

Chapter 2

1. (a) 3 significant figures; (b) 3 significant figures; (c) 5 significant figures; (d) 3 significant figures; (e) 4 significant figures; (f) 3 significant figures

For (d) and for (e), the zero in the tenths place is not a significant digit as its presence is needed only to indicate the position of the decimal point. If you write these using scientific notation, they are 9.03×10^{-2} and 9.030×10^{-2} , with three and four significant figures respectively.

2. (a) 0.894; (b) 0.893; (c) 0.894; (d) 0.900; (e) 0.0891
 3. (a) 12.01; (b) 16.0; (c) $6.022 \times 10^{23} \text{ mol}^{-1}$; (d) $9.65 \times 10^4 \text{ C/mol}$
 4. a. $4.591 + 0.2309 + 67.1 = 71.9219 \approx 71.9$
 b. $313 - 273.15 = 39.85 \approx 39.8$

Note that for (b) we retain an extra significant figure beyond that suggested by our simple rules so that the uncertainty in the final answer (1 part out of 398) is approximately the same as the most uncertain of our two measurements (1 part out of 313). Reporting the answer as 40, or 4.0×10^1 , as suggested by our simple rules, gives an uncertainty in the final result of 1 part out of 40, which is substantially worse than either of our two measurements.

- c. $712 \times 8.6 = 6123.2 \approx 6.1 \times 10^3$
 d. $1.43/0.026 = 55.00 \approx 55$
 e. $(8.314 \times 298)/96485 = 0.0256783 \approx 2.57 \times 10^{-2}$
 f. $\log(6.53 \times 10^{-5}) = -4.18509 = -4.185$

Note that when we take the logarithm of a number, any digits before the decimal point provide information on the original number's power of 10; thus, the 4 in -4.185 is not counted as a significant digit.

g. $10^{-7.14} = 7.244 \times 10^{-8} \approx 7.2 \times 10^{-8}$

Note that we take the antilog of a number, the digits before the decimal point provide information on the power of 10 for the resulting answer; thus, the 7 in -7.14 is not counted as a significant digit.

h. $(6.51 \times 10^{-5}) \times (8.14 \times 10^{-9}) = 5.29914 \times 10^{-13} \approx 5.30 \times 10^{-13}$

5. To find the %w/w Ni, we first subtract the mass of Co from point B from the combined mass of Ni and Co from point B, and then divide by the mass of sample; thus

$$\frac{(0.2306 - 0.0813) \text{ g Ni}}{12.1374 \text{ g sample}} \times 100 =$$

$$\frac{0.1493 \text{ g Ni}}{12.1374 \text{ g sample}} = 1.230 \text{ \%w/w Ni}$$

In problem 4 we use a bold red font to help us keep track of significant figures. For example, in (a) we mark the last significant digit common to the numbers we are adding together, and in (e), where we are multiplying and dividing, we identify the number with the smallest number of significant digits.

For problems in this chapter, all formula weights are reported to the number of significant figures allowed by the atomic weights in Appendix 18. As a compound's formula weight rarely limits the uncertainty in a calculation, in later chapters usually we will round formula weights to a smaller number of significant figures, chosen such that it does not limit the calculation's uncertainty.

For problem 7 we include an extra significant figure in each of the calculation's first two steps to avoid the possibility of introducing a small error in the final calculation as a result of rounding. If we need to report the result for an intermediate calculation, then we round that result appropriately; thus, we need to isolate 3.61×10^{-3} mol of BaCl_2 .

6. Using the atomic weights from Appendix 18, we find that the formula weight for $\text{Ni}(\text{C}_4\text{H}_{14}\text{N}_4\text{O}_4)_2$ is

$$(58.693) + (8 \times 12.011) + (14 \times 1.008) + (4 \times 14.007) + (4 \times 15.999) = 288.917 \text{ g/mole}$$

7. First we convert the mass of Cl^- to moles of Cl^-

$$256 \text{ mg Cl}^- \times \frac{1 \text{ g Cl}^-}{1000 \text{ mg Cl}^-} \times \frac{1 \text{ mol Cl}^-}{35.45 \text{ g Cl}^-} = 7.221 \times 10^{-3} \text{ mol Cl}^-$$

and then the moles of Cl^- to the moles of BaCl_2

$$7.221 \times 10^{-3} \text{ mol Cl}^- \times \frac{1 \text{ mol BaCl}_2}{2 \text{ mol Cl}^-} = 3.610 \times 10^{-3} \text{ mol BaCl}_2$$

and finally the moles of BaCl_2 to the volume of our BaCl_2 solution

$$3.610 \times 10^{-3} \text{ mol BaCl}_2 \times \frac{1 \text{ L}}{0.217 \text{ mol BaCl}_2} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 16.6 \text{ mL}$$

