

Making Connections

11. Forensic chemists analyze tissue for many things. A typical test would be for the presence and amount of arsenic. Quantity must be measured precisely to determine if a substance is present in a natural amount, or in an amount much greater, which could perhaps indicate foul play.

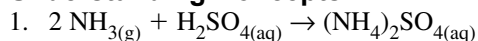
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7.6 QUANTITATIVE ANALYSIS

PRACTICE

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Understanding Concepts



24.4 mL 50.0 mL

2.20 mol/L C

$$n_{\text{NH}_3} = 0.0244 \text{ L} \times \frac{2.20 \text{ mol}}{1 \text{ L}} = 0.0537 \text{ mol}$$

$$n_{\text{H}_2\text{SO}_4} = 0.0537 \text{ mol} \times \frac{1}{2} = 0.0268 \text{ mol}$$

$$n_{\text{H}_2\text{SO}_4} = 0.0268 \text{ mol}$$

$$C_{\text{H}_2\text{SO}_4} = \frac{0.0268 \text{ mol}}{0.0500 \text{ L}}$$

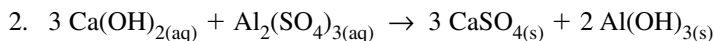
$$C_{\text{H}_2\text{SO}_4} = 0.537 \text{ mol/L}$$

or

$$C_{\text{H}_2\text{SO}_4} = 0.0244 \text{ L NH}_3 \times \frac{2.20 \text{ mol NH}_3}{1 \text{ L NH}_3} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NH}_3} \times \frac{1}{0.0500 \text{ L}}$$

$$C_{\text{H}_2\text{SO}_4} = 0.537 \text{ mol/L}$$

The concentration of sulfuric acid at this stage is 0.537 mol/L.



v 25.0 mL

0.0250 mol/L 0.125 mol/L

$$n_{\text{Al}_2(\text{SO}_4)_3} = 25.0 \text{ mL} \times \frac{0.125 \text{ mol}}{1 \text{ L}} = 3.13 \text{ mmol}$$

$$n_{\text{Ca}(\text{OH})_2} = 3.13 \text{ mmol} \times \frac{3}{1} = 9.38 \text{ mmol}$$

$$v_{\text{Ca}(\text{OH})_2} = 9.38 \text{ mmol} \times \frac{1 \text{ L}}{0.0250 \text{ mol}}$$

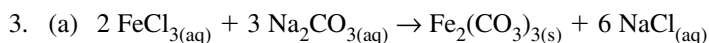
$$v_{\text{Ca}(\text{OH})_2} = 375 \text{ mL}$$

or

$$v_{\text{Ca}(\text{OH})_2} = 25.0 \text{ mL Al}_2(\text{SO}_4)_3 \times \frac{0.125 \text{ mol Al}_2(\text{SO}_4)_3}{1 \text{ mol Al}_2(\text{SO}_4)_3} \times \frac{3 \text{ mol Ca}(\text{OH})_2}{1 \text{ mol Al}_2(\text{SO}_4)_3} \times \frac{1 \text{ L Ca}(\text{OH})_2}{0.0250 \text{ mol}}$$

$$v_{\text{Ca}(\text{OH})_2} = 375 \text{ mL}$$

The volume of calcium hydroxide solution reacted is 375 mL.



75.0 mL v

$$n_{\text{FeCl}_3} = 75.0 \text{ mL} \times \frac{0.200 \text{ mol}}{1 \text{ L}}$$

$$n_{\text{Na}_2\text{CO}_3} = 15.0 \text{ mmol} \times \frac{3}{2}$$

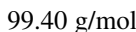
$$v_{\text{Na}_2\text{CO}_3} = 22.5 \text{ mmol} \times \frac{1 \text{ L}}{0.250 \text{ mol}}$$

$$v_{\text{Na}_2\text{CO}_3} = 75.0 \text{ mL} \cancel{\text{FeCl}_3} \times \frac{0.200 \text{ mol} \cancel{\text{FeCl}_3}}{1 \text{ L} \cancel{\text{FeCl}_3}} \times \frac{3 \text{ mol} \text{Na}_2\text{CO}_3}{2 \text{ mol} \cancel{\text{FeCl}_3}} \times \frac{1 \text{ L Na}_2\text{CO}_3}{0.250 \text{ mol Na}_2\text{CO}_3}$$

(b) A reasonable volume of sodium carbonate solution would be at least 100 mL, thus providing about a 10% *excess* to ensure a *complete* reaction.

