

Totally Epic AP Chem Review: Gravimetric Analysis!

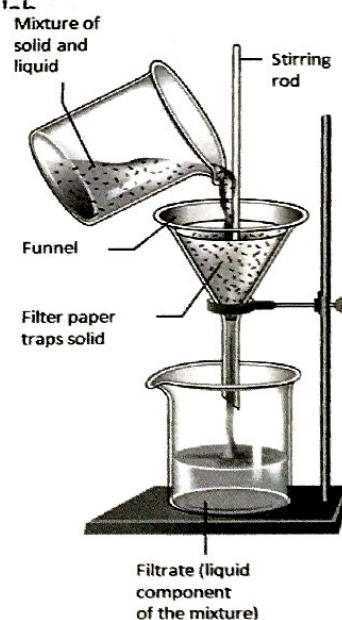
Gravimetric Analysis: analytical technique used to find the percent mass of an ion in an impure substance.

Method:

1. Reactant impure material with a known compound, forming a precipitate which contains the ion of interest.
2. Filter and dry precipitate, then measure its mass.
3. Use stoichiometry calculations to determine the mass of the ion of interest, using the balanced reaction to work backwards from the mass of the precipitate measured in the ¹⁻⁴

Steps in Gravimetric Analysis (Lab Technique)

1. Dry and weigh impure sample.
2. Dissolve sample in water.
3. Add precipitating reagent in excess.
4. Filter precipitate from solution.
5. Wash precipitate to remove soluble impurities.
6. Dry and weigh to constant mass



Steps in Gravimetric Analysis (Math-y Math)

1. Write the balanced precipitation reaction.
2. Use stoichiometry to convert from the mass of your precipitate to the mass of the ion of interest.

3. Calculate % by mass: $\% \text{ mass of ion} = \frac{\text{mass of ion (calculated in step 2)}}{\text{mass of original sample, NOT precipitate!}} \times 100$

Example: The following data were obtained when a sample of unknown chloride was analyzed:

Mass of watch glass	8.2030 g
Mass of unknown chloride sample + watch glass	8.3014 g
Mass of AgCl(s) + watch glass (after drying)	8.432 g

- a. Calculate the mass of the precipitate.

$$8.432 - 8.2030 = 0.229 \text{ g AgCl}$$

- b. Calculate the mass of chloride ions in the original sample.

$$0.229 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.32 \text{ g AgCl}} \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol AgCl}} \times \frac{35.45 \text{ g Cl}^-}{1 \text{ mol Cl}^-} = 0.0566 \text{ g Cl}^-$$

- c. Calculate the percent mass of chloride ions in the sample.

$$\% \text{ Cl}^- = \frac{0.0566 \text{ g Cl}^-}{\frac{8.3014 - 8.2030}{\text{Sample}}} \times 100 = 57.5\%$$

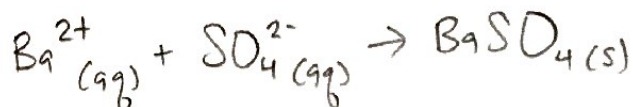
Common Lab Errors in Gravimetric Analysis

Error	Effect on Calculated % by Mass of Ion
Failure to add precipitating reagent in excess.	<ul style="list-style-type: none"> Not all of the ion of interest will precipitate, thus some of the ion will be lost in the filtrate. The calculated ion % will be \downarrow than the actual ion % by mass.
Failure to wash the precipitate to remove soluble impurities.	<ul style="list-style-type: none"> Some impurities will be left in the precipitate, so measured mass includes both the precipitate and these contaminants. The calculated ion % will be \uparrow than the actual ion % by mass.
The precipitate is excessively washed or washed with warm water.	<ul style="list-style-type: none"> Although the precipitate is "insoluble", small amounts of precipitate will dissolve into the wash water and be carried into the filtrate. The calculated ion % will be \downarrow than the actual ion % by mass.
The precipitate is not dried to "constant mass".	<ul style="list-style-type: none"> Not all of the water molecules will have been removed, so measured mass includes both the precipitate and remaining water molecules. The calculated ion % will be \uparrow than the actual ion % by mass.

Let's practice!

1. A 3.187 g impure sample containing sulfur and an inert material was to be analyzed for sulfur. As part of the procedure, the sample is dissolved, and the sulfur is converted to sulfate ion, SO_4^{2-} . Barium nitrate is added, which causes the sulfate to precipitate out as BaSO_4 . The dried BaSO_4 has a mass of 2.005 g.

- a. Write the balanced net ionic equation for the precipitation reaction.



- b. What mass of sulfur was in the original sample?

$$2.005 \text{ g BaSO}_4 \times \frac{1 \text{ mol BaSO}_4}{233.39 \text{ g BaSO}_4} \times \frac{1 \text{ mol S}}{1 \text{ mol BaSO}_4} \times \frac{32.06 \text{ g S}}{1 \text{ mol S}} = \boxed{0.2754 \text{ g S}}$$

- c. What is the percent of sulfur in the original ore?

$$\% \text{ S} = \frac{0.2754 \text{ g S}}{3.187 \text{ g Sample}} \times 100 = \boxed{8.641 \% \text{ S}}$$