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CONCEPT: SIGNIFICANT FIGURES

Significant figures indicate the level of ______________ involved with measurements and recordings.

□ A number with more significant figures is more __________.

□ Determining the number of significant figures for any given value can be easy depending on how you do it.

Significant Figure Rules (HARD)

1. Non-zero digits **ARE** significant.
2. Any zeros between two significant digits **ARE** significant.
3. Leading zeros are **NOT** significant.
4. A final zero or trailing zeros in the decimal portion **ARE** significant.
5. Trailing zeros in a whole number with the decimal point shown **ARE** significant.
6. Trailing zeros in a whole number with no decimal point are **NOT** significant.
7. For a number in scientific notation: N x 10^x, all digits comprising N ARE significant while following the first 6 rules.

Significant Figure Rules (EASY)

1. **If your number has a decimal point move from _______ to _______.**

□ Start counting once you get to your first non-zero number and keep counting until the end.

0.000250

2. **If your number has NO decimal point move from _______ to _______.**

□ Start counting once you get to your first non-zero number and keep counting until the end.

7055

EXAMPLE 1: How many sig figs does each number contain?

a) 0.0001010 m

b) 10030 s

c) 2.00 x 10^3 L

c) 14 people

● Multiplication/Division: Measurement with least ______________________________ determines final answer.

● Addition/Subtraction: Measurement with least ______________________________ determines final answer.

EXAMPLE 2: Perform the following calculations to the right number of sig figs:

$$\frac{(0.999+1.01)\cdot(15.1-0.815)}{(0.010)(16)+(33.3+7.5)}$$

CONCEPT: ADDITION AND SUBSTRACTION

When you add or subtract values in scientific notation they must have the same exponents. The coefficients add or subtract, but the exponents ___________________________.

$$\begin{array}{r} A \times 10^{x} \\ -\,B \times 10^{x} \\ \hline \end{array} \qquad \begin{array}{r} A \times 10^{x} \\ +\,B \times 10^{x} \\ \hline \end{array}$$

- If the exponents are not the same then we transform the _______________ value so that they do.
- Remember when adding or subtracting values that the final answer must have the ___________________.

EXAMPLE 1: Using the method discussed above, determine the answer to the following question.

$$\begin{array}{r} 8.17 \times 10^{8} \\ +1.25 \times 10^{9} \\ \hline \end{array}$$

EXAMPLE 2: Using the method discussed above, determine the answer to the following question.

$$\begin{array}{r} 9.08 \times 10^{-11} \\ -1.17 \times 10^{-12} \\ -3.35 \times 10^{-13} \\ \hline \end{array}$$

CONCEPT: MULTIPLICATION AND DIVISION

When you multiply values in scientific notation you ___________ the coefficients and ___________ the exponents.

$$(A \times 10^{x}) \cdot (B \times 10^{y}) =$$

When you divide values in scientific notation you ___________ the coefficients and ___________ the exponents.

$$\frac{(A \times 10^{x})}{(B \times 10^{y})} =$$

After multiplying and/or dividing remember that for the coefficient will have the ______________________________.

EXAMPLE 1: Using the method discussed above, determine the answer when the following values are multiplied.

$$(2.134 \times 10^{5}) \cdot (1.6 \times 10^{-3}) \cdot (3.07 \times 10^{6})$$

EXAMPLE 2: Using the methods discussed above, determine the answer for the following mixed operations question.

$$\frac{(7.33 \times 10^{8}) \cdot (9.89 \times 10^{-1})}{(6.12 \times 10^{11})}$$

PRACTICE: MULTIPLICATION AND DIVISION

EXAMPLE: Perform the following calculation to the right number of sig figs:

$$\frac{(9.12\times10^{-5})+(6.33\times10^{-3})}{(1.15\times10^{7})-(3.72\times10^{6})}$$

PRACTICE 1: Perform the following calculation to the right number of sig figs:

$$\left[\frac{(11.422-0.800)+(8.0+1.115)}{0.0720}\right]\cdot1.33\times10^{-5}$$

PRACTICE 2: Compute the following and determine the correct number of significant figures in the answer:

$$\frac{342.60\ (49.37+0.0063+897.10)+9.0287\ (87.001\times10^{2})}{(403.0\times10^{-19})+(-16\times10^{-20})}$$

CONCEPT: LOGARITHM & ANTI-LOGARITHMS

In General Chemistry, learned that pH equaled the negative (–) log of the concentration of hydronium ions and even then we had to take into account the number of significant figures.

When dealing with a number with a decimal point, like the one below, the portion to the left of the decimal point is referred to as the __________________ and the portion to the right of the decimal point is referred to as the __________________ .