Applying Inquiry Skills

4. (a) Prediction



$$n_{\text{NaOH}} = 20.0 \text{ mL} \times \frac{2.50 \text{ mol}}{1 \text{ L}}$$

$$n_{\text{Zn(OH)}_2} = 50.0 \text{ mmol} \times \frac{1}{2}$$

$$m_{\text{Zn(OH)}_2} = 25.0 \text{ mmol} \times \frac{99.40 \text{ g}}{1 \text{ mol}}$$

$$m_{\text{Zn(OH)}_2} = 20.0 \text{ mL } \cancel{\text{NaOH}} \times \frac{2.50 \text{ mol } \cancel{\text{NaOH}}}{1 \text{ L } \cancel{\text{NaOH}}} \times \frac{1 \text{ mol } \text{Zn(OH)}_2}{2 \text{ mol } \cancel{\text{NaOH}}} \times \frac{99.40 \text{ g } \text{Zn(OH)}_2}{1 \text{ mol } \cancel{\text{Zn(OH)}_2}}$$

Note: The calculated answer of 2.485 g can be rounded to 2.48 g or 2.49, depending on the rounding rule used in the classroom.

(b) Analysis

According to the evidence, the mass of zinc hydroxide that is actually produced is 2.39 g.

(c) Evaluation

$$\% \text{ difference} = \frac{0.10 \text{ g}}{2.48 \text{ g}} \times 100\% = 3.8\%$$

Note: The unrounded value of 2.485 g was used in the calculation. If the rounded value is used, the difference is 0.09 g and the percentage difference is 3.6%.

The prediction was 3.8% higher than the value obtained, and so is judged to be verified by the experimental results. A prediction within 5% is considered acceptably accurate (95% accurate) for this kind of lab work, with any difference probably just due to normal experimental error.

(d) The stoichiometric concept is supported by the results of this investigation, and judged to be acceptable because the prediction was verified. There is good confidence in this judgment, and no need is seen to modify the concept.

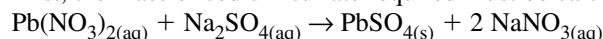
SECTION 7.6 QUESTIONS

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Applying Inquiry Skills

- Precipitating all the lead(II) ions will require adding an excess of a solution containing an anion (in this case, sulfate) that forms a low-solubility compound with lead(II). Making a sodium sulfate solution is a logical choice, since sodium sulfate is soluble, inexpensive, and easy to obtain.

First, the mass of sodium sulfate required must be calculated.



2.0 L *m*

0.34 mol/L 142.04 g/mol

$$n_{\text{Pb}(\text{NO}_3)_2} = 2.0 \cancel{\text{L}} \times \frac{0.34 \text{ mol}}{1 \cancel{\text{L}}} = 0.68 \text{ mol}$$

$$n_{\text{Na}_2\text{SO}_4} = 0.68 \text{ mol} \times \frac{1}{1} = 0.68 \text{ mol}$$

$$m_{\text{Na}_2\text{SO}_4} = 0.68 \cancel{\text{mol}} \times \frac{142.04 \text{ g}}{1 \cancel{\text{mol}}}$$

$$m_{\text{Na}_2\text{SO}_4} = 97 \text{ g}$$

$$\text{or } m_{\text{Na}_2\text{SO}_4} = \frac{0.34 \cancel{\text{mol Pb}(\text{NO}_3)_2}}{1 \cancel{\text{L Pb}(\text{NO}_3)_2}} \times \frac{1 \cancel{\text{mol Na}_2\text{SO}_4}}{1 \cancel{\text{mol Pb}(\text{NO}_3)_2}} \times \frac{142.04 \text{ g Na}_2\text{SO}_4}{1 \cancel{\text{mol Na}_2\text{SO}_4}}$$

$$m_{\text{Na}_2\text{SO}_4} = 97 \text{ g}$$

A *minimum* mass of 97 g of sodium sulfate must be used.

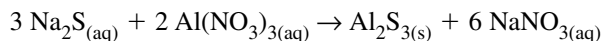
To ensure a complete precipitation of lead(II) ions, using an *excess* of sodium sulfate will be necessary. Commonly this amount should be about 10% more than the minimum required — say, 105 g of sodium sulfate, in this case.

The simplest process would be to obtain 105 g of sodium sulfate and dissolve it to make, say, 1.0 L of reacting solution. It is possible that not all of the solute will dissolve, depending on the water temperature. This problem could be solved by increasing the solvent volume.

- The experimental design is judged to be inadequate, because the mass of solid measured includes not only the precipitate but also the excess reactant and the second (soluble) product. The latter two chemicals crystallize out of the solution when the water is boiled away. The mass of solid remaining will be of more than one substance, with no way to calculate amounts from the value. The precipitate should have been separated by filtration and then dried.

