then 10 g. of acetone in 120 g. of water would lower the freezing point $X°$.

$$\tfrac{58}{1000}:1.86 = \tfrac{10}{120}:X$$

or

$$0.058:1.86 = .0833:X$$

Thus $X = \dfrac{0.0833 \times 1.86}{0.058} = 2.67$ deg.

= the freezing-point lowering by a concentration of 10 g. acetone in 120 g. water

Therefore the freezing point of the solution would be (0° − 2.67°) = −2.67°C. (NOTE: Dividing the weight of the solute by the weight of the solvent in each of the left-hand ratios of the above proportion gives the quotient, grams of solute per gram of solvent. Thus the two ratios are identical.)

It is possible to calculate the *molecular weight of an unknown solute* from the freezing point of its solution in a solvent whose K_f value is known.

2. Five grams of benzoic acid was dissolved in 80 g. of naphthalene, and the resulting solution froze at 76.7°C. What is the molecular weight of benzoic acid?

Solution. Summarize the data:

(1) K_f for naphthalene = 6.9°.
(2) The freezing point of naphthalene = 80.2°C.
(3) The lowering of the freezing point = (80.2 − 76.7) = 3.5°C.
(4) The concentration of the solution = 5 g. of benzoic acid in 80 g. of naphthalene.

So

$$\frac{M}{1000} : 6.9 = \frac{5}{80} : 3.5$$

and

$$\frac{3.5M}{1000} = \frac{6.9 \times 5}{80}$$

Or

$$0.0035M = 0.43125$$

and

$$M = 123+ = \text{molecular weight of benzoic acid}$$

NAME____ DESK No.____________________

PROBLEMS

1. Calculate the boiling point of the solution described in Example 1.

2. Calculate the boiling point of the solution described in Example 2.

3. A solution was made by dissolving 12 g. of glycerin, $C_3H_8O_3$, in 80 g. of water. At what temperature should the solution freeze? At what temperature should the solution boil?

4. A 5 per cent solution of sugar, $C_{12}H_{22}O_{11}$, should freeze at what temperature centigrade and at what temperature Fahrenheit?

5. The principal component of Prestone and Zerex, automobile-radiator antifreeze liquids, is ethylene glycol, $C_2H_6O_2$. What ratio by weight of glycol to water should be used to make a solution which will freeze at $-10°C.$?

6. If 5 g. of naphthalene, $C_{10}H_8$, is dissolved in 50 g. of nitrobenzene, at what temperature should the solution (*a*) boil; (*b*) freeze?

7. The large value of K_f for camphor is often taken advantage of in determining the approximate molecular weights of substances whose identity is unknown. Estimate the molecular weight of a substance 0.5 g. of which was dissolved in 5 g. of camphor, the resulting solution having a freezing point of 157.3°C.

8. When 1 g. of phthalic acid was dissolved in 12.5 g. of naphthalene, the solution was found to have a freezing point of 77.1°C. Estimate the molecular weight of the phthalic acid.

9. Two solutions were prepared. In one, 6 g. of benzene, C_6H_6, was dissolved in 75 g. of nitrobenzene, and in the other, 6 g. of nitrobenzene, $C_6H_5NO_2$, was dissolved in 75 g. of benzene. What is the freezing point of each?

10. Compare the respective weights of methyl alcohol, CH_4O, and ethyl alcohol, C_2H_6O, which must be dissolved in 500 g. of water to make a solution which will freeze at −5°C. Given the densities, methyl alcohol 0.793 g./ml. and ethyl alcohol 0.789 g./ml., compare the volumes required. What will be the freezing points of the solutions produced by dissolving 1 qt. of each of these two alcohols separately in 1 gal. of water?

11. (*a*) If 10 ml. of acetone, C_3H_6O (density = 0.792 g./ml.), was dissolved in 120 ml. of benzene (density = 0.894 g./ml.), at what temperature should the solution freeze? (*b*) When 5 g. of an unknown substance was dissolved in 120 g. of benzene, the freezing temperature of the resulting solution was found to be 3.05°C. What was the molecular weight of the solute?

12. One gallon each of methyl alcohol, CH_4O; ethyl alcohol, C_2H_6O; and ethylene glycol, $C_2H_6O_2$, was dissolved separately in 3-gal. portions of water. The densities of the solutes are methyl alcohol 0.793 g./ml., ethyl alcohol 0.789 g./ml., and ethylene glycol 1.115 g./ml. What will be the freezing temperatures of these solutions? In your estimation, which is the most efficient antifreeze per volume of solute used?

20. EQUIVALENT WEIGHTS AND NORMAL SOLUTIONS

The equivalent weight of an *element* is that amount of it which *combines* with or *displaces* 1.008 grams of hydrogen or its *equivalent* (see the discussion of equivalent weights in Chap. 10).

The equivalent weight of a *compound* is that amount of it which *contains* 1.008 grams of combined and replaceable hydrogen or its *equivalent.*

Since the valence of hydrogen is 1, this means that the atomic weight of a univalent element capable of replacing hydrogen, or in turn replaceable by hydrogen, is the equivalent weight of that element. The molecule of the compound hydrogen chloride, HCl, contains one atom of replaceable hydrogen. The gram-molecular weight of this compound is therefore its equivalent weight. Similarly, the molecule of sodium chloride, NaCl, contains one atom of sodium, a univalent element, which has substituted for the hydrogen atom of the hydrogen chloride molecule. Therefore, the gram-molecular weight of the sodium chloride contains the equivalent of 1.008 g. of hydrogen and so is the equivalent weight of the compound.

The molecule of sulfuric acid, H_2SO_4, contains two atoms of replaceable hydrogen (2.016 g. in the gram-molecular weight). The gram-molecular weight of this compound must be divided by two, in order to find the quantity of the substance which contains exactly 1.008 g. of replaceable hydrogen. The molecule of sodium sulfate, Na_2SO_4, does not contain any replaceable hydrogen, but it does contain two sodium atoms, each of which is equivalent to one hydrogen atom, and therefore the equivalent weight of this substance is obtained by dividing the gram-molecular weight by two.

As a last example, consider the molecule of aluminum chloride, $AlCl_3$. The valence of the aluminum atom is 3, and in the formation of this compound a single aluminum atom has replaced the hydrogen atoms of three molecules of hydrogen chloride. To find the equivalent weight of aluminum chloride, it is necessary to divide the gram-molecular weight by three.

NAME____ DESK No.______________

PROBLEMS

1. Calculate the equivalent weight of each of the following compounds: sodium chloride, $NaCl$; potassium sulfate, K_2SO_4; calcium chloride, $CaCl_2$; phosphoric acid, H_3PO_4; and potassium chlorate, $KClO_3$.

2. Calculate the equivalent weight of each of the following compounds: zinc sulfate, $ZnSO_4$; copper nitrate, $Cu(NO_3)_2$; silver nitrate, $AgNO_3$; and aluminum sulfate, $Al_2(SO_4)_3$.

3. How many equivalent weights does 100 g. of each of the following substances contain: sodium sulfate, Na_2SO_4; barium chloride, $BaCl_2$; oxalic acid, $H_2C_2O_4$; ferric nitrate, $Fe(NO_3)_3$; and ammonium chloride, NH_4Cl?

4. Calculate the equivalent weight of each of the following compounds: potassium oxalate, $K_2C_2O_4$; sodium phosphate, Na_3PO_4; sodium sulfite, Na_2SO_3; magnesium sulfate heptahydrate, $MgSO_4 \cdot 7H_2O$; and copper sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$.

5. What fraction of an equivalent weight is contained in 10.0 g. of each of the following compounds: sodium hydroxide, NaOH; potassium permanganate, $KMnO_4$; barium hydroxide, $Ba(OH)_2$; ferric phosphate, $FePO_4$; and potassium persulfate, $K_2S_2O_8$?

A solution which contains a solute (acid, base, or salt) dissolved at the rate of 1 *gram-equivalent weight per liter of solution* is said to be 1 *normal* (1*N*). It should be carefully noted that the *normal* solution is defined in terms of the total *volume* of the solution and *not* in terms of a quantity of the solvent. (Note the definition of the *molar* solution in Chap. 18.)

In preparing such a solution, only enough of the solvent is used to first dissolve the required amount of solute, and then this solution is diluted to exactly 1 liter of solution. A solution containing one-half of an equivalent weight of a given solute in 1 liter of solution is said to be one-half normal, or simply 0.5*N*. Similarly, if 2 gram-equivalent weights of a given solute are contained in 1 liter of solution, the concentration is said to be two normal, or 2*N*. The *normality* of a solution, then, is simply the number of gram-equivalent weights of the solute which are dissolved in 1 liter of the solution.

