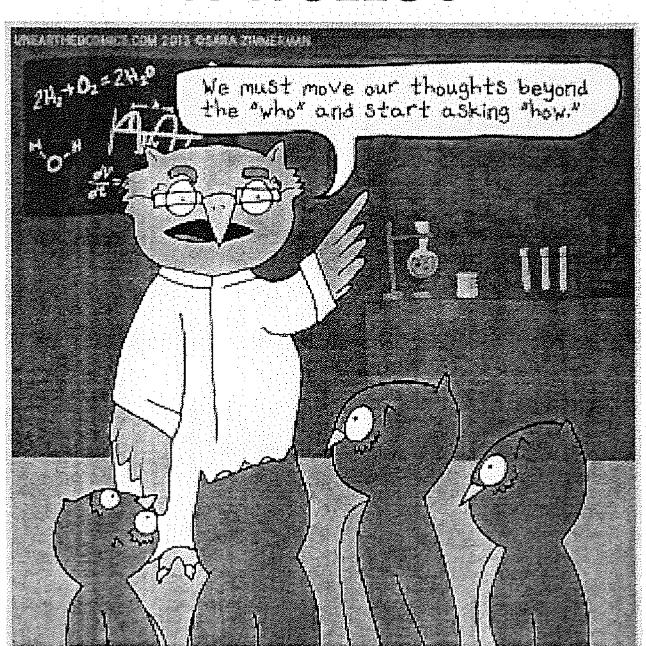
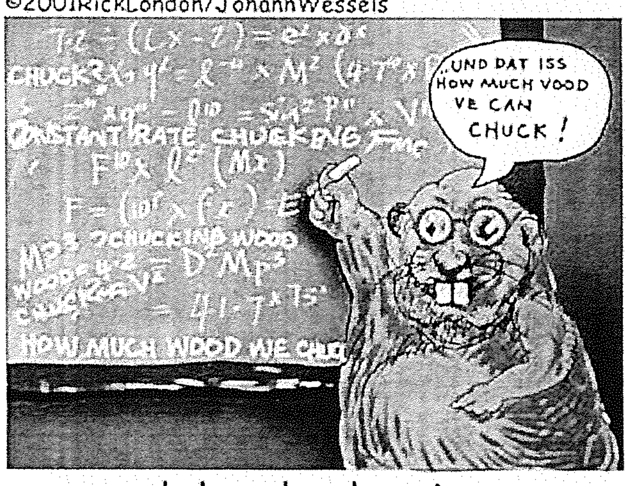
Measurements and Calculations Packet





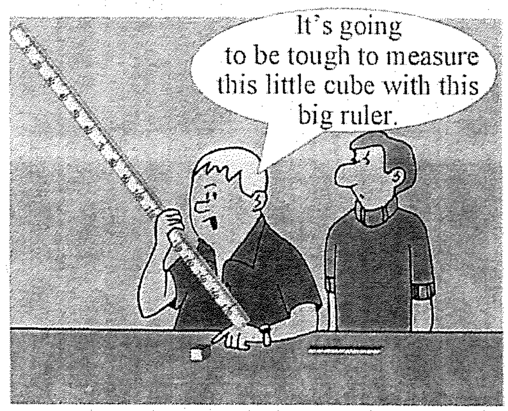
woodchuck physics

Name ______________________

Date ______________ Period ______

Metric Units and Prefixes

You wouldn't measure the thickness of an eyelash with a meter stick. It's just too big to be convenient. For that matter, a meter is too large to be a convenient unit for measuring the thickness of an eyelash. The metric system uses a system of prefixes to show fractions and multiples of the basic units. The basic units are meters (m) to measure distance, grams (g) to measure mass, and liters (L) to measure volume. Some of the important prefixes are shown below:



Kilo (k)	=	1,000	=	10^{3}
Deci (d)	=	0.1	=	10^{-1}
Centi (c)	=	0.01	=	10^{-2}
Milli (m)	=	0.001	=	10^{-3}
micro (μ)	=	0.000001	=	10^{-6}
nano (n)	=	0.000000001	=	10^{-9}
pico (p)	=	0.000000000001	=	10^{-12}

For each group of values listed below, write the items in order from largest to smallest. Then state whether the values represent distance, mass, or volume.

Values	Order	Type (distance, mass, or volume)
1. 10 km, 10 pm, 10μm, 10 dm	______________	______________
2. 0.5 μL, 0.5 nL, 0.5 cL	______________	______________
3. 1.2 mm, 1.2 km, 1.2 cm	______________	______________
4. 3.5 cg, 3.5 g, 3.5 ng, 3.5 μg	______________	______________
5. 0.25 μL, 0.25 dL, 0.25 cL	______________	______________
6. 7.3 g, 7.3 pg, 7.3 kg, 7.3 mg	______________	______________
7. 4 pL, 4μL, 4kL, 4mL, 4 dL	______________	______________
8. 8μm, 8 pm, 8 m, 8 km, 8 nm	______________	______________
9. 0.1 kL, 0.1 L, 0.1 μL, 0.1 pL	______________	______________
10. 5.6 dg, 5.6 kg, 5.6 pg, 5.6 μg	______________	______________

<u>Significant Digits</u>

Alternate Rule for Significant Digits

Here is an alternate rule for determining significant digits that Mr. McNamara taught me last year. He, in turn, learned it from a show on television. If I could credit the person who made it up, I would. The rule is really a "trick", which might allow students to get the correct answers without really understanding the concepts. I would recommend that students only use this as a secondary method, for the purpose of checking their answers.

