

1. Qualitative analysis:

A. Multiple choice:

C 1. Which of the following gives a strong yellow color to the flame in a flame test?

- a. Li^+ b. Na^{++} c. Na^+ d. Cr^{+++}

C 2. Which of the following ions can be precipitated by adding $\text{NH}_3(\text{aq})$?

- a. Li^+ b. NH_4^+ c. Fe^{+++} d. K^+

B 3. Which of the following ions will not produce a red color in a flame test?

- a. Li^+ b. K^+ c. Ca^{++} d. None of the above

D 4. A red precipitate is a positive confirmation test for which of the following ions?

- a. Li^+ b. Cr^{+++} c. Fe^{+++} d. Cu^{++}

C 5. What substance is added to Cr^{+++} ions to produce a yellow solution?

- a. $\text{K}_4\text{Fe}(\text{CN})_6$ b. KSCN c. H_2O_2 d. NH_3

B. General knowledge test.

Below are two pairs of ions. Give one test (and only ONE test, more than one test will result in no points) that will separate the two ions. Give one test for each ion in the pair that will confirm the presence of that ion.

i. Fe^{+++} and Cu^{++}

Separation test: Add $\text{NH}_3(\text{aq})$. This will ppt Fe^{3+} as $\text{Fe}(\text{OH})_3$ and leave the Cu^{2+} in solution as a complex ion.

Confirmation of Fe^{+++} : Acidify the ppt and add KSCN . The red-maroon color indicates a positive test

Confirmation of Cu^{++} : Acidify the supernatant and add $\text{K}_4\text{Fe}(\text{CN})_6$. The red ppt indicates a positive test.

ii. Cr^{+++} and Cu^{++}

Separation test: Same as above with the Cr^{3+} ppting as $\text{Cr}(\text{OH})_3$

Confirmation of Cr^{+++} : Add H_2O_2 the yellow solution indicates the presence of Cr^{3+}

Confirmation of Cu^{++} : Same as above.

2. A solution is 1.6 M in $[\text{Ag}(\text{NH}_3)_2^+]$ and 1.25 M in free NH_3 . What is the **maximum $[\text{Cl}^-]$** that can be maintained in the solution without forming a precipitate of $\text{AgCl}_{(s)}$?

First determine the **free Ag^+** concentration:



$$K_f = \frac{[\text{Ag}(\text{NH}_3)_2^+]}{[\text{Ag}^+][\text{NH}_3]^2} = 1.6 \times 10^7$$

$$[\text{Ag}^+] = \frac{[\text{Ag}(\text{NH}_3)_2^+]}{K_f [\text{NH}_3]^2} = \frac{1.6}{(1.6 \times 10^7)(1.25)^2} = 6.4 \times 10^{-8} \text{ M}$$

Use this $[\text{Ag}^+]$ to determine the $[\text{Cl}^-]$:



$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10}$$

$$[\text{Cl}^-] = \frac{K_{sp}}{[\text{Ag}^+]} = \frac{1.8 \times 10^{-10}}{6.4 \times 10^{-8}} = 2.8 \times 10^{-3} \text{ M}$$

3. A student mixes 78.43 mL of 0.1204 M Nitric acid with 44.54 mL of 0.4240 M Calcium Hydroxide solution. What is the pH of the resulting solution?

$$\text{mmol H}^+ = (78.43 \text{ mL})(0.1204 \text{ M}) = 9.443 \text{ mmol}$$

$$\text{mmol OH}^- = (44.54 \text{ mL})(0.4240 \text{ M}) \left(\frac{2 \text{ mol OH}^-}{1 \text{ mol Ca}(\text{OH})_2} \right) = 37.77 \text{ mmol}$$

$\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$	
9.443	37.77
-9.443	-9.443
0	28.33

$$pOH = -\log[\text{OH}^-] = -\log\left(\frac{28.33}{122.97}\right) = 0.6376$$

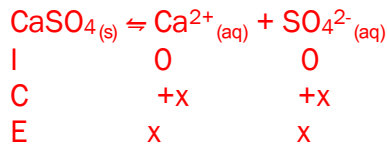
$$pH = 14.0000 - pOH = 14.0000 - 0.6376 = 13.3624$$

4. A particular water sample has 131 ppm of CaSO_4 (131 g CaSO_4 per 10^6 g water). If this water is boiled in a teakettle, approximately what **fraction of water must be evaporated** before $\text{CaSO}_4(\text{s})$ begins to precipitate? Assume that the solubility of $\text{CaSO}_4(\text{s})$ does not change much in the temperature range 0 to 100°C .