NAME________________________________ DESK No.____________

PROBLEMS

6. How is 1.5 liters of a 1*N* solution of sodium chloride prepared?

7. What weight of sodium hydroxide is needed for the preparation of 2.5 liters of 1*N* solution?

8. What weight of dissolved HCl does 500 ml. of a 0.5*N* solution of hydrochloric acid contain?

9. What weight of dissolved acid does 600 ml. of a 0.4*N* solution of sulfuric acid contain?

10. It is desired to make up 800 ml. of a 0.2*N* solution of sodium phosphate, Na_3PO_4. What weight of the salt is needed?

11. Write out complete directions for preparing 500 ml. of a 0.5N solution of sodium bisulfate, $NaHSO_4$.

12. Concentrated hydrochloric acid solution has a density of 1.18 g./ml. and contains 37.38 per cent HCl by weight. What is the *normal* concentration of this acid?

13. What weight of dissolved HCl does 50 ml. of 1N hydrochloric acid solution contain? Calculate from the balanced equation the amount of sodium hydroxide required to neutralize this quantity of HCl. If this required amount of sodium hydroxide was contained in 50 ml. of solution, what would be its *normal* concentration?

14. The density of concentrated sulfuric acid is 1.84 g./ml. and 96 per cent of the solution by weight is H_2SO_4. This concentrated acid contains how many gram-molecular weights of H_2SO_4 per liter of solution? What is its normality (*normal* concentration)?

15. What is the *normal* concentration of a solution which contains 2.49 g. of blue vitriol, $CuSO_4 \cdot 5H_2O$, dissolved in 100 ml. of solution? (Include the five molecules of water in the molecular weight, but only the copper need be considered in determining the equivalent weight of the substance.)

21. NEUTRALIZATION

By the definition of normal solution, we know that 1 liter of 1.0N acid contains the proper quantity of acid required to exactly neutralize 1 liter of 1.0N base. Each of these solutions would contain exactly 1 gram-equivalent weight of solute, acid, or base, respectively.

Also by definition, the gram-equivalent weight of any acid is exactly equal in its acid reaction capacity to the base reaction capacity of a gram-equivalent weight of any base.

Now suppose we were given 1 liter of 1.0N acid which is to be exactly neutralized by a solution of 0.5N base. Comparison of the normality figures for the two solutions shows us that the acid solution is twice as concentrated as the base solution. So, to neutralize 1 liter of 1.0N acid solution, we would have to use 2 liters of the 0.5N base solution. In other words, the 1 liter of 1.0N acid solution contains 1 gram-equivalent weight of acid, and to neutralize it, exactly 1 gram-equivalent weight of base must be used. Then to obtain the required quantity of the base, 2 liters of the 0.5N base solution must be used.

Note that the acid solution is twice as concentrated as the base solution, while the volume of the acid solution is half that of the base solution. A moment's reflection will enable us to recall that a similar relation was encountered in the expression of Boyle's law for gases. By analogy, then, we may write

$$V_a \times N_a = V_b \times N_b$$

where the subscript a refers to the acid solution and the subscript b refers to the base solution.

Or $$1 \text{ liter} \times 1.0N = 2 \text{ liters} \times 0.5N$$

and $$1 = 1$$

Examples

1. What volume 0.5N sodium hydroxide solution is required to neutralize 80 ml. of 1.0N hydrochloric acid?

Solution. Tabulation of data shows

$V_a = 80$ ml. acid $V_b = ?$

$N_a = 1.0N$ $N_b = 0.5N$

Since $$V_a \times N_a = V_b \times N_b$$

Then $$80 \times 1.0 = V_b \times 0.5$$

And $$80 = 0.5\ V_b$$

or $$V_b = 160 \text{ ml. of the } 0.5N \text{ base solution that must be used}$$

2. If 60 ml. of a 0.3N solution of hydrochloric acid solution was required to neutralize 45 ml. of a potassium hydroxide solution, what was the normality of the base solution?

From the statement of the problem it is seen that

$V_a = 60$ ml. $V_b = 45$ ml.

$N_a = 0.3N$ $N_b = ?$

Substituting in the general formula,

$$V_a \times N_a = V_b \times N_b$$

gives

$$60 \times 0.3 = 45 \times N_b$$

or

$$N_b = \frac{60 \times 0.3}{45} = \frac{18}{45} = 0.4$$

Therefore, the concentration of the potassium hydroxide solution is $0.4N$.

When acid and base solutions are to be used for the neutralization of one by the other, such that the concentration of one solution is known and is to be utilized in determining the concentration of the other, instruments capable of very accurate measurement of the reacting volumes are used. These instruments are called burettes, and the volumes of the solutions used can be measured to the nearest 0.01 ml. A dye whose color is very susceptible to the relative acidity or alkalinity of the solution is used to determine the exact neutralization point and is known in this application as an *indicator*. The process of so exactly determining the equivalent reacting volumes of solutions is called *titration*.

NAME______________________________ DESK No.______________

PROBLEMS

1. One liter of $0.1N$ solution of a base will neutralize what volume of (*a*) $0.1N$ acid; (*b*) $0.01N$ acid; (*c*) $0.2N$ acid; (*d*) $0.5N$ acid; (*e*) $0.15N$ acid? (These can be answered by inspection.)

2. What volume of $0.05N$ acid will be required to neutralize 1 liter of (*a*) $0.1N$ base; (*b*) of $0.02N$ base; (*c*) of $0.25N$ base? (These can be answered by inspection.)

3. What is the normality of an acid solution, 22.5 ml. of which was neutralized by 27.5 ml. of a $0.54N$ sodium hydroxide solution?

4. If 15.30 ml. of a solution of potassium hydroxide solution was neutralized by 21.50 ml. of $0.0964N$ sulfuric acid, what would be the normality of the potassium hydroxide solution?

5. A 10-ml. sample of vinegar was diluted to exactly 100 ml., and 18.66 ml. of the diluted vinegar solution required 16.78 ml. of $0.1051N$ sodium hydroxide solution for neutralization. What was the normality of the acetic acid of the vinegar?

6. A 2.766-g. sample of pure oxalic acid, $H_2C_2O_4 \cdot 2H_2O$, was dissolved in enough distilled water to make 500 ml. of solution, and 23.56 ml. of this acid solution was then used to neutralize 20.72 ml. of a sodium hydroxide solution. What is the normality of the base solution?

7. It is desired to prepare 10 liters of a sulfuric acid solution whose concentration is about 0.2*N*. What volume of 6*N* sulfuric acid should be used for dilution to the required volume? (The dilution formula is the same as that for neutralization.)

8. Concentrated hydrochloric acid has a density of 1.195 g./ml. and is 37.38 per cent HCl by weight. If 8.33 ml. of the concentrated acid solution was diluted to 1 liter volume, what was the normality of the resulting solution?

9. Twenty-five milliliters of an acid solution was diluted to exactly 250 ml. in a volumetric flask, and 27.67 ml. of the diluted solution required 19.35 ml. of a standardized base for neutralization. The base had been standardized by titration against acid potassium phthalate $H(KC_8H_4O_4)$, which acts as a monobasic acid. The data for the standardizing follow: 0.4085 g. of $H(KC_8H_4O_4)$ was dissolved in distilled water and required 21.80 ml. of the base for neutralization. What is the normality of the original acid solution?

NAME____ DESK No.__________________

10. A solution of hydrochloric acid is $0.25N$ in concentration.

(*a*) If 36.00 ml. of this acid solution was required for neutralizing 27.00 ml. of a sodium hydroxide solution, what was the normality of the sodium hydroxide solution?

(*b*) How many grams of dissolved HCl should 400 ml. of this acid solution contain?

(*c*) How many grams of dissolved NaOH would be contained in 750 ml. of the sodium hydroxide solution?

(*d*) How many moles of sodium chloride should be produced when 300 ml. of the hydrochloric acid solution is exactly neutralized by the sodium hydroxide solution?

11. A standard solution of barium hydroxide is exactly $0.05N$.

(*a*) How many grams of dissolved $Ba(OH)_2$ will 200 ml. of this solution contain?

(*b*) If 60 ml. of the barium hydroxide solution was used to neutralize 48 ml. of a solution of sulfuric acid, what was the normality of the acid?

(*c*) How many moles of dissolved $Ba(OH)_2$ would 10 liters of the barium hydroxide solution contain?

(*d*) How many liters of the acid solution is required to contain 1 gram-equivalent weight of H_2SO_4?

22. FARADAY'S LAWS OF ELECTRODEPOSITION

The conduction of the electric current through a solution is due to electrochemical action at the surfaces of the electrodes where the current enters and leaves the solution. Faraday first recognized the laws governing this phenomenon in 1832, when he stated (1) that the amount of the electrochemical action at the electrodes is in direct proportion to the quantity of current passing through the solution and (2) that chemically equivalent quantities of substances undergo transformation at the electrodes by a given quantity of current.