When you look at the number in question, you must determine if it has a decimal point or not. If it has a decimal, you should think of "**P**" for "**P**resent". If the number does not have a decimal place, you should think of "**A**" for "**A**bsent".

Example, for the number 35.700, think "**P**", because the decimal is **p**resent.

For the number 6500, you would think "**A**", because the decimal is **a**bsent.

Now, the letters "**A**" and "**P**" also correspond to the "**A**tlantic" and "**P**acific" Oceans, respectively. Now, assume the top of the page is north, and imagine an arrow being drawn toward the number from the appropriate coast. Once the arrow hits a nonzero digit, it and all of the digits after it are significant.

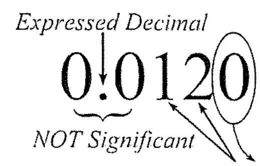
Each number that we record as a measurement contains a certain number of significant digits, which show accurate or estimated digits. When we do calculations our answers cannot be more accurate than the measurements that they are based on. We must be careful to follow the following rules whenever we perform calculations in Chemistry class.

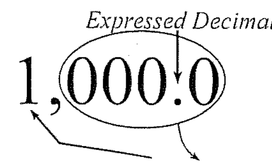
Chemistry: Form WS1.4.2B　　Name ______________________

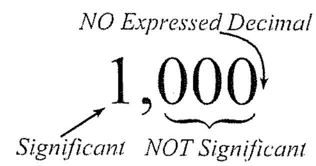
MATTER　　Date ______________ Period ____

Counting Significant Digits

An index card is 12.65 cm long ... approximately. The last digit is estimated since the smallest space on the ruler is 0.1 cm. The same index card is also 126,500 µm long. The 5 is still the estimated digit. The zeros are only place holders. They are not significant. Significant digits are the ones that are measured and the one (and *only* one) that is estimated. All nonzero numbers are significant. Place holding zeros, the leading zeros between the decimal and the first nonzero digit or the trailing zeros in a number that has no expressed decimal, are not significant. One way of recognizing significant digits is the Atlantic-Pacific rule. When the decimal is **a**bsent, count from the first nonzero digit toward the **A**tlantic coast. When the decimal is **p**resent, count from the first nonzero digit toward the **P**acific coast.







Pacific

Decimal **P**resent

Count toward the right from the first nonzero digit.

Atlantic

Decimal **A**bsent

Count toward the left from the first nonzero digit.

Tell the number of significant digits in each of the following measurements.

1. 48 cm ________
2. 306.2 g ________
3. 0.329 m ________
4. 83.9520 °C ________
5. 3700 mm ________
6. 400. cm^3 ________
7. 71.60 g ________
8. 0.00432 mm ________
9. 10.0 kg ________
10. 3.60×10^{15} sec ________
11. 6.24×10^{-4} m ________
12. 82.000 g ________

SIGNIFICANT DIGITS: PROBLEM SET I

I. Give the correct number of significant digits in each of the following numbers:

_____ 1. 6200.
_____ 2. 45,000,000
_____ 3. 0.0029
_____ 4. 10 grams
_____ 5. 30.00
_____ 6. 0.000420
_____ 7. 280 cars
_____ 8. 9.00
_____ 9. 6.023 x 10^{23}

10) 0.3760 _____
11. 860.001 _____
12. 19,000 _____
13. 620. _____
14) 50 dots _____
15. 1.0 x 10^{10} _____
16. 40 meters _____
17. 20,100,000 _____
18. 1,000.000 _____

II. Round each of the following numbers to **two** significant digits:

__________ 19. 63,488
__________ 20. 85.00
__________ 21. 0.00007449
__________ 22. 29.71

23. 59.98 __________
24. 0.98029 __________
25. 371,883 __________
26. 440.1 __________

III. Give the correct number of significant digits in the answer to each of the following problems. **DO NOT** calculate the answer.

_____ 27. 345.9 - 6
_____ 28. 3030. x 290
_____ 29. $(1.02)^5$
_____ 30. 0.25
_____ 31. 70 ÷ 512
_____ 32. 891.57 + 60

33. $(201)^4$ _____
34. 3 × 64 _____
35. 6680. + 23.1 _____
36) 27.90 ÷ 4060 _____
37. 102 x 6900 _____
38. 7402 + 20 _____

IV. Calculate each of the following problems and express your answer with the correct number of significant digits.

__________ 39. 36 ÷ 12
__________ 40. 0.36
__________ 41. 512.3 + 30
__________ 42. 3 x 16

43. $(2.4)^2$ __________
44. 123.09 - 1.7 __________
45. 7.01 x 6.3 __________
46. 845 ÷ 45.0 __________

Significant Digits

Multiplying and Dividing

RULE: When multiplying or dividing, your answer may only show as many significant digits as the multiplied or divided measurement showing the least number of significant digits.

Example: When multiplying 22.37 cm x 3.10 cm x 85.75 cm = 5946.50525 cm^3

We look to the original problem and check the number of significant digits in each of the original measurements:

22.37 shows 4 significant digits.

3.10 shows 3 significant digits.

85.75 shows 4 significant digits.

Our answer can only show 3 significant digits because that is the least number of significant digits in the original problem.