8. We can express a part per million in several ways—this is why some organizations recommend against using the abbreviation ppm—but here we must assume that the density of the solution is 1.00 g/mL and that ppm means mg/L or $\mu\text{g/mL}$. As molarity is expressed as mol/L, we will use mg/L as our starting point; thus

$$\frac{0.28 \text{ mg Pb}}{\text{L}} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} = 1.4 \times 10^{-6} \text{ M Pb}$$

9. (a) The molarity of 37.0% w/w HCl is

$$\frac{37.0 \text{ g HCl}}{1.00 \times 10^2 \text{ g solution}} \times \frac{1.18 \text{ g solution}}{\text{ml solution}} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} = 12.0 \text{ M HCl}$$

(b) To calculate the mass and volume of solution we begin with the molarity calculated in part (a). To avoid any errors due to rounding the molarity down to three significant, we will return one additional significant figure, taking the molarity as 11.97 M.

$$0.315 \text{ mol HCl} \times \frac{1 \text{ L}}{11.97 \text{ mol HCl}} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{1.18 \text{ g solution}}{\text{mL}} = 31.1 \text{ g}$$

$$0.315 \text{ mol HCl} \times \frac{1 \text{ L}}{11.97 \text{ mol HCl}} \times \frac{1000 \text{ mL}}{\text{L}} = 26.3 \text{ mL}$$

10. A volume of 1.0×10^3 mL is equivalent to 1.0 L; thus

$$1.0 \text{ L} \times \frac{0.036 \text{ mol NH}_3}{\text{L}} \times \frac{17.031 \text{ g NH}_3}{\text{mol NH}_3} \times \frac{1.00 \times 10^2 \text{ g solution}}{28.0 \text{ g NH}_3} \times \frac{1.00 \text{ mL}}{0.899 \text{ g solution}} = 2.4 \text{ mL}$$

11. As we have information about the solution's volume and no information about its density, we will assume that ppm and ppb are expressed as a mass of analyte per unit volume; thus,

$$\frac{45.1 \mu\text{g}}{250.0 \text{ ml}} \times \frac{1 \text{ g}}{1 \times 10^6 \mu\text{g}} \times 100 = 1.80 \times 10^{-5} \% \text{ w/w}$$

$$\frac{45.1 \mu\text{g}}{250.0 \text{ ml}} = 0.180 \text{ ppm}$$

$$\frac{45.1 \mu\text{g}}{250.0 \text{ ml}} \times \frac{1000 \text{ mL}}{\text{L}} = 1.80 \times 10^2 \text{ ppb}$$

12. To obtain a total concentration of 1.6 ppm F^- we must add sufficient NaF such that we increase the concentration of F^- by 1.4 ppm; thus

$$1 \text{ gal} \times \frac{3.785 \text{ L}}{\text{gal}} \times \frac{1.4 \text{ mg F}^-}{\text{L}} \times \frac{41.988 \text{ mg NaF}}{18.998 \text{ mg F}^-} = 12 \text{ mg NaF}$$

Note that "per gallon" implies that 1 gal is an exact number that does not limit the number of significant figures in our final answer.

13. $\text{pH} = -\log[\text{H}^+] = -\log(6.92 \times 10^{-6}) = 5.160$
 $[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-8.923} = 1.19 \times 10^{-9} \text{ M}$

14. When using a 25-mL graduated cylinder to measure 15 mL, the absolute uncertainty is $\pm 1\%$ of 25 mL, or ± 0.25 mL and the relative uncertainty is

$$\frac{\pm 0.25 \text{ mL}}{15 \text{ mL}} \times 100 = \pm 1.7\%$$

When using a 50-mL graduated cylinder to measure 15 mL, the absolute uncertainty is $\pm 1\%$ of 50 mL, or ± 0.50 mL and the relative uncertainty is

$$\frac{\pm 0.50 \text{ mL}}{15 \text{ mL}} \times 100 = \pm 3.3\%$$

15. First, we calculate the moles of $\text{K}_2\text{Cr}_2\text{O}_7$

$$9.67 \text{ g K}_2\text{Cr}_2\text{O}_7 \times \frac{1 \text{ mol K}_2\text{Cr}_2\text{O}_7}{294.181 \text{ g K}_2\text{Cr}_2\text{O}_7} = 0.03287 \text{ mol K}_2\text{Cr}_2\text{O}_7$$

and then we calculate the solution's molarity

$$\frac{0.03287 \text{ mol K}_2\text{Cr}_2\text{O}_7}{0.1000 \text{ L}} = 0.329 \text{ M K}_2\text{Cr}_2\text{O}_7$$