The total current is measured in ampere-seconds, or coulombs. That is, a current of 1 amp. flowing for 1 sec. equals 1 amp.-sec. of current, or 1 *coulomb*. Thus a current of 10 amp. flowing for 5 sec. equals (10 × 5) amp.-sec., or 50 *coulombs*. It has been found that 96,540 *coulombs* of electricity will cause 1 gram-equivalent weight of a metal (*e.g.*, 31.75 g. of copper, 107.88 g. of silver, or 1.0078 g. of hydrogen) to be plated out at the cathode, and the same amount dissolve simultaneously at the anode.

The generator (or battery) which is used as the source of the electric current may be considered as an electron pump which forces the circulation of the electrons *through* the wiring circuit *from* the anode *to* the cathode. Thus we can consider the anode as being *positive* to the solution because the electrons are continually pumped out of it into the wire circuit. Similarly we can consider the cathode as *negative* to the solution because it is coated with the electrons pumped to it from the anode by the generator (or battery).

The reactions occurring at the surface of the anode involve the *giving up of electrons* to the anode by the anions and hence are *oxidation* reactions. The reactions at the *cathode* surface are all *reduction* reactions and involve the *acceptance of electrons* by the cations. Thus we can say that anodic reactions are oxidation processes, while the cathodic reactions are reductions. Note, however, that the amount of oxidation that occurs at the anode must be equaled in amount by the reduction occurring simultaneously at the cathode.

Examples

1. When 2540 coulombs of electricity is passed through a cupric sulfate solution, using copper electrodes, what weight of copper should be plated out on the cathode? (That is, what should be the gain in weight of the copper cathode? Remember also that in this case an equal amount of copper dissolves from the anode.)

Solution. 96,540 coulombs causes 1 gram-equivalent weight of copper to plate out, (63.57 ÷ 2) = 31.78 g. Therefore 2540 coulombs will cause a proportional amount of copper to plate out (at the cathode).

$$96{,}540:31.78 = 2540:X$$

and

$$X = \frac{2540 \times 31.778}{96{,}540} = 0.8361 \text{ g. copper}$$

2. Using a steady current of 0.6 amp., how long will it take to plate out 0.1 g. of silver from a silver nitrate solution?

Solution. The valence of silver is 1; therefore the atomic weight of silver, 107.88, is also the equivalent weight. Then

$$107.88 \text{ g. Ag}:96{,}540 \text{ coulombs} = 0.1 \text{ g. Ag}:X \text{ coulombs}$$

and

$$X = \frac{0.1 \times 96{,}540}{107.88} = 894.4 \text{ coulombs required}$$

Now coulombs = seconds × amperes; so

$$894.4 \text{ coulombs} = X \text{ sec.} \times 0.6 \text{ amp.}$$

and

$$X = \frac{894.4}{0.6} = 1491 \text{ sec.} = 24 \text{ min., } 51 \text{ sec.}$$

NAME ______________________________ DESK No. ______________

PROBLEMS

1. What weights, respectively, of (*a*) silver; (*b*) zinc will 10,000 coulombs of electricity plate out from solutions of their ions?

2. How many coulombs of electricity is required to plate out 10 g. each of (*a*) hydrogen; (*b*) oxygen? Write the equations for the electrode reactions.

3. A current of 5 amp. was allowed to flow through a solution of cadmium cyanide for 1 hr. (*a*) How many coulombs was this? (*b*) What weight of the cadmium should have been deposited on the surface of a platinum cathode?

4. A current of 15 amp. was permitted to flow through a solution of silver nitrate for 30 min. What weight of silver should have been deposited at the cathode? What weight of oxygen should have been deposited simultaneously at the anode (platinum)?

5. With a current of 10 amp. flowing, how much time would be required for the electrodeposition of 1 g. each of (*a*) zinc; (*b*) mercury (ic)?

6. What length of time is required for the deposition of 5 g. of mercury (ic) upon a copper cathode, if the current strength is maintained at 4 amp.?

7. When platinum electrodes are used for the electrolysis of a dilute solution of sodium fluoride, hydrogen is plated out at the cathode and oxygen is plated out at the anode. (*a*) If a current of 5 amp. is passed through the solution for 30 min., what weight of each gas will be deposited? (*b*) What volume will each gas occupy under standard conditions? (*c*) What are their volumes at 724 mm. and 23°C.?

8. A current was passed through a silver coulometer for exactly 3 hr., and it was found that the silver cathode had gained 2.365 g. in weight. What was the average current (amperes) during the plating period? What weight of copper would have been deposited from a copper plating solution (cuprous cyanide) connected in the same circuit?

9. It is desired to cover one side of a sheet of copper, 4 by 3 ft., with a 0.001-in. plate of silver. Calculate the weight of silver and the quantity of electricity required. If the plating is done at a current density of 6 amp./sq. ft. of surface being plated, how much time will the plating process require? (The specific gravity of silver is 10.5.)

10. Chlorine is prepared commercially by the electrolysis of a brine solution, using carbon anodes and steel cathodes. (*a*) What quantity of electricity is required to liberate 100 lb. of chlorine? (*b*) What weight of hydrogen will be liberated at the cathode at the same time? (*c*) What volume will each of these liberated gases occupy at standard conditions of temperature and pressure? (*d*) What volume (cubic feet) should the hydrogen occupy when compressed at 1000 p.s.i. at 27°C.?

23. THE AVOGADRO NUMBER

Avogadro was a contemporary of Dalton, Gay-Lussac, Priestley, Proust, Lavoisier, and many others who were making contributions to the advancement of chemistry as a science at the beginning of the nineteenth century. In 1811 he published the results of some of his investigations and came to the conclusion that equal volumes of gases, measured under similar conditions of temperature, contain equal numbers of molecules. This statement was made by interpreting Gay-Lussac's law of reacting volumes of gases (the volumes of gases taking part in, or produced by, chemical reactions can always be expressed by the ratio of small whole numbers) in terms of the then new atomic-molecular theory of Dalton.

With the conclusions of Gay-Lussac and Avogadro available to them, it was only necessary for the chemists of that time to assume reasonably that oxygen gas existed as a molecule composed of two atoms, *viz.*, O_2, and the atomic theory and atomic-weight system could have been developed immediately. Unfortunately, Avogadro's work was neither widely known nor recognized for its full significance, and it remained for Cannizzaro to call it to the attention of the world of chemistry nearly 50 years later.

Today the statement of Avogadro is accepted as a matter of course, and various textbooks of chemistry describe it as a law, a principle, or as a hypothesis. As a sort of monument to Avogadro, modern chemists have designated the *number* of molecules known to be contained in a *gram-molecular weight* of any substance as the *Avogadro number.* This number has a constant value for all substances and has been accurately measured by a wide variety of methods. Its value is 6.023×10^{23}, or 602,300,000,000,000,000,000,000.

The following problems illustrate some of the uses and applications of the *Avogadro number* in chemical calculations.

NAME________________________________ DESK No.______________

PROBLEMS

1. (*a*) What fraction of a mole is 1 liter of oxygen, measured at 0°C. and 760 mm. pressure? (*b*) This volume (1 liter, standard temperature and pressure) contains how many molecules of oxygen?

2. (*a*) 1 kg. of water corresponds to how many moles of water? (*b*) How many molecules does 1 kg. of water contain?

3. (*a*) How many molecules of hydrogen are contained in 1 liter of the gas measured at standard temperature and pressure? (*b*) How many molecules are contained in 1 liter of the gas at 0°C. and 2 atm. pressure? (*c*) How many molecules are in 1 liter of the gas at 0°C. and 22.4 atm. pressure?

4. A 10-g. sample of methyl alcohol, CH_3OH, was dissolved in 100 g. water. What is the ratio of the number of alcohol molecules to water molecules in the solution?

5. Recall the definition of the molal solution (page 97). Consider a $1m$ solution of cane sugar, $C_{12}H_{22}O_{11}$, in water. What is the ratio of the number of solute molecules to the number of solvent molecules?

6. Given a $0.1m$ solution of hydrochloric acid: (*a*) how many HCl molecules are present in each gram of water? (*b*) If the HCl is considered 90 per cent ionized, how many hydrogen ions are in each gram of water?

7. The density of water at 4°C. is 1 g./ml. (*a*) How many molecules of water are contained in 1 ml. of water at this temperature? (*b*) How many molecules are in one drop of water at 25°C. when the density is 0.9907 g./ml.? (See the Appendix.)

8. The solubility of silver chloride at 10°C. is 0.000089 g. per 100 g. of water. (*a*) What is the molality of this solution? (*b*) If the AgCl is considered 100 per cent ionized, how many silver ions are contained in each gram of water?