12.005

When taking the log of a number:

1) The number of digits in the mantissa of your answer is equal to the number of significant figures in the problem.

EXAMPLE 1: Provide the answer with the correct number of significant figures for each of the following:

a) $\log(1.15 \times 10^{-5})$

b) log(100.)

The **antilogarithm** of a number x is equal to ____________. When taking the antilog of a number:

1) The number of significant figures in the answer is equal to the number of digits in the mantissa of the problem.

EXAMPLE 2: Provide the answer with the correct number of significant figures for each of the following:

a) antilog(– 4.18)

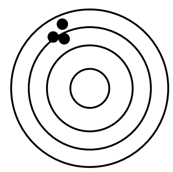
b) $10^{0.0033}$

CONCEPT: PRECISION & ACCURACY

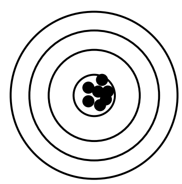
Even though we try to be as correct as possible every measurement or calculation we do in chemistry has some level of uncertainty called ___________________________.

When we investigate the quality of an experimental decision or calculation we take into account two major principles:

The first principle, which deals with the reproducibility of our calculations, is called ________________.



The second principle called ___________________ deals with how close our measured calculation is to the "actual" value.



EXAMPLE 1: A student must measure the weight of a sodium bicarbonate compound and obtains the following measurements: 23.12 g, 23.26 g, 23.08 g and 23.17g. If the true weight of the compound is 18.01 g what can be said about the student's results?

a) They are accurate and precise.
b) They are accurate, but not precise.
c) They are not accurate, but precise.
d) They are neither accurate or precise.

In the above equation it may seem to determine if the results are close to the "true" value by merely looking, but sometimes the determining if your results are precise or not may require more work.

CONCEPT: TYPES OF ERRORS

We've learned that all calculations made have some level of uncertainty or error associated with it.

A ___________________ error, also called ***indeterminate error***, occurs from uncontrollable variables in an experiment. It can occur at any time, in a positive or negative magnitude, can never be corrected and is not reproducible.

A ___________________ error occurs from a problem with the machinery or a design flaw in an experiment. It occurs in the same magnitude, can be corrected and is reproducible.

EXAMPLE 1: A uncalibrated pipet is used in the titration of 25 mL of 0.250 M $KMnO_4$ with 50.0 mL of HNO_3. If the pipet delivers 23.120 $\pm$ 0.02 mL what can be said about the possible error(s) observed?

EXAMPLE 2: State whether the errors are random or systematic for each of the following:

a) The analytical measuring pipet in the lab consistently delivers 25.0 $\pm$ 0.03 mL.

b) I weigh an analyte sample 4 times and obtain the following numbers: 1.110, 1.392, 1.040 and 1.850.

CONCEPT: TYPES OF UNCERTAINTY

Absolute Uncertainty is the ± value associated with any numerical calculation. If a student delivers 25.00 ± 0.02 mL of water to a mixture the ± value of 0.02 mL would be the uncertainty.

Relative Uncertainty is the absolute uncertainty divided by the associated measurement.

$$\text{Relative Uncertainty} = \frac{\text{Absolute Uncertainty}}{\text{Magnitude of Measurement}} = \frac{0.02\ \text{mL}}{25.00\ \text{mL}} = 0.001$$

The *Percent Relative Uncertainty* would be the relative uncertainty multiplied by 100.

$$\text{Percent Relative Uncertainty} = 0.001 \times 100 = 0.1\%$$

EXAMPLE: Calculate the relative and percent relative uncertainty from the given problem.

a) 3.25 (±0.03)

PRACTICE: Calculate the absolute uncertainty from the given problem.

a) 6.77 (±5.6%)

CONCEPT: PROPAGATION OF UNCERTAINTY FROM RANDOM ERROR

With addition and subtraction or multiplication and division we will use certain rules for propagation of our answer.

- **The Real Rule:** The first digit of the absolute uncertainty is the last significant digit in the answer.

Addition & Subtraction

For addition and subtraction, the uncertainty in our final answer is determined from each individual absolute uncertainty.

1

$$y = x_1 + x_2 \qquad e_y = \sqrt{e_{x1}^2 + e_{x2}^2}$$

$$y = x_1 - x_2 \qquad e_y = \sqrt{e_{x1}^2 + e_{x2}^2}$$

2

$$\begin{array}{r} 5.22\ (\pm 0.04) \\ +3.80\ (\pm 0.03) \\ -2.17\ (\pm 0.03) \\ \hline 6.85\ (\pm e_4) \end{array}$$

3

$$e_4 = \sqrt{(0.04)^2 + (0.03)^2 + (0.03)^2} = \sqrt{0.0034} = 0.05_8$$

4

$$6.85(\pm 0.06)$$

Multiplication & Division:

For multiplication and division, we must first convert the absolute uncertainties into percent relative uncertainties.