16. Given that the uncertainty in the volume and in the concentration is 1% ($1.0 \text{ L} \pm 0.1 \text{ L}$ or $0.10 \text{ M} \pm 0.01 \text{ M}$), we can prepare these

Given the uncertainty in the volume and in the concentration, there is no advantage to taking the extra time needed to measure the solid's mass to three or four decimal places, to quantitatively transfer the solid to a volumetric flask, and dilute to volume.

solution by weighing out the appropriate amount of solid to two significant figures, place it in a 1-L beaker or bottle, and dissolve in 1000 mL of water. What remains is to calculate the amount of reagent for each solution; thus: (a) for KCl we have

$$1.0 \text{ L} \times \frac{0.10 \text{ mol K}^+}{\text{L}} \times \frac{1 \text{ mol KCl}}{\text{mol K}^+} \times \frac{74.55 \text{ g KCl}}{\text{mol KCl}} = 7.5 \text{ g KCl}$$

$$1.0 \text{ L} \times \frac{1.0 \times 10^2 \text{ mg K}^+}{\text{L}} \times \frac{1 \text{ g K}^+}{1000 \text{ mg K}^+} \times \frac{74.55 \text{ g KCl}}{39.098 \text{ g K}^+} = 0.19 \text{ g KCl}$$

$$1.0 \text{ L} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{1.0 \text{ g K}^+}{1.0 \times 10^2 \text{ mL}} \times \frac{74.55 \text{ g KCl}}{39.098 \text{ g K}^+} = 19 \text{ g KCl}$$

and (b) for K_2SO_4 we have

$$1.0 \text{ L} \times \frac{0.10 \text{ mol K}^+}{\text{L}} \times \frac{1 \text{ mol K}_2\text{SO}_4}{2 \text{ mol K}^+} \times \frac{174.25 \text{ g K}_2\text{SO}_4}{\text{mol K}_2\text{SO}_4} = 8.7 \text{ g K}_2\text{SO}_4$$

$$1.0 \text{ L} \times \frac{1.0 \times 10^2 \text{ mg K}^+}{\text{L}} \times \frac{1 \text{ g K}^+}{1000 \text{ mg K}^+} \times \frac{174.25 \text{ g K}_2\text{SO}_4}{78.196 \text{ g K}^+} = 0.22 \text{ g K}_2\text{SO}_4$$

$$1.0 \text{ L} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{1.0 \text{ g K}^+}{1.0 \times 10^2 \text{ mL}} \times \frac{174.25 \text{ g K}_2\text{SO}_4}{78.196 \text{ g K}^+} = 22 \text{ g K}_2\text{SO}_4$$

and (c) for $\text{K}_3\text{Fe}(\text{CN})_6$ we have

$$1.0 \text{ L} \times \frac{0.10 \text{ mol K}^+}{\text{L}} \times \frac{1 \text{ mol K}_3\text{Fe}(\text{CN})_6}{3 \text{ mol K}^+} \times \frac{329.247 \text{ g K}_3\text{Fe}(\text{CN})_6}{\text{mol K}_3\text{Fe}(\text{CN})_6} = 11 \text{ g K}_3\text{Fe}(\text{CN})_6$$

$$1.0 \text{ L} \times \frac{1.0 \times 10^2 \text{ mg K}^+}{\text{L}} \times \frac{1 \text{ g K}^+}{1000 \text{ mg K}^+} \times \frac{329.247 \text{ g K}_3\text{Fe}(\text{CN})_6}{117.294 \text{ g K}^+} = 0.28 \text{ g K}_3\text{Fe}(\text{CN})_6$$

$$1.0 \text{ L} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{1.0 \text{ g K}^+}{1.0 \times 10^2 \text{ mL}} \times \frac{329.247 \text{ g K}_3\text{Fe}(\text{CN})_6}{117.294 \text{ g K}^+} = 28 \text{ g K}_3\text{Fe}(\text{CN})_6$$

