SCIENTIFIC (EXPONENTIAL) NOTATION

Putting Ordinary Numbers into Scientific Notation:

Scientists (an those studying science) frequently must deal with numbers that are very large or very small. Have you met Avogadro's number (6.02×10^{23})? Or have you calculated the wavelength of red light (6.10×10^{-7} m)? If those numbers weren't written the way they are, all of us who must deal with them would be spending much of our time just counting the zeros that separate the figures from the decimal point. To avoid that kind of time wasting, a method of writing very large and very small numbers was invented. It is called scientific or exponential notation.

The rules for writing numbers in scientific notation are

1. The first figure is a number from 1 to 9.
2. The first figure is followed by a decimal point and then the rest of the figures.
3. Then multiply by the appropriate power of 10.

Here are some examples: $425 = 4.25 \times 100 = 4.25 \times 10^2$ $0.00098 = 9.8 \times 1/10000 = 9.8 \times 0.0001 = 9.8 \times 10^{-4}$

$36000 = 3.6 \times 10000 = 3.6 \times 10^4$ $0.0135 = 1.35 \times 1/100 = 1.35 \times 0.01 = 1.35 \times 10^{-2}$

Putting Scientific Notation into Ordinary Numbers:

Going from exponential notation to everyday expression is a bit easier than writing an ordinary number in scientific notation because the number that you are dealing with comes with built-in instructions. If the number ends with x 103 (a positive exponent), it is simply telling you to multiply by 1000 or move the decimal point three places to the right. If the number ends with x 10-4 (a negative exponent), it is telling you to divide by 10000, or move the decimal point four places to the left

Here are some examples: $3.24 \times 10^2 = 3.24 \times 100 = 324$ $7.8156 \times 10^3 = 7.8156 \times 1000 = 7815.6$

Sometimes, we have to put in placeholder zeros to show the decimal value of the number. Remember, that these placeholder zeros are not significant.

$9.2 \times 10^4 = 9.2 \times 10000 = 92000$ $5.68 \times 10^{-3} = 5.68 \times 1/1000 = 5.68 \times 0.001 = 0.00568$

$4.235 \times 10^6 \times 4.235 \times 1000000 = 4235000$ $2.9 \times 10^{-5} = 2.9 \times 1/100000 = 2.9 \times 0.00001 = 0.000029$

I. Express each of the following numbers in scientific notation. (Remember to pay attention to zeros. include them if they are significant, do not include them If they are not)

1) 325 __________ 3) 70 __________ 5) 96,400 __________

2) 0.361 __________ 4) 0.0428 __________ 6) 0.00573 __________

II. Write each of the following as ordinary numbers. (Watch for significant zeroes)

7) 3.64×10^4 _________________ 10) 3.88×10^{-2} _________________

8) 2.97×10^{-4} _________________ 11) 6.285×10^3 _________________

9) 3.9734×10^5 ._________________ 12) 5.65×10^{-1} ___________________

III. Round each of the following to the nearest whole number.

13) 56.912 __________ 15) 3.4125 __________ 17) 251.7817 __________

14) 0.5182 __________ 16) 112.511 __________ 18) 63.541 __________

SIGNIFICANT DIGITS / SIGNIFICANT FIGURES

The accuracy of the final answer to a problem depends upon the accuracy of the numbers used to express each measurement used. The accuracy of any measurement depends upon the instrument, which is used, and upon the observer. The digits in an answer which imply more accuracy than the measurements justify are not significant and should dropped so that those digits which remain truly imply the accuracy of the original measurements. The remaining digits are called significant digits or significant figures. Significant digits or significant figures consist of the definitely known digits plus one estimated digit.

Measured numbers -- For example, if the mass of a chemical sample is 31.72 grams, the measurement is said to have four significant digits. The last digit, 2, has probably been estimated but the mass of the chemical sample is definitely between 31.7 g and 31.8 g. Any zeros present to the right of a decimal point are important, because they indicate where the uncertain or estimated digit is.

Exact numbers have an infinite number of significant digits. Exact numbers are "counts," not measurements. You may have exactly 24 students in a class, but you cannot have 24.3 students. Furthermore, relationships such as 1 day = 24 hours contain exact numbers. Exact numbers can have any number of zeros to the right of the decimal point or last nonzero digit. Exact numbers do not limit the number of significant digits in a calculation.

Determining which zeros are significant -- because zeros must be written both as placeholders and as indicators of the precision of the measurement, we must learn how to distinguish between them. The following rules are used to determine the number of significant digits:

1. All non-zero digits are significant:
   Ex - 421 - three significant digits.

2. Zeros between significant digits are significant ("trapped zeros")
   Ex - 4.021; 4.201, 402.1 - four significant digits

3. Zeros after (to the right of) the decimal point are significant ("trailing zeros")
   Ex - 42.100; 421.00; 4210.0 - five significant digits

4. Zeros after (to the right of) the decimal point and before (to the left of) nonzero digits are not significant ("leading zeros").
   They are placeholders!
   Ex – 0.00421 - three significant digits

5. A zero standing alone before (to the left of) the decimal point is not significant. It is a placeholder!
   Ex - 0.421 - three significant digits

6. Zeros after (to the right of) all nonzero digits when no decimal point is present are not significant. They are placeholders.
   Ex – 421,000 - three significant digits

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When the decimal is Present, start counting with the first nonzero number on the left. Keep counting until you fall off.	When the decimal is Absent, begin counting with the first non-zero number on the right. Keep counting until you falloff.

