Rules for Reporting Significant Figures

1. Nonzero digits always count as significant figures

2. Zeros are what mix people up. There are three situations in which they can occur.

leading zeros precede all nonzero digits and are **never** significant (i.e., 0.000182 has three sign. figs.)

captive zeros are between nonzero digits and are **always** significant (i.e., 1008.02 has six sign. figs.)

trailing zeros are significant **only if** a number contains a decimal point (i.e., 1200 has two sign. figs.; 1200.00 has six sign. figs., 1.200×10^3 has four sign. figs.)

*** Note here the advantage of using exponential (i.e., scientific) notation to clear up these ambiguities!**

3. Exact numbers have no bearing on the number of significant figures in a calculated result. Examples of these are the following;

conversion factors such as 1 L = 1000 mL
numbers reflecting an exact count such as 8 stones or 16 people
stoichiometry in chemical reactions involves exact numbers

Operations

Rules for multiplication/division

The answer to contain the same **number of sign. figs.** as the least precise measurement used in the calculation.

72.5674	*six sign. figs*	
x 3.34	*three sign. figs (limiting term)*	
242.3751160		*initial answer (must be rounded off to three sign. figs.)*
		Final Answer = 2.42×10^2

Rules for addition/subtraction

The answer to contain the same **number of decimal places** as the least precise measurement used in the calculation.

456.367963	
- 452.1	*least number of decimal places (limiting term)*
4.267963	*initial answer (must be rounded off to one decimal place)*
	Final Answer = 4.3 !

Rules for logarithms

In logarithmic values, only those numbers to the right of the decimal place count as significant. For example,

pH = 10.26 has only two significant figures and corresponds to a $[H^+] = 5.5 \times 10^{-11}$ **M**
$pK_a = 4.730$ has three significant figures and corresponds to $K_a = 1.86 \times 10^{-5}$

What is the pH if the concentration of H^+ is measured to be 1.25×10^{-6} **M** ?

What is the $[OH^-]$ if the pH has been determined to be 9.32 ?

Rules for Propagation of Uncertainty from Random Error

Addition and Subtraction - the squares of the absolute errors are additive (i.e., add the variances)

$y = x_1 + x_2$

$\rightarrow e_y = [(e_{x1})^2 + (e_{x2})^2]^{1/2}$

where e_y is the absolute error in y, and e_{x1} is the absolute error in x_1

Multiplication and Division - the squares of the relative errors are additive

$y = x_1 * x_2$

$\rightarrow e_y/y = [(e_{x1}/x_1)^2 + (e_{x2}/x_2)^2]^{1/2}$

where e_y/y is the relative error in y, and (e_{x1}/x_1) is the relative error in x_1

Exponents and Logarithms

$y = \log x$

$\rightarrow$ absolute error in $y = e_y = 0.43\ (e_x/x)$

$y = 10^x$

$\rightarrow$ relative error in $y = e_y/y = 2.3\ (e_x)$

where e_x = absolute error in x
and e_x/x = relative error in x

E.g. If the pH of a lake sample is measured to be 7.88 ± 0.02, what is the $[H^+]$ and the associated uncertainty?