5946.50525 shows 9 significant digits, we must round to the tens place in order to show only 3 significant digits. Our final answer becomes 5950 cm^3.

Adding and Subtracting

RULE: When adding or subtracting your answer can only show as many decimal places as the measurement having the fewest number of decimal places.

Example: When we add 3.76 g + 14.83 g + 2.1 g = 20.69 g

We look to the original problem to see the number of decimal places shown in each of the original measurements. 2.1 shows the least number of decimal places. We must round our answer, 20.69, to one decimal place (the tenth place). Our final answer is 20.7 g

Chemistry: Form WS1.4.2C   Name ____________________

MATTER   Date ______________ Period ____

Calculating With Significant Digits

Every measurement has some error associated with it. Even if you are extremely careful, the best you can do is estimate the last digit beyond where your measuring tool measures. This causes some trouble with calculations. If you are finding the area of a piece of land, for example, when you multiply the length by the width, you are multiplying estimates by estimates. This can only multiply the uncertainty. There are rules to keep extra uncertain numbers from cropping up in your calculations.

★ **multiplication and division** - the number of significant figures in a product or quotient is the same as the measurement with the smaller number of significant figures

> *Problem*
> 3.1415 × 2.25 = 7.068375
>
> *Correct number of Significant Figures* = 3
>
> *Solution* 7.07

★ **addition and subtraction** - the number of decimal places in the sum or difference is equal to the number of decimal places in the measured quantity with the smallest number of decimal places

> *Problem*
> 6.357 - 2.4 = 3.957
> *Correct number of Decimal Places* = 1
>
> *Solution* 4.0

Perform each of the following calculations, expressing the answer with the correct number of significant digits.

1) 3.482 cm + 8.51 cm + 16.324 cm ____________

2) 48.0032 g + 9.17 g + 65.4321 g ____________

3) 80.4 cm - 16.532 cm ____________

4) 106.5mL - 30. mL ____________

5) 48.2 cm × 1.6 cm × 2.12 cm ____________

6) 8.3 m × 4.0 m × 0.9823 m ____________

7) 64.34 cm^3 ÷ 8.149 cm ____________

8) 4.93 mm^2 ÷ 18.71 mm ____________

9) $\mathbf{0.57\,mL \times \frac{760\,mm}{740\,mm} \times \frac{273\,K}{250\,K}}$ ____________

10) $\mathbf{5.13\,g \times \frac{44.962\,a.m.u.}{115.874\,a.m.u.}}$ ____________

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Name ____________________

Date ______________ Period ______

Working With Scientific Notation

Once you know how to read and write numbers in scientific notation, you are ready to learn how to do calculations in scientific notation. Addition and subtraction follow a few simple rules: (1) numbers must be a multiple of the same power of 10; (2) the first factor can then be added or subtracted; and (3) the power of 10 is not affected. See the example to the right

Example

$1.35 \times 10^5 + 2.9 \times 10^4$

Procedure

① $2.9 \times 10^4 = 0.29 \times 10^5$

② 0.29×10^5
$\underline{+\,1.35 \times 10^5}$
1.64×10^5

Multiplication follows a different set of rules: (1) multiply the first factors; and (2) add the exponents. See below.

Example

$(2 \times 10^4) \times (1.5 \times 10^3)$

Result

3.0×10^7

Division follows still a third set of rules: (1) divide the first factors; and (2) subtract exponents

Example

$$\frac{3.0 \times 10^5}{2.0 \times 10^3}$$

Result

1.5×10^2

Based on the instructions and examples above, answer the questions below.

1) $2.1 \times 10^3 + 2.1 \times 10^2$ ____________
2) $3.25 \times 10^5 - 5.2 \times 10^3$ ____________
3) $8.7 \times 10^6 + 3.1 \times 10^7$ ____________
4) $9.63 \times 10^5 + 8.81 \times 10^4$ ____________
5) $1.3 \times 10^{-4} - 5.6 \times 10^{-5}$ ____________
6) $9.11 \times 10^{-1} + 3.27 \times 10^{-2}$ ____________
7) $4.2 \times 10^1 + 9.7 \times 10^{-1}$ ____________
8) $6.7 \times 10^0 - 1.3 \times 10^{-1}$ ____________
9) $5.55 \times 10^3 + 9.46 \times 10^4$ ____________
10) $7.0 \times 10^5 + 8.1 \times 10^3$ ____________
11) $3.72 \times 10^{-2} - 8.45 \times 10^{-4}$ ____________
12) $8.7 \times 10^{-9} + 9.4 \times 10^{-10}$ ____________
13) $(5.7 \times 10^4) \times (3.1 \times 10^2)$ ____________
14) $(4.5 \times 10^3) \div (1.5 \times 10^1)$ ____________
15) $(8.8 \times 10^4) \div (2.2 \times 10^5)$ ____________
16) $(2.3 \times 10^{-1}) \times (5.0 \times 10^{-1})$ ____________
17) $(1.24 \times 10^2) \div (4.0 \times 10^{-2})$ ____________
18) $(3.0 \times 10^{-4}) \times (5.5 \times 10^2)$ ____________
19) $(6.4 \times 10^3) \times (1.5 \times 10^8)$ ____________
20) $(3.3 \times 10^3) \times (2.0 \times 10^{-5})$ ____________
21) $(4.9 \times 10^{-2}) \div (7.0 \times 10^{-3})$ ____________
22) $(1.1 \times 10^7) \div (5.5 \times 10^4)$ ____________
23) $(7.2 \times 10^{-6}) \times (1.4 \times 10^9)$ ____________
24) $(9.6 \times 10^5) \div (1.2 \times 10^7)$ ____________

SCIENTIFIC NOTATION: PROBLEM SET I

I. Convert each of the following decimal numbers into the exponential form. Express your answers with one significant digit to the left of the decimal point. Show all significant digits.