1

$$y = x_1 \cdot x_2 \qquad \%e_y = \sqrt{\%e_{x1}^2 + \%e_{x2}^2}$$

$$y = \frac{x_1}{x_2} \qquad \%e_y = \sqrt{\%e_{x1}^2 + \%e_{x2}^2}$$

2

$$\frac{3.68\ (\pm 0.05) \times 1.15\ (\pm 0.06)}{0.92\ (\pm 0.06)} = 4.6 \pm e_4$$

3

$$\frac{0.05}{3.68} \times 100 = 1._4\% \qquad \frac{0.06}{1.15} \times 100 = 5._2\% \qquad \frac{0.06}{0.92} \times 100 = 6._5\%$$

4

$$\%e_4 = \sqrt{(1._4\%)^2 + (5._2\%)^2 + (6._5\%)^2} = \sqrt{71.25\%} = 8._4\%$$

5

$$\frac{8._4\%}{100} \times 4.6 = 0.3_9$$

6

$$4.6\ (\pm 8\%)$$

$$4.6\ (\pm 0.4)$$

PRACTICE: PROPAGATION OF UNCERTAINTY CALCULATIONS 1

EXAMPLE: Determine the absolute and relative uncertainty to the following addition problem.

1.511 (± 0.02) + 2.53 (± 0.01) + 0.987 (± 0.01)

PRACTICE: Determine the absolute and relative uncertainty to the following addition and subtraction problem.

8.88 (± 0.03) - 3.29 (± 0.10) + 6.43 (± 0.001)

PRACTICE: PROPAGATION OF UNCERTAINTY CALCULATIONS 2

EXAMPLE: Determine the absolute and relative uncertainty to the following multiplication and division problem.

$$\frac{7.73\,(\pm\,0.02) \times 8.921\,(\pm\,0.02)}{9.17\,(\pm\,0.03)}$$

PRACTICE: Determine the absolute and relative uncertainty to the following multiplication and division problem.

$$\frac{1.12\,(\pm\,0.01) \times 0.546\,(\pm\,0.01)}{3.12\,(\pm\,0.02) \times 1.12\,(0.03)}$$

PRACTICE: PROPAGATION OF UNCERTAINTY CALCULATIONS 3

EXAMPLE: Two students wish to prepare a stock solution for their lab experiment. Student A uses an un-calibrated pipet that delivers 50.00 ($\pm$ 0.02) mL to deliver 200 mL to a container. Student B uses a calibrated pipet that delivers 40.00 ($\pm$ 0.01) mL to deliver 200 mL to a container.

a) Calculate the absolute uncertainty in each of their deliveries.

PRACTICE: Based on the previous example calculate the molarity value for each student if they dissolve 0.300 ($\pm$0.03) moles of analyte.

PRACTICE: PROPAGATION OF UNCERTAINTY CALCULATIONS 4

EXAMPLE 1: A "Class A" 50-mL buret is certified by the manufacturer to deliver volumes within a "tolerance" (i.e. uncertainty) of ± 0.05 mL. The smallest graduations on the buret are 0.1 mL. You use the buret to titrate a solution, adding 5 successive volumes to the solution. The following volumes were added:

Addition	Volume (mL)
1	6.73
2	8.92
3	7.52
4	2.48
5	5.15

What is the total volume added, and **what is the uncertainty** associated with this final volume?

EXAMPLE 2: A Class A 250 mL volumetric flask has an uncertainty of ± 0.15 mL, and a 50 mL volumetric pipet has an uncertainty of ± 0.05 mL. If I fill a 250 mL volumetric flask to the line and remove four 50 mL aliquots with my volumetric pipet, I should have 50 mL of solution remaining in the flask. What is the absolute and relative uncertainty in the remaining volume?

PRACTICE: PROPAGATION OF UNCERTAINTY CALCULATIONS 5

PRACTICE 1: I am making a 0.1 M KCl (molar mass 74.551) solution for an experiment. To measure the mass of the KCl, I will use an analytical balance that is only accurate to ± 0.01 g. I place a piece of paper on the balance and set the tare to read 0.00. I then put the KCl on the balance until it reads 6.79 g. What is the uncertainty in this mass?

PRACTICE 2: The volume of the solution I am making is 2.5 L. To measure this volume I will use a large graduated cylinder that can measure volume to ± 10 mL. What is the absolute uncertainty in my concentration?