NAME____ DESK NO.______________________

9. Recently a new unit of weight has been adopted, the *avogram*. It is defined as the division of the *gram* by the Avogadro number.

$$\frac{1\text{ g.}}{6.023 \times 10^{23}} = 1 \text{ avogram}$$

What is the weight of the sodium atom, expressed in this new unit? The weight of the oxygen atom?

10. Concentrated sulfuric acid is 93.19 per cent pure H_2SO_4, and its density is 1.835 g./ml. (*a*) What is the molarity of this solution? (*b*) What is the molality? (*c*) How many molecules of H_2SO_4 are there per 100 molecules of water in this solution?

11. A container of pure nitrogen was exhausted by a vacuum pump to a pressure of 0.76 mm. (*a*) Express this pressure in atmospheres. (*b*) How many molecules were left in each liter in the tank?

12. The density of table salt, occurring as the mineral *halite*, is 2.16 g./cc. (*a*) How many atoms of sodium are contained in a 1 cc. of this substance? (*b*) Inspect the diagram of the sodium chloride crystal in your text, and then calculate the distance in centimeters between the centers of the sodium atoms.

24. CHEMICAL EQUILIBRIUM: SOLUBILITY PRODUCT

In a qualitative sense it is said that the salt, silver chloride, is insoluble. Actually this is not true because its solubility, though very small, has been measured and is known to be 0.000,157,7 g. per 100 g. of water. By using this solubility figure along with the knowledge gained from experiments on reaction control (*i.e.*, concentration and temperature effects), calculations can be made which furnish much important information regarding the entire process of precipitation.

Now 0.000,157,7 g. of AgCl per 100 g. of water is the same ratio as 0.001577 g. of AgCl per 1000 g. of water, or 0.001577 g. of AgCl per liter of water (if we assume the density of water to be 1.000 g./ml.). This solubility figure can be converted to moles of AgCl per liter by dividing by the molecular weight of the silver chloride (107.88 + 35.457 = 143.337).

$$\frac{0.001577 \text{ g./liter}}{143.337 \text{ g./mole}} = 0.000011 \text{ mole AgCl/liter} = 1.1 \times 10^{-5} \text{ moles/liter}$$

If this dissolved AgCl is 100 per cent ionized, and according to theory it should be, then the ionization equation shows the relation

$$\underset{\substack{(1.1 \times 10^{-5} \text{ mole/liter,} \\ 100\% \text{ ionized})}}{AgCl} \rightarrow \underset{\substack{(1.1 \times 10^{-5} \\ \text{mole/liter)}}}{Ag^+} + \underset{\substack{(1.1 \times 10^{-5} \\ \text{mole/liter)}}}{Cl^-}$$

Note that in this solution there are no molecules at all because the dissolved silver chloride has become ionized completely, yielding 1.1×10^{-5} mole of Ag^+ and 1.1×10^{-5} mole of Cl^- per liter of water.

Now the probability that any one silver ion may recombine with a chloride ion to form a molecule of solid AgCl is dependent upon the number of chloride ions present. Similarly we must say that the probability that any one chloride ion may recombine with a silver ion to form a molecule of solid AgCl is dependent upon the number of silver ions present. In the saturated solution of silver chloride we see that 1.1×10^{-5} mole of Ag^+ can be in equilibrium with 1.1×10^{-5} mole of Cl^- without precipitating any solid silver chloride. This relation is so exact that we can describe the reaction with a mathematical equation

$$[Ag^+][Cl^-] = K$$

where $[Ag^+]$ and $[Cl^-]$ represent the maximum permissible molar concentrations of the silver ion and chloride ion, respectively, that can exist in equilibrium with one another in the same solution without forming any precipitate. Now K has a constant value for each temperature and can be evaluated as the product obtained by multiplying the solubility concentration of the silver ions by the solubility concentration of the chloride ions, as shown in the ionization equation above. Since many mathematical constants are commonly designated by K, we shall call this one the solubility-product constant, and let it be represented by K_{sp}(AgCl). Then

$$[Ag^+][Cl^-] = (1.1 \times 10^{-5})(1.1 \times 10^{-5}) = 1.21 \times 10^{-10} = K_{sp}(\text{AgCl})$$

Going back to the saturated solution of silver chloride, experiment shows that the addition of a single drop of hydrochloric acid solution (or a drop of a solution of any other chloride) results in the increase of the chloride-ion concentration so that the chances, or probability, of any one silver ion recombining with a chloride ion are greatly increased, thus causing the formation of some solid AgCl which appears as a precipitate. This action removes some of the silver ions along with some of the chloride ions, until an equilibrium between the ions left in the solution is again established. Once more we can write

$$[Ag^+][Cl^-] = K_{sp} = 1.21 \times 10^{-10}$$

It must be noted here that the new values of $[Ag^+]$ and $[Cl^-]$ are not the same as those for the saturated solution of the silver chloride. But the same K_{sp} value does hold, for it is dependent only on the temperature, which we are considering as constant throughout our considerations. Because of this relation, it is possible to calculate the concentration conditions under which a precipitate may be formed and how much of one ion may be left in solution when the concentration of the other ion is increased, as just described. In the following examples various applications of the solubility-product principle will be illustrated.

Attention might well be called to the fact that the mathematical relation just described is not new to you. Review of Chap. 13 will show that in the case of Boyle's law,

$$P_0 \times V_0 = P_1 \times V_1 = P_2 \times V_2 = \text{———} = K \text{ (temperature constant)}$$

Examples

1. Hydrochloric acid is to be added to a solution of silver nitrate, $AgNO_3$, until the chloride-ion concentration reaches the value of 0.1 mole/liter. What will be the concentration of the silver ion left in the resulting solution?

Solution. The equation for the reaction shows

$$HCl + AgNO_3 \rightarrow HNO_3 + AgCl \text{ (insoluble)}$$

We have seen in the preceding discussion that the value of $K_{sp}(AgCl)$ is 1.21×10^{-10}, or 0.000,000,-000,121, and now $[Cl^-]$ is to be maintained at 0.1 mole/liter. Substitution in the solubility-product formulation gives

$$[Ag^+][Cl^-] = [Ag^+](0.1) = 1.21 \times 10^{-10} = K_{sp}(AgCl)$$

Solving for the concentration of the silver ion gives

$$Ag^+ = \frac{1.21 \times 10^{-10}}{0.1} = 1.21 \times 10^{-9} \text{ mole of Ag per liter}$$

Since the atomic weight of silver is 107.88, then

$$(1.21 \times 10^{-9})(107.88) = 1.3 \times 10^{-7} \text{ g. silver ion per liter}$$

If the concentration of the silver ion is desired as grams per 100 g. of water, then dividing by 10 gives 1.3×10^{-8} g. of the silver ion left in each 100 g. of water.

2. The solubility of barium chromate, $BaCrO_4$, is 0.00393 g./liter at 25°C. Calculate the value of K_{sp}.

Solution. Convert the given solubility concentration to moles per liter by dividing by the molecular weight of the barium chromate, 253.4.

$$\frac{0.00393 \text{ g./liter}}{253.4 \text{ g./mole}} = 0.0000155 \text{ mole } BaCrO_4 \text{ per liter} = 1.55 \times 10^{-5} \text{ mole/liter}$$

According to the ionization equation for $BaCrO_4$,

$$\underset{\substack{(1.55 \times 10^{-5} \text{ moles/liter,} \\ 100\% \text{ ionized)}}}{BaCrO_4} \rightarrow \underset{\substack{(1.55 \times 10^{-5} \\ \text{moles/liter)}}}{Ba^{++}} + \underset{\substack{(1.55 \times 10^{-5} \\ \text{moles/liter)}}}{CrO_4^{--}}$$

Now $$K_{sp}(BaCrO_4) = [Ba^{++}][CrO_4^{--}] = (1.55 \times 10^{-5})(1.55 \times 10^{-5})$$

and $$K_{sp} = 2.4 \times 10^{-10} \text{ (at 25°C.)}$$

3. The solubility-product constant for lead carbonate, at 25°C., is known to be 4.0×10^{-14}. What is the solubility of lead carbonate in (*a*) moles/liter; (*b*) grams per 100 g. of water?