1. 7.90
2. 960,000
3. 0.00000182
4. 0.415
5. 216,000,000,000
6. 0.00000030

II. Convert each of the following exponential numbers into the decimal form. Show all significant digits.

7. 7.12×10^{-2}
8. 8.90×10^{1}
9. 3.3×10^{-1}
10. 6.75×10^{6}
11. 4×10^{0}
12. 5.00×10^{-5}

III. Shift the decimal point in each of the following numbers expressing your answer with **one** significant digit to the left of the decimal point. Show all significant digits.

13. 23.9×10^{-3}
14. 0.006×10^{4}
15. 0.00075×10^{-6}
16. 9500×10^{5}
17. 17.60×10^{0}
18. 0.03×10^{0}

IV. Shift the decimal point in each of the following exponential numbers and give the **missing number**. Show all significant digits.

19. $7.5 \times 10^{3} = ? \times 10^{1}$
20. $7.89 \times 10^{-6} = ? \times 10^{-7}$
21. $0.15 \times 10^{22} = 15 \times 10^{?}$
22. $450 \times 10^{-10} = 4.5 \times 10^{?}$
23. $0.00012 \times 10^{-2} = ? \times 10^{-5}$
24. $0.004 \times 10^{2} = 4 \times 10^{?}$
25. $4.87 \times 10^{7} = ? \times 10^{9}$
26. $25.0 \times 10^{-5} = ? \times 10^{-4}$
27. $782 \times 10^{24} = 0.782 \times 10^{?}$
28. $0.89 \times 10^{-45} = 89 \times 10^{?}$
29. $800 \times 10^{1} = ? \times 10^{5}$
30. $60 \times 10^{1} = 0.0006 \times 10^{?}$

SCIENTIFIC NOTATION: PROBLEM SET II

Calculate each of the following problems. Express your answer in the **exponential form** with **one significant digit to the left of the decimal point** and with the **correct number of significant digits.**

1. $(2.0 \times 10^{3})(4.24 \times 10^{5}) =$

2. $(4.5 \times 10^{6}) + (1.5 \times 10^{6}) =$

3. $(4.0 \times 10^{-3})(1.2 \times 10^{-10}) =$

4. $(5.6 \times 10^{16}) \div (8 \times 10^{12}) =$

5. $(2.4 \times 10^{-5})^{2} =$

6. $(1.80 \times 10^{-12}) \div (1.5 \times 10^{-7}) =$

7. $(3.5 \times 10^{-8}) - (1.3 \times 10^{-8}) =$

8. $(4.9 \times 10^{19})^{1/2} =$

9. $(4 \times 10^{-2})(4.5 \times 10^{8})^{2} =$

10. $\dfrac{(1.5 \times 10^{-6})^{3}}{5 \times 10^{-2}} =$

11. $(2.16 \times 10^{-22})^{1/3} =$

12. $(8.3 \times 10^{-3}) + (2 \times 10^{-4}) =$

13. $(4.0 \times 10^{5})^{4} =$

14. $(6.22 \times 10^{18})(8.2 \times 10^{-3}) =$

15. $\dfrac{7.2 \times 10^{17}}{8.0 \times 10^{-1}} =$

DATE ________ PERIOD ________ NAME ____________________

Chem

1 Skill

Use with Chapter 2.

MATHEMATICS ASSESSMENT

Write the following numbers in scientific notation.

1. 156.90 ____________________
2. 12 000 ____________________
3. 0.0345 ____________________
4. 0.008 90 ____________________

Expand the following numbers.

5. 1.23×10^6 ____________________
6. 2.5×10^{-3} ____________________
7. 1.54×10^4 ____________________
8. 5.67×10^{-1} ____________________

Solve the following and put your answer in scientific notation.

9. $\frac{6.6 \times 10^{-8}}{3.3 \times 10^{-4}} =$ ____________________
10. $\frac{7.4 \times 10^{10}}{3.7 \times 10^{3}} =$ ____________________
11. $\frac{2.5 \times 10^{8}}{7.5 \times 10^{2}} =$ ____________________
12. $(2.67 \times 10^{-3}) - (9.5 \times 10^{-4}) =$ ____________________
13. $(1.56 \times 10^{-7}) + (2.43 \times 10^{-8}) =$ ____________________
14. $(2.5 \times 10^{-6}) \times (3.0 \times 10^{-7}) =$ ____________________
15. $(1.2 \times 10^{-9}) \times (1.2 \times 10^{7}) =$ ____________________
16. $(2.3 \times 10^{4}) \times (2.0 \times 10^{-3}) =$ ____________________

Give the number of significant digits in the following measurements.

17. 2.9910 m ____________________
18. 5600 km ____________________
19. 0.006 70 kg ____________________
20. 809 g ____________________

Solve the following problems and give the answer in the correct number of significant digits.

21. $\frac{2.674 \text{ m}}{2.0 \text{ m}} =$ ____________________
22. 5.25 L × 1.3 L = ____________________
23. 9.0 cm + 7.66 cm + 5.44 cm = ____________________
24. 10.07 g − 3.1 g = ____________________

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Chem

1 Skill

NAME ______________________________

Solve for x in the following problems.

