IB Chemistry SL Unit 0 Practice

Significant Digits

1) The measurement, 206 cm, has how many significant (measured) digits? 3
2) The measurement, 206.0 °C, has how many significant digits? 4
3) The measurement, 0.00206 g, has how many significant digits? 3
4) The measurement, 0.0020600 mole, has how many significant digits? 5
5) The measurement, 2.060×10^{-3} coulombs, has how many significant digits? 4
6) The measurement, 20600 molecules, has how many significant digits? 3
7) Add the following three numbers and report your answer using significant figures:
   2.5 cm + 0.50 cm + 0.055 cm = ? 3.055 → 3.1 cm
8) Subtract the following numbers and report your answer using significant figures:
   416 g – 210. g = ? 206 g
9) Multiply the following three numbers and report your answer to the correct number of significant figures:
   0.020 cm x 50. cm x 11.1 cm = ? 11.1 → 11 cm
10) Divide the following three numbers and report your answer to the correct number of significant figures:
   0.530 g / 0.1010 mL = ? 5.25 mL

Propagation of Uncertainties

Adding and subtracting values: *Rule: Add uncertainties!*

1) 1.00 ± 0.05 g of Copper is added to 10.00 g ± 0.05 g of Iron. What is the mass of the two, together?
   11.00 ± 0.10 g ← keep 2 dig so we don't loose sig figs
2) The starting volume on a buret is 40.00 ± 0.02 mL. The final volume on the buret is 32.60 ± 0.02 mL. What volume of liquid was removed from the buret? 7.40 ± 0.04 mL
3) The weighing boat weighs 2.45 ± 0.01 g. The total weight, after adding Calcium Carbonate to the weighing boat is 4.62 ± 0.01 g. What is the mass of the calcium carbonate? 2.17 ± 0.02 g

Multiplying and dividing by a constant value (constant value means the value has no uncertainty). *Rule: Multiply or Divide the uncertainty by the constant value (in other words, do the same thing to the uncertainty that you did to the value!)*

1) Unit conversion
   a) How many milligrams is 0.000120 ± 0.000005 g?
   b) How many milligrams is $1.20 \times 10^{-4} \pm 5 \times 10^{-6}$ g?
   c) How many milligrams is $(1.20 \pm 0.05) \times 10^{-4}$ g?
   d) All of the above are the same problem, which is the best way to write the question?

see paper

2) Mole conversions
   a) How many moles in 12.01 ± 0.02 g of Aluminum?
   b) How many atoms in 1.45 ± 0.04 moles of Aluminum?
   c) How many atoms in 3.24 ± 0.01 g of Aluminum?

Multiplying and dividing when both values have uncertainty *Rule: 1) Find the fractional uncertainty for each value given in the problem, 2) Add the fractional uncertainties together 3) Find the absolute uncertainty of the desired quantity by multiplying the result of step 2) times the desired quantity.*

1) Molarity: Calculate the molarity when 2.0 ± 0.1 moles of NaCl is dissolved in 100. ± 1 mL of water.
2) Density: Calculate the density of an object with mass 25.0 ± 0.2 g and volume 50.0 ± 0.5 mL
3) Density: Calculate the density of a rectangular object with mass of 32.0 ± 0.2 g, length of 1.25 ± 0.05 cm, width of 2.25 ± 0.05 cm and height of 3.35 ± 0.05 cm.

Topic 11 Practice (Unit 0)

$\times/\div$ by a constant (conversions)

1) a) $0.000120 \pm 0.000005\,g \left(\frac{1000\,mg}{1\,g}\right) = 0.120 \pm 0.005\,mg$

b) $1.20\times10^{-4} \pm 5\times10^{-6}\,g \left(\frac{1000\,mg}{1\,g}\right) = 0.120 \pm 0.005\,mg$

c) $(1.20 \pm 0.05)\times10^{-4}\,g \left(\frac{1000\,mg}{1\,g}\right) = 0.120 \pm 0.005\,mg$

d) discuss.

2) a) $12.01 \pm 0.02\,g\,Al \left(\frac{1\,mol}{26.98\,g}\right) = 0.4451 \pm 0.0007\,mol\,Al$

b) $1.45 \pm 0.04\,mol\,Al \left(\frac{6.02\times10^{23}\,atom\,Al}{1\,mol\,Al}\right) = 8.73\times10^{23} \pm 0.24\times10^{23}\,atoms$

c) $3.24 \pm 0.01\,g\,Al \left(\frac{1\,mol}{26.98\,g\,Al}\right)\left(\frac{6.02\times10^{23}\,atom}{1\,mol}\right) = 7.22\times10^{22} \pm 0.02\times10^{22}\,atoms$

$\times/\div$ w/ both uncertain values

1) $\left(\frac{2.0 \pm 0.1\,mol\,NaCl}{100. \pm 1\,mL}\right)\left(\frac{1000\,mL}{1\,L}\right) =$ need % uncertainty

$\left(\frac{2.0 \pm 5\%\,mol\,NaCl}{100. \pm 1\%\,mL}\right)\left(\frac{1000\,mL}{1\,L}\right) = 20. \pm 6\%\,M\,NaCl$

$\downarrow$

$20. \pm 1.2\,M \rightarrow$

$\downarrow$

$\boxed{20. \pm 1\,M\,NaCl}$

$$\left(\frac{x}{\text{value}}\right)100 = \%$$

2) $\frac{25.0 \pm 0.2\,g}{50.0 \pm 0.5\,mL}$ = need % uncertainty

$$\frac{25.0 \pm 0.8\%\,g}{50.0 \pm 0.4\%\,mL} = 0.500 \pm 1.2\%\ g/mL$$

$\downarrow$

$$\boxed{0.500 \pm 0.006\ g/mL}$$

3) volume = $l \times w \times h$, switch to % uncertainty

$$(1.25 \pm 4\%\,cm)(2.25 \pm 2\%\,cm)(3.35 \pm 1\%\,cm) = 9.42 \pm 7\%\,cm^3$$

now density: $\frac{32.0 \pm 0.6\%\,g}{9.42 \pm 7\%\,cm^3} = 3.40 \pm 7.6\%\ g/cm^3$

$\downarrow$

$$\boxed{3.40 \pm 0.26\ g/cm^3}$$

∴ when determining uncertainty in calculations keek as many digits as you need to keep all sig figs in the numerical answer.