17. For a serial dilution, we need the concentration of solution A to calculate the concentration of solution B, and the concentration of solution B to calculate the concentration of solution A; thus

$$\text{Solution A: } 0.100 \text{ M} \times \frac{10.00 \text{ mL}}{250.0 \text{ mL}} = 4.00 \times 10^{-3} \text{ M}$$

$$\text{Solution B: } 4.00 \times 10^{-3} \text{ M} \times \frac{25.00 \text{ mL}}{100.0 \text{ mL}} = 1.00 \times 10^{-3} \text{ M}$$

$$\text{Solution C: } 1.00 \times 10^{-3} \text{ M} \times \frac{20.00 \text{ mL}}{500.0 \text{ mL}} = 4.00 \times 10^{-5} \text{ M}$$

18. When we dissolve 1.917 g NaCl in 50 mL of water measured using a graduated cylinder, the reported concentration of NaCl is limited to just two significant figures because the uncertainty for the graduated cylinder is approximately 1%, or 1 part in 50; thus

$$\frac{1.917 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}}}{0.050 \text{ L}} = 0.66 \text{ M NaCl}$$

When we quantitatively transfer this solution to a 250.0 mL volumetric flask and dilute to volume, we can report the concentration of NaCl to four significant figures because the uncertainty in the volumetric flask is ± 0.01 mL, or 1 part in 2500; thus

$$\frac{1.917 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}}}{0.2500 \text{ L}} = 0.1312 \text{ M NaCl}$$

Note that the second calculation does not begin with the concentration from the previous calculation, as we did in problem 17 for a serial dilution. A quantitative transfer is not a serial dilution; instead, all 1.917 g of NaCl added to the 50-mL beaker is transferred to the 250.0 mL volumetric flask, so we begin our calculation with this mass of NaCl.

19. First, we calculate the moles of NO_3^- in 50.0 mL of 0.050 M KNO_3 and in 40.0 mL of 0.075 M KNO_3 .

$$0.0500 \text{ L} \times \frac{0.050 \text{ mol NO}_3^-}{\text{L}} = 2.50 \times 10^{-3} \text{ mol NO}_3^-$$

$$0.0400 \text{ L} \times \frac{0.075 \text{ mol NO}_3^-}{\text{L}} = 3.00 \times 10^{-3} \text{ mol NO}_3^-$$

Next, we add together these results to obtain the total moles of NO_3^- in the combined solutions, and then divide by the total volume to find the concentration of NO_3^- and pNO_3^- .

$$\frac{2.50 \times 10^{-3} \text{ mol NO}_3^- + 3.00 \times 10^{-3} \text{ mol NO}_3^-}{0.0500 \text{ L} + 0.0400 \text{ L}} = 0.061 \text{ M NO}_3^-$$

$$\text{pNO}_3^- = -\log(\text{NO}_3^-) = -\log(0.061) = 1.21$$

20. First, we calculate the moles of Cl^- in 25.0 mL of 0.025 M NaCl and in 35.0 mL of 0.050 M BaCl_2 .

As we are interested only in the concentration of NaCl in our final solution, there is no particular reason for us to complete the intermediate calculation; we did so here simply to make this point: The uncertainty in a calculated result is determined by the measurements that contribute to the calculation only, and is not affected by other measurements that we happen to make. What matters in this case is that 1.917 g of NaCl are dissolved in 250.0 mL of water. If we fail to complete a quantitative transfer, then our calculated concentration is in error, but this is an error in the accuracy of our work, not an uncertainty in the inherent precision of the balance or volumetric pipet. We will have more to say about accuracy and precision in Chapter 4.

Here, again, we keep an extra significant figure through the intermediate steps of our calculation.

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$$0.0250 \text{ L} \times \frac{0.025 \text{ mol NaCl}}{\text{L}} \times \frac{1 \text{ mol Cl}^-}{\text{mol NaCl}} = 6.25 \times 10^{-4} \text{ mol Cl}^-$$

$$0.0350 \text{ L} \times \frac{0.050 \text{ mol BaCl}_2}{\text{L}} \times \frac{2 \text{ mol Cl}^-}{\text{mol BaCl}_2} = 3.50 \times 10^{-3} \text{ mol Cl}^-$$

Next, we add together these results to obtain the total moles of Cl^- in the combined solutions, and then divide by the total volume to find the concentration of Cl^- and pCl .

$$\frac{6.25 \times 10^{-4} \text{ mol Cl}^- + 3.50 \times 10^{-3} \text{ mol Cl}^-}{0.0250 \text{ L} + 0.0350 \text{ L}} = 0.069 \text{ M Cl}^-$$

$$\text{pCl}^- = -\log(\text{Cl}^-) = -\log(0.069) = 1.16$$

21. The concentration is

$$0.0844 \text{ M ethanol} \times \frac{0.500 \text{ L}}{0.0050 \text{ L}} = 8.44 \text{ M ethanol}$$