25. $\frac{3x}{y} = \frac{6g}{b}$ ______________________________

26. $d = \frac{t}{x}$ ______________________________

27. $\frac{2x^2}{3} = dg$ ______________________________

28. $\frac{2\sqrt{x}}{c} = y$ ______________________________

Make the following conversions.

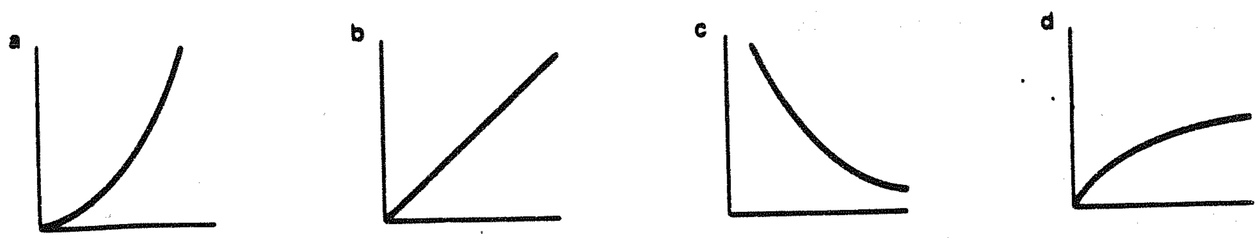
29. 4008 g = ______________________________ mg

30. 48 mL = ______________________________ L

31. 239 mm = ______________________________ cm

32. 38 kg = ______________________________ mg

Answer the questions that refer to the following graphs.



33. Which graph represents an inverse relationship? ______________________________

34. Which of the graphs could have the equation $y = kw^2$? ______________________________

35. Plot a graph of the data given in the following table.

x	y
0	2
0.5	8
1	14
2	26
3	38
4	50

36. What is the slope of the line? ______________________________

37. What is the value of y when $x = 4$? ______________________________

38. What is the value of y when $x = 6$? ______________________________

39. What is the value of x when $y = 0$? ______________________________

SCIENTIFIC NOTATION

Name ____________________

Scientists very often deal with very small and very large numbers, which can lead to a lot of confusion when counting zeros! We have learned to express these numbers as powers of 10.

Scientific notation takes the form of $M \times 10^n$ where $1 \leq M < 10$ and "n" represents the number of decimal places to be moved. Positive n indicates the standard form is a large number. Negative n indicates a number between zero and one.

Example 1: Convert 1,500,000 to scientific notation.
We move the decimal point so that there is only one digit to its left, a total of 6 places.

$1{,}500{,}000 = 1.5 \times 10^6$

Example 2: Convert 0.000025 to scientific notation.
For this, we move the decimal point 5 places to the right.

$0.000025 = 2.5 \times 10^{-5}$

(Note that when a number starts out less than one, the exponent is always negative.)

Convert the following to scientific notation.

1. 0.005 = ____________________
2. 5,050 = ____________________
3. 0.0008 = ____________________
4. 1,000 = ____________________
5. 1,000,000 = ____________________
6. 0.25 = ____________________
7. 0.025 = ____________________
8. 0.0025 = ____________________
9. 500 = ____________________
10. 5,000 = ____________________

Convert the following to standard notation.

1. 1.5×10^3 = ____________________
2. 1.5×10^{-3} = ____________________
3. 3.75×10^{-2} = ____________________
4. 3.75×10^2 = ____________________
5. 2.2×10^5 = ____________________
6. 3.35×10^{-1} = ____________________
7. 1.2×10^{-4} = ____________________
8. 1×10^4 = ____________________
9. 1×10^{-1} = ____________________
10. 4×10^0 = ____________________

DIMENSIONAL ANALYSIS (FACTOR LABEL METHOD)

Name ________________

Using this method, it is possible to solve many problems by using the relationship of one unit to another. For example, 12 inches = one foot. Since these two numbers represent the same value, the fractions 12 in/1 ft and 1 ft/12 in are both equal to one. When you multiply another number by the number one, you do not change its value. However, you may change its unit.

Example 1: Convert 2 miles to inches.

$$2 \text{ miles} \times \frac{5{,}280 \text{ ft}}{1 \text{ mile}} \times \frac{12 \text{ inches}}{1 \text{ ft}} = 126{,}720 \text{ in}$$

(Using significant figures, 2 mi = 100,000 in.)

Example 2: How many seconds are in 4 days?

$$4 \text{ days} \times \frac{24 \text{ hrs}}{1 \text{ day}} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ sec}}{1 \text{ min}} = 345{,}600 \text{ sec}$$

(Using significant figures, 4 days = 300,000 sec.)

Solve the following problems. Write the answers in significant figures.