Solution. (*a*) Reference to the ionization equation of lead carbonate shows

$$PbCO_3\text{(dissolved solid)} \rightarrow Pb^{++} + CO_3^{--}$$

Then $$K_{sp}(PbCO_3) = [Pb^{++}][CO_3^{--}] = 4.0 \times 10^{-14}$$

According to the ionization equation, whatever number of moles of lead carbonate are dissolved when ionization is complete will give to the solution exactly that same number of moles of barium ion *and* that same number of moles of the carbonate ion. Thus we may write

$$[Pb^{++}] = [CO_3^{--}] = \sqrt{4.0 \times 10^{-14}} = \sqrt{4.0} \times \sqrt{10^{-14}} = 2 \times 10^{-7} \text{ mole/liter}$$

and

$$PbCO_3 \text{ (dissolved solid)} = 2 \times 10^{-7} \text{ mole/liter}$$

(*b*) Now the molecular weight of $PbCO_3$ is 267.2 (207.2 + 12 + 48 = 267.2), and

$(2 \times 10^{-7})(267.2) = 5.34 \times 10^{-5}$ g. of lead carbonate dissolved in 1 liter of the water solution

The solubility in grams per 100 g. of water would be one-tenth as much, or 5.34×10^{-6} g.

4. The solubility of lead iodide, PbI_2, is 0.0701 g. per 100 g. of water at 25°C. What is the solubility-product constant?

Solution. Convert the solubility concentration to moles per liter. (The molecular weight of PbI_2 is 461.22.)

$$\frac{10 \times 0.0701}{461.22} = 0.00152 = 1.52 \times 10^{-3} \text{ mole/liter}$$

Write the ionization equation.

PbI_2	$\rightarrow$	Pb^{++} +	I^-	$+ I^-$
(0.00152 mole/liter, 100% ionized)		(0.00152 mole/liter)	(0.00152 mole/liter)	(0.00152 mole/liter)

In this case PbI_2 furnishes the same number of moles of each of the three ions. Note that two of the three ions have the same identity, and that the *total* concentration of the iodide ion therefore will be (2×0.00152) mole/liter. These iodide ions have no intelligence, of course, and do not know that they are different, or that we may refer to them as the first iodide ion and the second iodide ion. And they really are not different at all—they all act simply as iodide ions. Thus in the re-forming of the PbI_2 molecule, all the iodide ions present represent chances for any lead ion to react.

$$Pb^{++} + I^- \rightarrow PbI^+$$

and all the remaining iodide ions then are influential in providing opportunity for the reaction

$$PbI^+ + I^- \rightarrow PbI_2 \text{ (solid)}$$

In other words the total concentration of the iodide ion must be considered effective in causing each of these two stepwise reactions to occur. The formulation for calculating the solubility-product constant of PbI_2 is

$$[Pb^{++}][I^-][I^-] \rightarrow [Pb^{++}][I^-]^2 = K_{sp}(PbI_2)$$

Substitution of the values we have calculated for the concentrations of the lead and iodide ions gives

$$K_{sp}(PbI_2) = (0.00152)(2 \times 0.00152)^2 = (0.00152)(0.00304)^2$$
$$= (1.52 \times 10^{-3})(9.24 \times 10^{-6}) = 14 \times 10^{-9}$$

or

$$K_{sp}(PbI_2) = 1.4 \times 10^{-10}$$

NAME____ DESK No.____________________

PROBLEMS

1. At room temperature calcium carbonate is soluble only to the extent of 0.0011 g. per 100 g. of water. Express this solubility as moles per liter, and then calculate the solubility-product constant.

2. Solubility measurements show that 0.000165 g. of silver bromide dissolves in 1 liter of water at 25°C. Calculate the solubility-product constant for this salt.

3. Only 0.000233 g. of barium sulfate can be dissolved in 100 g. of water at room temperature. (*a*) Calculate the solubility product for this substance. (*b*) What is the equilibrium concentration of the barium ion in a solution where the sulfate-ion concentration is to be maintained at 0.001 mole/liter?

4. The solubility-product constant for barium carbonate at 25°C. is 8.1×10^{-9} (see the Appendix). Calculate the solubility of this salt in moles per liter.

5. Barium sulfate is to be precipitated by adding barium chloride to a solution known to contain dissolved sulfates (*e.g.*, $NaHSO_4$ and $MgSO_4$, as in a water analysis). If the solution is made $0.01N$ with respect to the barium ion, at 25°C., to what concentration will the sulfate ion be reduced? [See the Appendix for $K_{sp}(BaSO_4)$ at 25°C.]

6. The solubility of silver chromate, Ag_2CrO_4, at 0°C., is 1.4×10^{-3} g. per 100 g. of water. Calculate the value of the solubility-product constant at this temperature.

7. At 10°C. only 0.000089 g. of silver chloride dissolves in 100 g. of water. At 18°C. the solubility is 0.000157 g. per 100 g. of water. And at 25°C., 0.000187 g. of AgCl will dissolve in 100 g. of water. (*a*) Calculate the solubility-product constant for each of these temperatures, and (*b*) then predict what the value should be at 30°C.

8. What weight of silver iodide can be dissolved in 2 liters of water at 25°C.? [See the Appendix for the value of $K_{sp}(AgI)$.]

NAME______________________________ DESK NO.______________

9. Ferrous hydroxide, $Fe(OH)_2$, at 10°C. has a measured solubility of 6.7×10^{-3} g./liter of water. Calculate its solubility-product constant.

10. What can be the maximum concentration of the zinc ion in moles per liter in a solution in which the hydroxide ion is to be raised and maintained at $0.01N$? (K_{sp} for zinc hydroxide is 5.0×10^{-17}.)

11. In a solution saturated with silver bromide at 25°C., how many silver ions are present in each milliliter of the solution? [See the Appendix for K_{sp}(AgBr). HINT: Remember Avogadro's number?]

12. Using the data of Problem 5 above, calculate the number of sulfate ions that would be present in 1 ml. of saturated solution of barium sulfate at 25°C., and then compare with the number present after the barium-ion concentration had been raised to $0.01N$.

13. Consider a solution of silver nitrate, $AgNO_3$, which is $0.01M$ and is 100 per cent ionized. (*a*) How many moles of silver ion is contained in 1 ml. of the solution? (*b*) How many individual silver ions are present in 1 drop of the solution? (1 ml. = 20 drops, Appendix.)

14. One drop of $0.1M$ hydrochloric acid was added to 10 ml. of water. Ignoring the slight increase in volume, calculate (*a*) the molarity of the resulting acid solution; (*b*) the concentration of the chloride ion in moles per liter.

15. One drop of $0.1M$ hydrochloric acid was added to 20 ml. of a saturated silver chloride solution. Ignoring the slight increase in volume, calculate (*a*) the total resulting chloride concentration and (*b*) the concentration of the silver ion remaining in solution in equilibrium. (*c*) What weight of AgCl should have precipitated?

16. Five milliliters of $0.1M$ barium chloride solution was added to 15 ml. of $0.01M$ sodium sulfate solution. Assume the volumes are additive, and calculate the weight of the barium sulfate that should have precipitated.

25. CHEMICAL EQUILIBRIUM: IONIZATION CONSTANT

The relations that exist in a dilute solution of a weak electrolyte can be expressed in a formulation somewhat similar to that we have used for the slightly soluble substances.

Consider a $0.1M$ solution of acetic acid. Accurate physicochemical measurements show that the acetic acid is 1.34 per cent ionized at this concentration. This is to say that 1.34 per cent of the total acetic acid present, or 0.00134 mole ($0.0134 \times 0.1 = 0.00134$), is acting as free ions, H^+ and $C_2H_3O_2^-$, while the remainder, 0.09866 mole ($0.1 - 0.00134 = 0.09866$), acts as undissociated molecules of the acid. Application of these data to the ionization equation shows that at equilibrium

(Total = 0.1 mole/liter, 1.34 per cent ionized)

$$\underset{(0.09866\ \text{mole/liter})}{HC_2H_3O_2} \rightarrow \underset{(0.00134\ \text{mole/liter})}{H^+} + \underset{(0.00134\ \text{mole/liter})}{C_2H_3O_2^-}$$

Note: The 0.00134 mole of the acetic acid which has become ionized has produced equimolar quantities of the hydrogen ion and acetate ion.

The formulation which we are to use is

$$\frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = K_i$$

where the concentrations of each component are expressed in moles per liter, as indicated by our use of the brackets, and K_i is known as the ionization constant. Substitution of the concentration values we obtained above gives

$$\frac{(0.00134)(0.00134)}{(0.09866)} = 0.0000183 = K_i \text{ (acetic acid)}$$

Now this value of K_i for acetic acid, while obtained with the measured values of concentrations for a $0.1M$ solution, actually holds and is constant for all concentrations of the acid, as long as we continue to work at the same, or constant, temperature. Thus if we desire to know the degree of ionization that exists in a solution of acetic acid which is $0.01M$, we can use the value of K_i which we have just now obtained in making the calculations necessary to yield the answer we want.

Example. Calculation of the hydrogen-ion concentration in a $0.01M$ solution of acetic acid.