First calculate the molarity of the water sample:

$$\frac{131 \text{ g } \cancel{\text{CaSO}_4}}{10^6 \text{ g } \cancel{\text{H}_2\text{O}}} \times \frac{1 \text{ mol } \text{CaSO}_4}{136.141 \text{ g } \cancel{\text{CaSO}_4}} \times \frac{1 \text{ g } \cancel{\text{H}_2\text{O}}}{1 \text{ mL } \cancel{\text{H}_2\text{O}}} \times \frac{1 \text{ mL } \cancel{\text{H}_2\text{O}}}{10^{-3} \text{ L } \cancel{\text{H}_2\text{O}}} = 9.62 \times 10^{-4} \text{ M}$$

Then calculate the molarity of a saturated CaSO_4 solution:



$$K_{sp} = [\text{Ca}^{2+}][\text{SO}_4^{2-}] = x^2 = 1.2 \times 10^{-6} \Rightarrow x = 1.1 \times 10^{-3} \text{ M}$$

Now it's a dilution problem:

$$M_1 V_1 = M_2 V_2$$

$$\frac{V_1}{V_2} = \frac{M_2}{M_1} = \frac{9.62 \times 10^{-4}}{1.1 \times 10^{-3}} = 0.87$$

We must evaporate 87% of the original solution volume.

5. How many grams of sodium nitrite must be added to 150.0 mL of 0.1115 M Nitrous acid to make a solution that has a pH of 3.50?

The $\text{p}K_a$ of nitrous acid is 3.37. Because the pH is close to this the solution is a buffer.

$$\text{pH} = \text{p}K_a + \log \frac{\text{base}}{\text{acid}} = 3.50 = 3.37 + \log \frac{x}{16.72 \text{ mmol}}$$

$$0.13 = \log \frac{x}{16.72 \text{ mmol}}$$

$$x = 16.72 \text{ mmol} \times 10^{0.13} = 22.56 \text{ mmol } \text{NO}_2^- \times \frac{1 \text{ mmol } \text{NaNO}_2}{1 \text{ mmol } \text{NO}_2^-} \times \frac{10^{-3} \text{ mol}}{1 \text{ mmol}} \times \frac{68.9953 \text{ g } \text{NaNO}_2}{1 \text{ mol } \text{NaNO}_2}$$

$$= 1.6 \text{ g } \text{NaNO}_2$$

6. A 0.0536 M solution of a weak base has a pH of 10.05. Calculate K_b for the base.



ICE Table:

	B	OH ⁻	BH ⁺
I	0.0536	10^{-7}	0
C	-x	+x	+x
E	0.0536-x	$10^{-7} + x$	x

Assume that the 10^{-7} is negligible. This gives.

$$[OH^-] = 10^{-pH} = 10^{-(14.00-10.05)} = 10^{-3.95} = 0.000112$$

$$K_b = \frac{[OH^-][BH^+]}{[B]} = \frac{(0.000112)(0.000112)}{(0.0536)} = 2.3 \times 10^{-7}$$

7. A 1.0-L solution that is 4.2 M in ammonia is mixed with 26.7 g of ammonium chloride.

- a. What is the hydroxide ion concentration of this solution?

Ammonia and ammonium make a buffer.

$$\text{mol NH}_3 = 1.0 \text{ L} \times \frac{4.2 \text{ mol NH}_3}{\text{L}} = 4.2 \text{ mol NH}_3$$

$$\text{mol NH}_4^+ = 26.7 \text{ g NH}_4\text{Cl} \times \frac{1 \text{ mol NH}_4\text{Cl}}{53.491 \text{ g NH}_4\text{Cl}} \times \frac{1 \text{ mol NH}_4^+}{1 \text{ mol NH}_4\text{Cl}} = 0.499 \text{ mol NH}_4^+$$

$$pH = pK_a + \log \frac{\text{base}}{\text{acid}} = (14.00 - (-\log(1.8 \times 10^{-5}))) + \log \frac{4.2 \text{ mol}}{0.499 \text{ mol}}$$

$$= 10.18$$

$$pOH = 14.00 - pH = 3.82$$

$$[OH^-] = 10^{-pOH} = 10^{-3.82} = 1.52 \times 10^{-4} \text{ M}$$

- b. 0.075 mol of $MgCl_2$ is added to the above solution. Assume that there is no volume change. After $Mg(OH)_2$ has precipitated, what is the molar concentration of magnesium ion? What percent of the Mg^{2+} is removed from the solution?

The hydroxide ion concentration will not change significantly because the solution is buffered. We can calculate the magnesium ion concentration directly.

$$K_{sp} = [Mg^{2+}][OH^-]^2 = 1.8 \times 10^{-11}$$

$$[Mg^{2+}] = \frac{K_{sp}}{[OH^-]^2} = \frac{1.8 \times 10^{-11}}{(1.52 \times 10^{-4})^2} = 7.9 \times 10^{-4} \text{ M} \times 1.0 \text{ L} = 7.9 \times 10^{-4} \text{ mol}$$

$$\% \text{ removed} = \frac{(0.075 \text{ mol} - 7.9 \times 10^{-4} \text{ mol})}{0.075 \text{ mol}} \times 100 = 99\%$$