1. 3 hrs = ________ sec
2. 0.035 mg = ________ cg
3. 5.5 kg = ________ lbs
4. 2.5 yds = ________ in
5. 1.3 yrs = ________ hr (1 yr = 365 days)
6. 3 moles = ________ molecules (1 mole = 6.02×10^{23} molecules)
7. 2.5×10^{24} molecules = ________ moles
8. 5 moles = ________ liters (1 mole = 22.4 liters)
9. 100. liters = ________ moles
10. 50. liters = ________ molecules
11. 5.0×10^{24} molecules = ________ liters
12. 7.5×10^{3} mL = ________ liters

METRICS AND MEASUREMENT

Name ____________________

In the chemistry classroom and lab, the metric system of measurement is used, so it is important to be able to convert from one unit to another.

mega	kilo	hecto	deca	**Basic Unit**	deci	centi	milli	micro
(M)	(k)	(h)	(da)	gram (g)	(d)	(c)	(m)	(μ)
1,000,000	1000	100	10	liter (L)	.1	.01	.001	.000001
10^6	10^3	10^2	10^1	meter (m)	10^{-1}	10^{-2}	10^{-3}	10^{-6}

Factor Label Method

1. Write the given number and unit.
2. Set up a conversion factor (fraction used to convert one unit to another).
   a. Place the given unit as denominator of conversion factor.
   b. Place desired unit as numerator.
   c. Place a "1" in front of the larger unit.
   d. Determine the number of smaller units needed to make "1" of the larger unit.
3. Cancel units. Solve the problem.

Example 1: 55 mm = _____ m

$$55\ \cancel{mm} \times \frac{1\ m}{1000\ \cancel{mm}} = 0.055\ m$$

Example 2: 88 km = _____ m

$$88\ \cancel{km} \times \frac{1000\ m}{1\ \cancel{km}} = 88{,}000\ m$$

Example 3: 7000 cm = _____ hm

$$7000\ \cancel{cm} \times \frac{1\ \cancel{m}}{100\ \cancel{cm}} \times \frac{1\ hm}{100\ \cancel{m}} = 0.7\ hm$$

Example 4: 8 daL = _____ dL

$$8\ \cancel{daL} \times \frac{10\ \cancel{L}}{1\ \cancel{daL}} \times \frac{10\ dL}{1\ \cancel{L}} = 800\ dL$$

The factor label method can be used to solve virtually any problem including changes in units. It is especially useful in making complex conversions dealing with concentrations and derived units.

Convert the following.

1. 35 mL = __________ dL
2. 950 g = __________ kg
3. 275 mm = __________ cm
4. 1,000 L = __________ kL
5. 1,000 mL = __________ L
6. 4,500 mg = __________ g
7. 25 cm = __________ mm
8. 0.005 kg = __________ dag
9. 0.075 m = __________ cm
10. 15 g = __________ mg

FLM / Dimensional analysis PROBLEM SET

1. How many ounces is 45.0 tons?

2. How many liters are in 8.55 cubic miles of ocean water?

3. Water has a density of one gram (1.00g) per cubic centimeter. What is that in pounds per gallon?

4. A road sign posts a 50.0 MPH limit. What is that in KPS?

5. How large in square feet is a 525 square meter house?

6. How many square yards of paint are needed on the ceiling of a room that is 24.0 feet wide, 51.0 feet long and ten feet high?

7. Gasoline has a density of 5.83 pounds per gallon. What volume tank would you need to store three tons of gasoline?

8. The bathtub in the residential suite of the White House had to be enlarged for William Howard Taft. He weighed 370 pounds. The bathtub is four feet wide, six feet long and three feet high. How many gallons of water are needed to fill the WHT Memorial Bathtub? (1 gal = 231 cu.in.)

9. The density of water is 8.3 pounds/gal. What weight of water is needed to fill the WHT Memorial Bathtub?

10. Mercury is 13.6 kg/liter. What mass of mercury is needed to fill the WHT Memorial Bathtub?

11. The velocity of light is 186,000 mi/sec. It takes radio signals 17.6 minutes to go from Earth to Jupiter at their closest approach. How far apart are the two planets at that time?

12. A car takes on 22.7 L of gasoline. How many oz. is that?

13. Osmium metal is 22.0 g/mL. What's the volume of 2.50 kg of it?

14. Concentrated sulfuric acid is 1.84 kg/L. What is the mass of 500 mL of it?

15. A man is 5 ft. 11 inches tall. How many meters is that?

Chemistry: Form WS1.4.1A Name ______________________

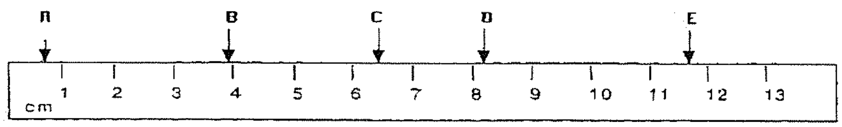
MATTER AND ENERGY Date ______________ Period ____

Measurement: Estimating One Place Beyond

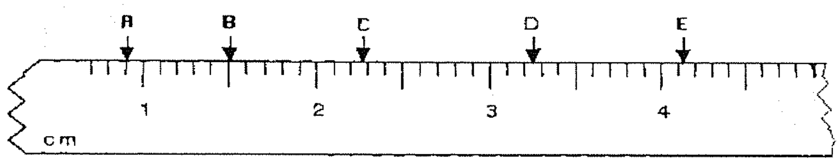
The accuracy of measurement is limited by the tools we use. If a clock doesn't have a second hand, we can estimate the number of seconds past the minute by looking at the minute hand. If a clock doesn't even have a minute hand, we can estimate time in hours and minutes, but certainly not seconds!!