Let X equal the number of moles of acetic acid that has ionized in the $0.01M$ solution. Then $(0.01 - X)$ equals the residual number of moles of acetic acid in the molecular form in the solution, and X equals the number of moles of each hydrogen ion and acetate ion that have been produced. Reference to the ionization equation then shows

(Total = 0.01 mole/liter)

$$\underset{(0.01 - X)\ \text{moles/liter}}{HC_2H_3O_2} \rightarrow \underset{X\ \text{moles/liter}}{H^+} + \underset{X\ \text{moles/liter}}{C_2H_3O_2^-}$$

Setting these values for the concentration of the various constituents into the formulation for K_i gives

$$\frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = 0.000{,}018{,}3 = \frac{(X)(X)}{(0.01 - X)} = \frac{X^2}{(0.01 - X)}$$

Performing the indicated operations and then transforming the equation yields

$$X^2 + 0.000,018,3X - 0.000,000,183 = 0$$

This last is a quadratic equation, which must be solved if we are to find the answer to our question. A solution for the value of X is not readily apparent by either the method of factoring or of completing the square. Since the equation is already in the general form,

$$aX^2 + bX + c = 0$$

the *quadratic formula* can be applied easily.

$$X = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-0.000,018,3 \pm \sqrt{(0.000,018,3)^2 - 4(-0.000,000,183)}}{2}$$

and
$$X = \frac{-0.000,018,3 \pm (0.000,855)}{2}$$

Choosing the positive value of the second term of the numerator, since the negative would have no significance, we obtain

$$X = \frac{0.000,837,2}{2} = 0.000418, \text{ or } 0.00042 \text{ mole/liter}$$

Therefore the degree, or percentage, of the acetic acid that is ionized in the 0.01M solution is

$$\frac{0.00042 \text{ mole/liter}}{0.02 \text{ mole/liter}} = 0.042, \text{ or } 4.2\%$$

Short-cut Method of Calculation. Study carefully the original equation obtained by substituting the X values in the formulation for K_i.

$$K_i = 0.000,018,3 = \frac{X^2}{(0.01 - X)}$$

If the value of X is very small when compared with the value 0.01, then the actual value of $(0.01 - X)$ will be almost equal to 0.01 itself. Therefore, dividing the number $(0.01 - X)$ into the quantity X^2, as indicated, will give a quotient which is not greatly different from that which would be obtained by dividing the number, 0.01, into X^2. If we then ignore the X value in the divisor, our equation becomes

$$0.000,018,3 = \frac{X^2}{0.01}$$

Transformation gives

$$X^2 = \sqrt{0.000,000,183}$$

and taking square roots yields

$$X = 0.000427 \text{ mole/liter}$$

Comparison of this value for X, with that previously obtained by the more rigorous and exact algebraical method, shows that the error in the less exact method is rather small and usually can be ignored. In using this method of calculation, remember the basic assumption: the actual value of the X in the divisor must be small in comparison with the number from which it is subtracted.

Effect of a Common Ion

Thus far we have confined our attention to a dilute solution containing the equilibrium system composed of acetic acid and its ions. What should be the effect on the equilibrium in a 0.01M solution of acetic acid if sufficient potassium acetate is added so that the resulting concentration of the acetate ion reaches and is maintained at 0.1M? A summary of the data at hand shows that we know (1) the equilibrium formulation for the acetic acid solution, (2) the value of K_i for acetic acid already evaluated, and (3) now a new and definite value for the concentration of the acetate ion. We can calculate the new values of concentrations of the hydrogen ion and of the acetic acid molecule only if we can express these values in terms of each other. (The formulation for K_i consists of a total of four terms, and we know only two—the value of K_i itself, 1.83×10^{-5}, and of $[C_2H_3O_2^-]$, 0.1 mole/liter. A single equation cannot be solved for two unknown terms unless those two unknowns are related and can be expressed in terms of one of them, thus reducing the actual unknown to a single value. Fortunately in this case it is possible to do this.)

Let $[H^+] = X$ in the potassium acetate–acetic acid solution. Then $[HC_2H_3O_2] = (0.01 - X)$ moles/liter. Since we have the fixed condition that $[C_2H_3O_2^-] = 0.1$ mole/liter, we may write

$$K_i = 1.83 \times 10^{-5} = \frac{(X)(0.1)}{(0.01 - X)} = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

Successive transformations give

$$0.1X = (1.83 \times 10^{-5})(0.01 - X) = (1.83 \times 10^{-7}) - 0.000{,}018{,}3X$$

$$0.1X + 0.000{,}018{,}3X = 1.83 \times 10^{-7}$$

Note here that the addition of 0.000,018,3X to the term 0.1X does not give a total X value that is materially greater than the value of 0.1X and so can be neglected. (In other words, the sum 0.100,018,3X is negligibly larger than the value of 0.1X. Further, the use of the term 0.100,018,3X in solving for the value of X implies an accuracy that is not consistent with the other figures used, either in this calculation or in the original determination of the value of K_i.) Solving for X yields

$$X = \frac{1.83 \times 10^{-7}}{0.1} = 1.83 \times 10^{-6} \text{ moles } H^+ \text{ per liter}$$

Thus, by increasing the concentration of the acetate-ion concentration, through the addition of potassium acetate, to approximately 230 times that in the original 0.01M acetic acid solution (0.00042 mole/liter to 0.1 mole/liter), we reduced, or repressed, the concentration of the hydrogen ion to about $\frac{1}{230}$ of its original value (0.00042 mole/liter to 0.000,001,83 mole/liter).

This action, known as the common-ion effect, is perfectly general in its application. The addition of a common ion, in the form of some soluble and highly ionized compound, to a solution of a slightly ionized substance, causes a depression of the ionization of that already slightly ionized substance. The addition of ammonium chloride to a dilute solution of ammonium hydroxide (common ion = ammonium) depresses the ionization of the slightly ionized ammonium hydroxide and reduces the actual hydroxide-ion concentration in the resulting solution to a figure that is much smaller than that of the original ammonium hydroxide solution.

NAME________________________________ DESK NO.________________

PROBLEMS

1. Nitrous acid ionizes according to the equation $HNO_2 \rightarrow H^+ + NO_2^-$. In its $0.1M$ solution it is 6.5 per cent ionized. Calculate the ionization constant K_i for this acid.

2. Hydrocyanic acid, HCN, is a very weak acid and is ionized only to the extent of 0.00046 per cent in its $1.0M$ solution. Calculate the ionization constant K_i for this acid. ($HCN \rightarrow H^+ + CN^-$.)

3. Calculate the concentration of the hydrogen ion in a $1.0M$ solution of acetic acid, using the value of K_i as obtained in the introduction to this chapter. What is the percentage of ionization exhibited by $1.0M$ acetic acid?

4. Calculate the concentration of the hydrogen ion in $0.001M$ acetic acid solution by both the exact and the less exact methods. Is there any appreciable difference in your answers? Why?

5. It happens that the value of K_i for ammonium hydroxide is practically identical with that of acetic acid, 1.83×10^{-5}. Calculate the concentration of the hydroxide ion in a $0.05M$ solution of this weak base.

6. Cyanic acid, HCNO, is 3.2 per cent ionized in its $0.1M$ solution. Calculate the value of K_i for this acid. ($HCNO \rightarrow H^+ + CNO^-$.)

NAME____ DESK No.____________________

7. Carbonic acid, H_2CO_3, ionizes in two steps: $H_2CO_3 \rightarrow H^+ + HCO_3^-$ and $HCO_3^- \rightarrow H^+ + CO_3^{--}$. The value of K_i for the equilibrium of the first-step reaction is 3.5×10^{-7}. What is the concentration of the bicarbonate ion that exists in a 0.01*M* solution of the acid? (Ignore the second-step reaction.)

8. If 5 g. of ammonium chloride was added to 1 liter of 0.02*M* ammonium hydroxide solution and it was assumed that the NH_4Cl was completely ionized, calculate the concentration of the hydroxide ion in the resulting solution.

9. Calculate the hydroxide-ion concentration in a solution of ammonium hydroxide which is 0.004*M*. Use both the exact and the less exact methods of calculation and compare the answers.

10. A solution was prepared so that it contained both sodium hydroxide and ammonium hydroxide in equal concentrations, $0.01M$ each. Assume that the sodium hydroxide is completely ionized, and calculate the concentration of the ammonium ion in the solution.

11. The value of K_i for benzoic acid, $HC_7H_5O_2$, is 6.4×10^{-5}, and the ionization reaction is written $HC_7H_5O_2 \rightarrow H^+ + C_7H_5O_2^-$. Calculate the actual number of hydrogen ions that is present in 1 ml. of $0.01M$ solution.

12. Calculate the actual numbers of potassium ions, hydrogen ions, acetate ions, and acetic acid molecules in 1 ml. of the potassium acetate–acetic acid solution described in the introduction to this chapter.