If we take the reaction in Question 3 as an example, precipitating and then crystallizing would result in a mixture of the solids, aluminum nitrate (the excess reactant), aluminum sulfide (the precipitate), and sodium nitrate (the second (but soluble) product).

- (a) **Prediction**



20.0 mL *m*

0.210 mol/L 150.14 g/mol

$$n_{\text{Na}_2\text{S}} = 20.0 \cancel{\text{mL}} \times \frac{0.210 \text{ mol}}{1 \cancel{\text{L}}}$$

$$n_{\text{Na}_2\text{S}} = 4.20 \text{ mmol}$$

$$\begin{aligned}
 n_{\text{Al}_2\text{S}_3} &= 4.20 \text{ mmol} \times \frac{1}{3} \\
 n_{\text{Al}_2\text{S}_3} &= 1.40 \text{ mmol} \\
 m_{\text{Al}_2\text{S}_3} &= 1.40 \text{ mmol} \times \frac{150.14 \text{ g}}{1 \text{ mol}} \\
 m_{\text{Al}_2\text{S}_3} &= 210 \text{ mg} = 0.210 \text{ g}
 \end{aligned}$$

or

$$\begin{aligned}
 m_{\text{Al}_2\text{S}_3} &= 20.0 \text{ mL Na}_2\text{S} \times \frac{0.210 \text{ mol Na}_2\text{S}}{1 \text{ L Na}_2\text{S}} \times \frac{1 \text{ mol Al}_2\text{S}_3}{3 \text{ mol Na}_2\text{S}} \times \frac{150.14 \text{ g Al}_2\text{S}_3}{1 \text{ mol Al}_2\text{S}_3} \\
 m_{\text{Al}_2\text{S}_3} &= 210 \text{ mg} = 0.210 \text{ g}
 \end{aligned}$$

According to the stoichiometric method, the mass of aluminum sulfide produced is predicted to be 0.210 g.

(b) **Analysis**

$$m_{\text{Al}_2\text{S}_3} = 1.17 \text{ g} - 0.97 \text{ g} = 0.20 \text{ g}$$

According to the evidence, the mass of aluminum sulfide that is actually produced is 0.20 g.

(c) **Evaluation**

The design of the experiment is judged to be adequate, with no obvious flaws. It allowed the question to be answered easily with simple materials, concepts, and procedures.

$$\text{difference} = |0.20 \text{ g} - 0.210 \text{ g}| = 0.01 \text{ g}$$

$$\% \text{ difference} = \frac{0.01 \text{ g}}{0.210 \text{ g}} \times 100\% = 5\%$$

The prediction was 5% higher than the value obtained, and is judged to be verified by the experimental results. A prediction within 5% is considered acceptably accurate (95% + ...) for this kind of lab work, with any difference probably just due to normal experimental error.

The stoichiometric concept is the authority for this investigation. It is supported by the results of this investigation, and judged to be acceptable because the prediction was verified. There is good confidence in this judgment, and there is no need to modify the concept.

4. (a) **Analysis**

$$v_{\text{AgNO}_3} = 100 \text{ mL} = 0.100 \text{ L}$$

$$m_{\text{AgNO}_3} = 6.74 \text{ g} - 1.27 \text{ g} = 5.47 \text{ g}$$

$$M_{\text{AgNO}_3} = 169.88 \text{ g/mol}$$

$$n_{\text{AgNO}_3} = 5.47 \text{ g} \times \frac{1 \text{ mol}}{169.88 \text{ g}}$$

$$= 0.0322 \text{ mol}$$

$$C_{\text{AgNO}_3} = \frac{0.0322 \text{ mol}}{0.100 \text{ L}}$$

$$C_{\text{AgNO}_3} = 0.322 \text{ mol/L}$$

The molar concentration of silver nitrate in solution is 0.322 mol/L.

Making Connections

- The most common way to check the concentration of antifreeze is to measure its density, since density and concentration are proportional. Battery acid is measured the same way (to determine battery charge). The higher the concentration of solute, the denser the solution will be, and the higher in the solution a buoyant object will float. A calibrated device used to measure liquid density in this way is called a hydrometer. Common density units for solutions are g/mL and kg/L, but concentrations of automotive fluids are usually expressed in terms of what they are intended to do: lowest working temperature (for antifreeze), or charge condition (for battery acid). It is also common to state solution densities as a ratio with the density of pure water — giving a numerical value greater or less than 1 — called the “specific gravity” of the solution.

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