Directions: **Based on the diagrams below, make the best estimate for each of the measurements indicated by the arrows.**

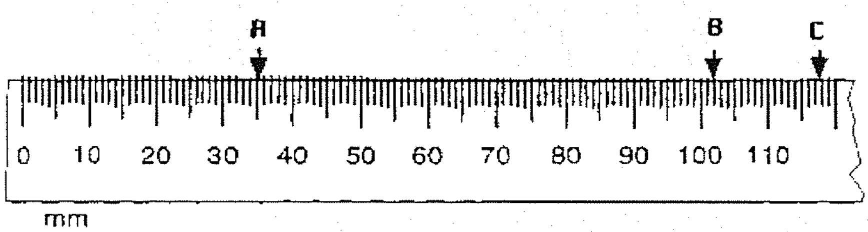
1. Estimate the number of centimeters indicated by each of the arrows below.



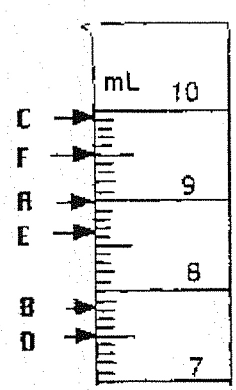
2. Estimate the number of centimeters indicated by each of the arrows below.



3. Estimate the number of millimeters and the number of centimeters indicated by each of the arrows below.



4. Estimate the number of milliliters indicated by each of the arrows below.



1. [a] ______________
   [b] ______________
   [c] ______________
   [d] ______________
   [e] ______________

2. [a] ______________
   [b] ______________
   [c] ______________
   [d] ______________
   [e] ______________

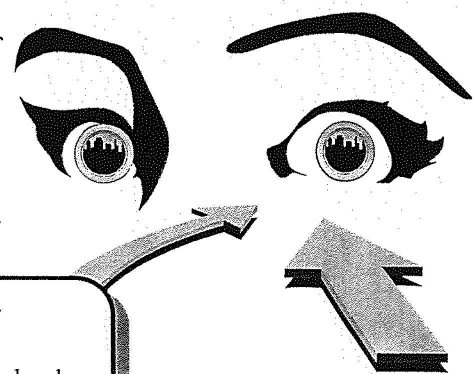
3. [a] ______________ mm
   ______________ cm
   [b] ______________ mm
   ______________ cm
   [c] ______________ mm
   ______________ cm

4. [a] ______________
   [b] ______________
   [c] ______________
   [d] ______________
   [e] ______________
   [f] ______________

Chemistry: Form WS1.4.3A Name ______________________

MATTER Date ______________ Period ____

Calculating Errors of Measurement



If you are trying to get to New York City and you are 5 cm off course, you will still arrive in New York City. If you are trying to take some dirt out of your eye and you are 5 cm off course, you are working on the wrong eye. The size of the error is the same. The size of the error compared to the size of the target is not the same. The actual size of the error – the difference between the observed value and the true value – is known as the **absolute error**. The sign of the absolute error is not important. The size of the error is more important than whether the value is over or under.

★ **Observed value** - value based on laboratory measurements

★ **True value** - most probable value or accepted value based on references

$$\text{Absolute error} = |\text{Observed value} - \text{True value}|$$

The real measure of how far off a value is, is the percentage error. It is the size of the error, the absolute error, compared to the true value.

$$\textit{Percent error} = \frac{|\textit{observed value} - \textit{true value}|}{\textit{true value}} \times 100\%$$

Example: Aluminum has a density of 2.7 g/mL. A student measured the density to be 2.5 g/mL. What is the percentage error?

$$\textit{Percent error} = \left|\frac{2.5\,{}^{g}\!/_{mL} - 2.7\,{}^{g}\!/_{mL}}{2.7\,{}^{g}\!/_{mL}}\right| \times 100\% = \left|\frac{-0.2\,{}^{g}\!/_{mL}}{2.7\,{}^{g}\!/_{mL}}\right| \times 100\% = 7.\overline{407}\,\% = 7\,\%$$

Answer the questions below based on your understanding of errors.

1. The freezing point of water is 273.2 K, but it was measured at 250.1 K. What is the percentage error?

2. The mass of a penny is 2.67 g, but it was measured at 2.55 g. What is the percentage error?

3. The air pressure was 101.3 kPa, but the weatherman said it was 1001.3 kPa. What is the percentage error?

4. The amount of heat released when 1 mole of CO_2 forms is 393.5 kJ, but it was measured at 378.2 kJ. What is the percentage error?

PERCENTAGE ERROR

Name ____________________

Percentage error is a way for scientists to express how far off a laboratory value is from the commonly accepted value.

The formula is:

$$\% \text{ error} = \left| \frac{\text{Accepted Value} - \text{Experimental Value}}{\text{Accepted Value}} \right| \times 100$$

→ absolute value

Determine the percentage error in the following problems.