NAME____ DESK No.____________________

26. PROBLEMS: GENERAL AND REVIEW

1. An oxygen sample measured 600 ml. at 15°C. and 750 mm. pressure, dry (no water vapor). What weight of mercuric oxide was decomposed in obtaining this sample? ($2HgO \rightarrow 2Hg + O_2$.)

2. What volume of hydrogen, to be collected over water at 17°C. and 800 mm. pressure, should be displaced by 100 g. of aluminum in the reaction $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$?

3. Ammonia gas can be prepared in the laboratory by the reaction $NaOH$ (solid) + NH_4Cl (solid) + heat $\rightarrow$ $NaCl$ (solid) + H_2O + NH_3. What weight of ammonium chloride must be used for the preparation of 5 liters of ammonia which is to be measured in the dry condition at 37°C. and 5 atm. pressure?

4. Given the reaction $FeS + 2HCl \rightarrow FeCl_2 + H_2S$. (*a*) How many moles of hydrogen sulfide could be obtained from 1 lb. of iron sulfide which is 87 per cent pure FeS? (*b*) This hydrogen sulfide would occupy what volume at 33°C. and 3 atm. pressure?

5. A sodium chloride solution is 8 per cent NaCl by weight, and its density is 1.056 g./ml. Express its concentration in both *molal* and in *molar* units.

6. Concentrated hydrochloric acid is 37.38 per cent pure HCl, and its density is 1.19 g./ml. Calculate (and compare) both the molarity and molality of this solution.

NAME________________________________ DESK No.________________

7. A 5-g. sample of naphthalene, $C_{10}H_8$, was dissolved in 80 g. of benzene. Calculate the freezing point of the solution.

8. Calculate the boiling point of a solution which contains 40 g. of glycerin, $C_3H_8O_3$, in 80 g. of water. A radiator solution of similar composition would freeze at what Fahrenheit temperature?

9. Calculate the freezing point of a solution which contains 40 g. of methyl alcohol, CH_4O, in 80 g. of water. Compare this solution with that of the preceding problem in the following respects: percentage composition by weight; mole ratio, solute to solvent; and advantages as radiator antifreeze solution.

10. A 30-g. sample of a certain substance was dissolved in 180 g. of water, and the resulting solution froze at $-1.72°C$. Calculate the molecular weight of the dissolved substance.

11. When 5 g. of an unknown substance was dissolved in 150 g. of carbon tetrachloride, the resulting solution had a boiling point of 80.58°C. What is the approximate molecular weight of the substance?

12. Camphor has a freezing point of 178.4°C. and a K_f value of 37.7°C. When a 0.2-g. sample of a new compound was dissolved in 10 g. of camphor, the resulting solution exhibited a freezing point of 172.5°C. What is the molecular weight of the new substance?

13. If 360 ml. of a gaseous substance, measured at 710 mm. pressure and at 33°C., weighed 1.600 g., what was the approximate molecular weight of the compound?

NAME____ DESK No.__________________

14. The formula for ammonia gas is NH_3. What should be the weight of 500 ml. of this gas at standard temperature and pressure? What is the weight of 500 ml. of this gas measured at 37°C. and 700 mm. pressure?

15. A 0.700-g. sample of gas occupied a volume of 640 ml. when measured at 27°C. and 730 mm. pressure. What is the approximate molecular weight of the gas?

16. A compound contains 26.31 per cent calcium, 42.10 per cent sulfur, and 31.57 per cent oxygen by weight. What is its simplest formula?

17. Total analysis of a compound showed it to be water of crystallization 6.57 per cent, potassium 28.15 per cent, hydrogen 0.73 per cent, carbon 17.50 per cent, and oxygen 46.68 per cent. What is its simple formula?

18. A 0.1187-g. sample of a certain metal displaced 0.2157 g. of silver from $AgNO_3$ solution. What is the equivalent weight of the metal?

19. A sample of a metal weighed 0.545 g. and displaced 219 ml. of hydrogen which was collected over water at 27°C. and 740 mm. pressure. Calculate the equivalent weight of the metal.

20. A certain metallic element was found to have a specific heat of 0.093 cal./g./deg., and a 0.5-g. sample displaced 1.650 g. of silver from silver nitrate solution. What is the exact atomic weight of the metal, and what is its valence?

21. A compound was found to be 92.28 per cent carbon and 7.72 per cent hydrogen. When 4.5 g. of the substance was dissolved in 60 g. of carbon tetrachloride, the resulting solution boiled at 81.63°C. What is the exact formula and molecular weight of the compound?

NAME____ DESK No.____________________

22. Calculate the exact atomic weight of an element from the following data: specific heat is 0.24 cal./g./deg., and 0.180 g. of the element (metal) displaces 0.6356 g. of copper from a solution of $CuSO_4$.

23. Two compounds, the first a gas, the second a liquid, have identical percentage composition, 85.7 per cent carbon and 14.3 per cent hydrogen. A 331-ml. sample of the gas measured at 33°C. and 720 mm. pressure weighed 0.35 g. One gram of the liquid was dissolved in 15 g. of carbon tetrachloride, and the resulting solution had a freezing point of −25.05°C. What is the exact formula and molecular weight of each of these two substances?

24. We have found that the gram-molecular volume of a gas is 22.4 liters at 0°C. and 1 atm. pressure. Under these same conditions of temperature and pressure, calculate the ounce-molecular volume of a gas, expressing the volume in cubic feet.

25. How many cubic feet of nitrogen, measured at standard conditions, should be released by the heat decomposition of 4 lb. of ammonium nitrite? (NH_4NO_2 + heat $\rightarrow$ $N_2 + 2H_2O$.)

26. Given the reaction, $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$. What weight of sodium carbonate will be needed to yield 100 cu. ft. of carbon dioxide gas which is to be measured at 18°C. and 0.9 atm. pressure?

27. A 100-g. sample of sodium chloride was heated with concentrated sulfuric acid. (*a*) What weight of HCl gas should have been produced? (*b*) The HCl gas was dissolved in sufficient water to make 500 ml. of solution. What was the normality of the solution?

NAME________________________________ DESK No.______________

28. Suppose you have exactly 6 g. of pure silver nitrate, $AgNO_3$. (*a*) This is how many moles? (*b*) Dissolved in 500 ml. of solution, what is the resulting normality? (*c*) What was the normality of the acid solution if 33.65 ml. of this silver nitrate solution was required for complete reaction by a 50-ml. sample of a hydrochloric acid solution?

29. What is the (*a*) molarity and (*b*) normality of a phosphoric acid solution which is 70 per cent pure H_3PO_4 and has a density of 1.526 g./ml.?

30. Glacial acetic acid is 100 per cent pure $HC_2H_3O_2$ and has a density of 1.050 g./ml. What is the normality of this "solution"?

31. A solution of sodium hydroxide is 10 per cent NaOH by analysis, and its density is 1.109 g./ml. What is the molar concentration of the solution?

32. What volume of 1.5N NH_4OH solution can be produced from the NH_3 resulting from the hydrolysis of 400 g. of Mg_3N_2?

33. If 5.160 g. of pure and dry sodium carbonate was dissolved in distilled water to make exactly 500 ml. of solution and if 27.86 ml. of this solution was then used in the titration of 31.24 ml. of a hydrochloric acid solution, calculate the normality of the acid solution. ($Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$.)

34. A 21.25-ml. sample of 0.185*N* sulfuric acid solution was used to neutralize 19.15 ml. of a sodium carbonate solution. (*a*) What is the normality of the sodium carbonate solution? (*b*) What weight of H_2SO_4 was contained in the 21.25 ml. of the solution used in the titration? (*c*) What weight of Na_2CO_3 would be contained in 1 liter of its solution? (*d*) What is the molar concentration of the sodium carbonate solution?

35. A 48.4-ml. sample of a hydrochloric acid solution required 36.3 ml. of 0.28*N* sodium hydroxide solution for neutralization. (*a*) What volume of the sodium hydroxide solution is required to contain exactly 1 gram-equivalent weight of NaOH? (*b*) Each milliliter of the sodium hydroxide solution contains what weight of dissolved NaOH? (*c*) What is the normality of the hydrochloric acid solution? (*d*) What weight of dissolved HCl does 1 liter of the hydrochloric acid solution contain?

36. A 36.8-ml. sample of $0.21N$ HCl solution was required for neutralizing 27.6 ml. of a sodium hydroxide solution. (*a*) What is the concentration (normality) of the NaOH solution? (*b*) What weight of dissolved HCl does 2 liters of the hydrochloric acid solution contain? (*c*) What volume of the acid solution is required to contain exactly 1 gram-equivalent weight of HCl? (*d*) What weight of dissolved NaOH does 1 ml. of the titrated sodium hydroxide solution contain?