Rules for Rounding:

1. If the eliminated digit is less than 5, round down.
2. If the eliminated digit is greater than 5, round up.

Addition and Subtraction of Significant Figures

The least accurate measurement determines the accuracy of answer in an addition or subtraction problem. The answer may contain only as many decimal places as the least accurate measurement. In other words, when adding and subtracting, round your answer to the least number of decimal places in any of the numbers.

Multiplication and Division of Significant Figures

In multiplication and division, the answer should have the same number of significant figures and the factor having the least number of significant digits (the least precise measurement) in the problem. In other words, when multiplying and dividing, round your answer to the least number of significant figures in any of the factors.

Determine the number of significant digits in each of the following:

19) 0.02 _______
20) 142 _______
21) 0.020 _______
22) 0.073 _______
23) 501 _______
24) 1.071 _______

Round each of the following to three significant digits.

25) 88.473 _______
26) 69.95 _______
27) 8505 _______
28) 0.000056794 _______
29) 976450 _______
30) 67.048 _______

Solve the following problems using your calculator and express each answer with the correct number of significant figures.

31) 1234
32) 0.023
33) 890
34) 91010
35) 9010.0
36) 1090.0010
37) 0.00120
38) 3.4×10^{4}
39) 9.010×10^{-2}
40) 1020010
41) 780.
42) 1000
43) 918.010
44) 0.0001
45) 8120
46) 7.991×10^{-10}
47) 23.4 x 14
48) 7.895 + 3.4
49) 0.0945 x 1.47
50) 0.005 - 0.0007
51) 7.895 / 34
52) 0.2 / 0.0005
53) 350.0 - 200
54) 27.68 - 14.369
55) 3.08 x 5.2
56) 0.0036 x 0.02
57) 4.35 x 2.74 x 3.008
58) 35.7 x 0.78 x 2.3
59) 3.76 / 1.62
60) 0.075 / 0.030

Unit Conversion

The metric system is based on units of ten. This makes converting between units by dimensional analysis very easy to do. Below are some of the most common metric prefixes. These prefixes can be applied to any unit desired. You will be expected to memorize the ones in bold type.

Prefix	Symbol	Equivalency Statement		
kilo	**k**	**1 km**	**=**	**1000 m**
hecto	h	1 hm	=	100 m
deca	da	1 dam	=	10 m
deci	**d**	**1 dm**	**=**	**0.1 m**
centi	**c**	**1 cm**	**=**	**0.01 m**
milli	**m**	**1 mm**	**=**	**0.001 m**
micro	µ	1 µm	=	10^{-6} m
nano	**n**	**1 nm**	**=**	**10^{-9} m**

In addition, there are three other conversion factors you should memorize. They are:

1 cm^{3} = 1 mL **1 dm^{3} = 1 L** **1 mole = 6.02 x 10^{23} molecules**

Unit Conversion Practice

YOU MUST SHOW ALL WORK (WITH UNITS!) IN ORDER TO RECEIVE CREDIT.

61) 35 mL = __________ dL

62) 950 g = __________ kg

63) 4500 mg = __________ g

64) 25 cm = __________ mm

65) 0.025 mg = _________ cg

66) 1000 L = __________ kL

DENSITY

I. Units of Density

A. **Mass**: amount of matter contained in any object. Does not depend on gravity; therefore the mass of an object is the same anywhere in the universe.

B. **Volume**: The amount of space that an object takes up. There are three basic ways to measure volume: 1) Direct measurement and calculations; used for regular shaped solid, 2) volume displacement; used for irregular shaped solids, and 3) direct measurement with a graduated cylinder, used for liquids.

Metric units of volume: cubic centimeters (cm^3) liters (L), milliliters (mL)

C. **Density**: ratio of mass to volume. The more closely packed the atoms or molecules of a substance are packed the higher the density. There are two ways to determine density from experimental data: 1) calculation from specific mass and volume measurements; or 2) using a graph of mass vs. volume from a series of measurements - the density of the substance is determined by the slope of the line. The greater the slope the higher the density of the substance.

Note about calculations of density: If the correct units are used, the formula **D=** ***M/V*** *can* be used. However, since chemistry teachers are notorious for **NOT** giving their students the same units they want the answer in. **DIMENSIONAL ANALYSIS** is the best problem-solving method!

Metric units of density: Solids - g/cm^3; Liquids - g/mL; Gases - g/L

Density Homework

*Show all work, **INCLUDING UNITS**. Do not forget to perform any necessary conversions.*

67) A sample of seawater weighs 158 g and has a volume of 156 mL. What is the density, in g/mL, of seawater?

68) Table salt has a density of 2.16 g/cm^3. A cylindrical box holds 425 g of salt. What is the volume, in cm^3, occupied by the salt in the box (ignore the spaces between the crystals).

69) What is the mass of ethyl alcohol that exactly fills a 200.0 mL container. The density of ethyl alcohol is 0.789 g/mL.

70) A block of lead has the dimensions 4.5 cm by 5.2 cm by 6.0 cm. The block weighs 1.587 kg. From this information, calculate the density of lead in g/cm^3.

71) Ethyl alcohol is added to a beaker with mass 204.88 g until the beaker and alcohol together weigh 253.2 g. What volume, in cL, was added to the beaker? The density of ethyl alcohol is 0.789 g/mL.

72) A student determines that the mass of a metal rod is 12.0 g. He places 10.0 mL of water in a graduated cylinder. When the rod is placed in the water, the graduated cylinder then reads 15.0 mL. What is the density of the metal rod?