1. Experimental Value = 1.24 g
   Accepted Value = 1.30 g

   Answer: __________

2. Experimental Value = 1.24×10^{-2} g
   Accepted Value = 9.98×10^{-3} g

   Answer: __________

3. Experimental Value = 252 mL
   Accepted Value = 225 mL

   Answer: __________

4. Experimental Value = 22.2 L
   Accepted Value = 22.4 L

   Answer: __________

5. Experimental Value = 125.2 mg
   Accepted Value = 124.8 mg

   Answer: __________

Density Problems

Name:

1. What is the density of an element if a sample having a mass of 43.2g has a volume of 96.5mL?
2. A sample of gas has a volume of 4L and a mass of 4.922g. What is its density?
3. Mercury has a density of 13.6g/mL. What is the volume of a sample that has a mass of 2242g?
4. If a liquid has a density of .88g/cm^3, what volume of this liquid would have a mass of 54g?
5. What is the mass of 84mL if a liquid if its density is 1.25g/mL?
6. What is the mass of 25mL of oxygen gas if its density is 1.43g/mL?
7. A gas is confined in a rectangular tank 25cm long, 8cm high and 10.4cm wide. If the density of the gas is 19.3g/L, what is the mass of the gas?
8. The density of an acid is 1.85g/mL. What volume of the acid would have a mass of 64g?
9. If 40mL of a liquid with a mass of 44.8g was mixed with 50mL of a liquid having a mass of 48g, what would the density of the resulting liquid be?
10. An object has a mass of 57.7g and occupies a volume of 21.65cm^3. Calculate its density.
11. A sample of a substance whose density is 4.19g/cm^3 occupies .11cm^3. What is the mass of this sample?
12. What is the volume of a 29.6g sample of a metal that is known to have a density of 5.15g/cm^3?
13. If the density of silver is 10.5g/cm^3, what is the mass of a sample of silver that occupies 965 cm^3?
14. A certain gas under given conditions has a density of 1.34×10^{-4}g/cm^3. What volume will 250g of this gas occupy under the same conditions?
15. An object is found to have a mass of 1.934kg and occupies a volume of 542cm^3. Calculate its density in g/cm^3?

Name ______________________ Class______ Date__________

Activity 5-7
Density of Solutions

Density is defined as mass of a sample per unit volume. Density is an intensive property of pure substances and mixtures. Density of water solutions and most other solutions is usually reported as grams of solution per milliliter of solution. Another unit that expresses the same property, mass of solution per unit volume of solution, is kilograms per liter. Density does not describe concentration since it does not relate quantity of solute to either quantity of solvent or quantity of solution. Density is a physical property of a given solution.

Sample Problem 1 An aqueous solution of sodium hydroxide is 26.0% by mass sodium hydroxide. Its density is 1.29 g/mL. What is the mass of sodium hydroxide in 1.00 liter of this solution? What is the corresponding mass of water?

Solution Given: density, percentage by mass, and quantity of solution.
Find: mass of solute and mass of solvent.

$$1.00\text{ L} \times \frac{1000\text{ mL}}{1\text{ L}} \times \frac{1.29\text{ g solution}}{1\text{ mL}} \times \frac{26.0\text{ g NaOH}}{100\text{ g solution}} = 335\text{ g NaOH}$$

$$1.00\text{ L} \times \frac{1000\text{ mL}}{1\text{ L}} \times \frac{1.29\text{ g solution}}{1\text{ mL}} \times \frac{74.0\text{ g H}_2\text{O}}{100\text{ g solution}} = 955\text{ g H}_2\text{O}$$

Sample Problem 2 A sample of aqueous solution has a mass of 712 g and contains 242 g ethanol. Its volume is 0.750 L. What is the mass of water in the sample? What is the percentage by mass ethanol? What is the density of the solution in grams per milliliter?

Solution Given: mass of solution, mass of solute, volume of solution.
Find: (a) mass of solvent, (b) percent by mass solute, (c) density.

(a) 712 g solution − 242 g ethanol = 470 g water

(b) $$\frac{242\text{ g ethanol}}{712\text{ g solution}} \times 100 = 34.0\%\text{ ethanol}$$

(c) $$\frac{712\text{ g solution}}{0.750\text{ L solution}} \times \frac{1\text{ L}}{1000\text{ mL}} = \frac{0.949\text{ g}}{1\text{ mL}}$$

Practice problems

Solve the following problems. In the space below each problem, show a labeled setup. Do any necessary arithmetic on scrap paper. Write your answers in the spaces at the right.

1. A solution of ethanol in water has a density of 0.949 g/mL. It is 34.0% by mass ethanol.
   a. What is the mass of 8.00 liters of this solution? 1. a. __________

   b. What is the mass of water in 8.00 liters of this solution? b. __________

2. A solution of sulfuric acid has a density of 1.61 g/mL. This solution contains 70.0% by mass sulfuric acid.
   a. What volume of this solution contains 2.50 kilograms of sulfuric acid?

2. a. ____________

   b. What mass of this solution contains 2.50 kilograms of sulfuric acid?

b. ____________

3. A solution is prepared by mixing 0.767 kg sulfuric acid with 2.71 kg of water to make 3.00 liters of solution.
   a. What is the density of this solution in grams per milliliter?

3. a. ____________

   b. What is the percentage by mass of sulfuric acid in this solution?

b. ____________

4. An aqueous solution is prepared using 1.00 kg of ethanol. The resulting solution has a density of 0.874 g/mL and a volume of 1.68 L.
   a. What is the mass of the solution?

4. a. ____________

   b. What mass of water is required to prepare this solution?

b. ____________

   c. What is the percentage by mass of ethanol in this solution?

c. ____________

5. An aqueous solution is prepared that has a volume of 1.0 L and is 10% by mass sodium chloride. Its density is 1.07 g/mL. How does each procedure affect the properties listed in the table? Use the symbols I—increases, D—decreases, RTS—remains the same, to complete the table.

Procedure	Masses			Volume of the Solution	Density of the Solution
	Solute	Solvent	Solution		
Addition of NaCl(s)					
Evaporation of H_2O					
Addition of H_2O					