37. What volume of commercial phosphoric acid of density 1.69 g./ml. and containing 85 per cent H_3PO_4 can be obtained by the action of sulfuric acid on 500 kg. of phosphate rock which is 87 per cent $Ca_3(PO_4)_2$? ($Ca_3(PO_4)_2 + 3H_2SO_4 \rightarrow 3CaSO_4 + 2H_3PO_4$.)

38. What weight of ammonium chloride must be used for the preparation of 4.5 liters of $2.5N$ NH_4OH solution? ($NH_4Cl + NaOH \rightarrow NaCl + H_2O + NH_3$; $NH_3 + H_2O \rightarrow NH_4OH$.)

NAME________________________________ DESK No.______________

39. What volume of NH_3 gas, to be measured over mercury at 21°C. and 720 mm. pressure, can be prepared by heating 112 g. of ammonium chloride, 96 per cent pure NH_4Cl, with solid NaOH?

40. A 100-g. sample of zinc oxide was dissolved in hydrochloric acid by the reaction $ZnO + 2HCl \rightarrow ZnCl_2 + H_2O$. The hydrochloric acid solution used was 10 per cent HCl by weight. What weight of the acid solution was required for the reaction?

41. A 100-g. sample of sodium acetate, $NaC_2H_3O_2$, is to be heated with concentrated sulfuric acid. (*a*) What weight of pure acetic acid should be obtained in the distillation? (*b*) What volume of sulfuric acid, specific gravity 1.86 and 96 per cent pure H_2SO_4, will be consumed in the process? (*c*) If the acetic acid thus produced is dissolved in 2 liters of solution, what is the resulting molarity?

42. Nitric acid is to be made by distillation with concentrated sulfuric acid and 1700 g. of sodium nitrate. What volume of 6*N* nitric acid solution can be prepared from the product? ($NaNO_3 + H_2SO_4 \rightarrow NaHSO_4 + HNO_3$.)

43. Sodium iodate, $NaIO_3$, occurs as an impurity in Chile saltpeter to the extent of 0.2 per cent by weight. If the recovery processes are 93 per cent efficient, how much elementary iodine could be recovered from 10,000 tons of the saltpeter?

44. What weight of bone ash, which is 70 per cent $Ca_3(PO_4)_2$, will be required to supply the phosphorus necessary for preparing 20,000 lb. of 85 per cent phosphoric acid solution? If the specific gravity of this 85 per cent acid solution is 1.689, what is its volume in gallons?

NAME______________________________ DESK No.______________

45. One ton of orpiment ore contained 576 lb. of arsenic metal. If experience shows that the metallurgical processes of recovery are 91 per cent efficient, what per cent of the ore was actually As_2S_3?

46. What weight of limestone, 90 per cent pure calcium carbonate, must be treated with hydrochloric acid to obtain 200 liters of dry carbon dioxide, measured at 3 atm. pressure and at a temperature of 31°C.?

47. How many pounds of pure nitric acid, HNO_3, can be prepared from 1 ton of sodium nitrate, $NaNO_3$, which is 89 per cent pure? How many gallons of solution, 68 per cent pure HNO_3 and having specific gravity of 1.405, can be made from this acid? (Density of water = 8.33 lb./gal.)

48. What volume of hydrogen would you expect to obtain, collected over water at 30°C. and 724 mm. pressure, from the reaction of hydrochloric with 100 g. of an alloy which is 80 per cent zinc and 20 per cent aluminum?

49. A retort was charged with 3000 kg. of Chile saltpeter, 96 per cent pure $NaNO_3$, for making nitric acid by reaction with concentrated sulfuric acid. Allowing a 10 per cent excess, what volume of the sulfuric acid, in gallons, should be used if the acid were 95 per cent pure H_2SO_4 and had a specific gravity of 1.83? What is the volume in liters?

50. Assume that the reacting gaseous nitrogen, N_2, is converted quantitatively (100 per cent) in the cyanamide process into ammonia, NH_3. (*a*) How many cubic feet of ammonia gas, measured under standard conditions of temperature and pressure, should 1 lb. of nitrogen yield? (*b*) To what weight of ammonium sulfate, $(NH_4)_2SO_4$, is 1 lb. of nitrogen equivalent?

51. What volume of concentrated sulfuric acid, specific gravity 1.825 and containing 94 per cent H_2SO_4, is required to convert 1000 lb. of phosphate rock, 88 per cent $Ca_3(PO_4)_2$, into superphosphate, $Ca(H_2PO_4)_2(caSO_4)_2$?

NAME____ DESK NO.____________

52. Given the reaction: $Cu + 4HNO_3$ (conc.) $\rightarrow Cu(NO_3)_2 + 2H_2O + 2NO_2$. (*a*) What weight of pure HNO_3 will be used in dissolving 100 g. of copper? (*b*) What volume of concentrated nitric acid solution, 70 per cent pure HNO_3 and density of 1.42 g./ml., is required to react with 1500 g. of copper? (*c*) When cooled to −9.3°C., the NO_2 becomes a liquid whose density is 1.45 g./ml., and whose formula is written N_2O_4. What volume of N_2O_4 could be produced from the NO_2 released in the reaction described in (*b*)?

53. Iron sulfide is to be used as a source for hydrogen sulfide which is to be used in turn for precipitating antimony sulfide.

$$FeS \text{ (solid)} + 2HCl \rightarrow FeCl_2 + H_2S \text{ (gas)}$$

$$3H_2S \text{ (gas)} + 2SbCl_3 \rightarrow Sb_2S_3 \text{ (solid)} + 6HCl$$

(*a*) What volume of hydrogen sulfide, measured at standard conditions, should 11 g. of iron sulfide yield? (*b*) The volume of hydrogen sulfide produced from 1 mole of iron sulfide should precipitate what weight, in grams, of the antimony sulfide? (*c*) One liter of 0.1N antimony chloride solution should require what volume of hydrogen sulfide (standard temperature and pressure) to convert the antimony to insoluble sulfide?

54. Copper reacts with hot concentrated sulfuric acid according to the equation $Cu + 2H_2SO_4$ (hot, conc.) $\rightarrow CuSO_4 + 2H_2O + SO_2$. (*a*) What volume of sulfur dioxide gas, measured at standard conditions, should 250 g. of copper liberate? (*b*) What weight of the pure H_2SO_4 will be consumed in reacting with the 250 g. of copper? (*c*) If the copper sulfate formed (from the 250 g. of copper) was dissolved in 10 liters of water solution, what would be the molarity of the solution? (*d*) If the concentrated sulfuric acid solution is 96 per cent pure H_2SO_4 and its density is 1.86 g./ml., what volume of the acid should be consumed in reacting with the 250 g. of copper?

NAME______________________________ DESK No.__________

55. In the electrolytic purification of copper, the electrode reactions are anode, Cu (metal) $\rightarrow$ $Cu^{++} + 2e$; cathode, $Cu^{++} + 2e \rightarrow$ Cu (metal). (*a*) If 100,000 coulombs passed through the solution, what weight of copper should deposit at the cathode? (*b*) What weight of copper becomes copper ion at the anode when a current of 100 amp. passes through the solution for 5 hr.?

56. In the Nelson-Vorce-Hooker cell reaction, $2NaCl + 2H_2O$ + electric current $\rightarrow H_2 + Cl_2 + 2NaOH$, the cathode reaction is $2H^+ + 2e^- \rightarrow H_2$, and the anode reaction is $2Cl^- \rightarrow Cl_2 + 2e^-$. (*a*) How many coulombs of electricity is required for the conversion of 100 kg. of NaCl into the corresponding electrolytic products? (*b*) The consumption of 1000 kg. of sodium chloride in this reaction should produce how many tons of sodium hydroxide? (*c*) With the consumption of 100 lb. of sodium chloride in this reaction, how many cubic feet of hydrogen, measured at 27°C. and 5 atm., should be produced? (*d*) If the chlorine liberated with the consumption of 100 kg. of sodium chloride was compressed to a liquid, density 1.557 g./ml., what would be its volume?

Desk No.__________

1. Assume that gaseous nitrogen is converted quantitatively (100 per cent) into nitric acid, HNO_3, the Haber-Ostwald process. (*a*) What volume of nitrogen, at standard conditions of mperature and pressure, is required to produce 1 gram-equivalent weight of nitric acid? (*b*) What v e of nitrogen, in cubic feet and at standard conditions, is necessary for the production of 1 ton of nitric acid? (*c*) If 1 mole of nitric acid is dissolved in 250 ml. of solution and then is neutralized by 150 ml. of a sodium hydroxide solution, what must be the normality of the NaOH solution? (*d*) How many moles of ammonia, NH_3, should be produced by 1000 liters of nitrogen at 200 atm. pressure and 500°C.? (See text for equations, if necessary.)