Today: Titrations

Main Points:
1. Draw Picture
2. Chemistry Process
3. Notes
4. Examples

For titrations, life is always easier with a picture:
- There are hard problems out there, but this makes them easy:

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\[ \text{Titrant} \]
\[ \text{Conc. } \quad \text{Always a liquid} \]
\[ \text{Volume } \]

\[ \text{Titrand or Analyte} \]
\[ \text{Conc. } \quad \text{or mass } \]
\[ \text{Volume } \]
\[ \text{Liquid or solid} \]
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Titrant:
- Usually well known
- If not, this we are "standardizing"

Volume:
- This is determined by experiment
- \( (V_f - V_i) \pm 0.02 \text{ mL} \), reading bottom of the meniscus

Titrand:
- Concentration unknown unless standardizing
- Volume: Known - how much did you add?
- Mass: Amount of solid added

Titrations only work if 1 unknown variable
Titrations are based on fast (immediate), reversible reactions.

Ex: Acid-base neutralization.
This means that you don't need to redissolve while titrating!

We use some chemical (or physical) means to tell if reaction is done (this is in the bottom flask).

Stage 1: Lots of unreacted titrant.
Stage 2: Less unreacted titrant.
Stage 3: Exactly done (Equivalent Point).
Stage 4: Over done.

Example: Titrating 50.00 mL of 0.08 M HCl with 0.100 M NaOH.

Stage 1: Setup:
- Put 50.00 mL HCl in flask.
- Put NaOH in buret, flush buret (not into HCl), remove V-inch.
- Add a drop or 2 of phenolphthalein.

Observation: Clear. 50.00 mL HCl, 0.04 M H⁺.

Stage 1: Start adding titrant:
- 10 mL NaOH.

Observe: clear liquid.

Added 0.01 L NaOH (0.100 M): 0.01 mol NaOH
\[ H^+ + OH^- \rightarrow H_2O \]

0.01 mol H⁺ - 0.01 mol (H⁺) = 0.003 mol H⁺

\[ \text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]
Stage 2: Kept adding NaOH (smaller increments)
   - Added 38.00 mL NaOH (total)
   - N\text{H}_3 \text{O}^+ \text{ mole} @ \text{start}
   - Added 0.0380 M \times 0.1000 M = 0.0038 \text{ mole OH}^-

0.004 - 0.0038 = 0.0002 \text{ mole H}^+

→ observe: as the solns mix, get some pink color

Stage 3: @ endpoint
   - Added 40.00 mL NaOH (0.004 \text{ molar})
   - exactly neutralized the H\text{H}^+ \rightarrow 0 \text{ mols H}^+ \text{ left}
   - observe faint pink that fades in 30 sec.

\[ \text{CO}_2(g) + \text{OH}^- \rightarrow \text{HCO}_3^- + \text{OH}^- \]

Stage 4: past endpoint
   - added 45.00 mL NaOH (0.0045 \text{ molar})
   - less 0.004 \text{ mol H}^+ \text{ leaves} 0.0005 \text{ mol OH}^-
   - observe strong pink!
Notes

- This is a big concept (but simple)

- Titrations are all about moles (note prev. example)

- We are putting x moles of acid in the bottom and adding base until it is neutralized

- Adding 2-3 drops of indicator doesn't matter (does dilute the sample)

- Usually we add 50-150 mL H₂O to bring up volume for each working

- When adding the solid we just toss in the H₂O, not measured

Titration are the mole superduperimagery!

Examples

Calc. volume of 1.420 M NaOH needed to titrate 25.00 mL of 4.500 M H₃SO₄

A solution of this was prepared w/ 1.023 g Tris in 100.0 mL H₂O. 5.00 mL of the soln. was titrated w/ aqueous HNO₃ to reach the endpoint at 42.09 mL (methyl red)

What was [HNO₃]? 

H₂N⁺C(4H₂O)³⁻ FW=121.136

This (hydroxyethyl) amnomethane, "nonprotic"