

Topics:

1) Common Ion Effect in Acid-Base Reactions

Examples: weak acid mixed with salt containing its conjugate base.
weak base mixed with salt containing its conjugate acid.

- Can use ICE or can use Henderson-Hasselbalch Equation to solve. $pH = pK_a + \log \frac{[A^-]}{[HA]}$
 - Remember: In common ion situations, the equilibrium does not shift much so “-x and +x” factors are negligible. Thus, you will use initial concentrations (or moles) in the HH equation.
 - Common ion effect is negligible when a weak acid is added to a strong acid (pH can be determined by strong acid only.) or when a weak base is added to a strong base (pH can be determined by strong base only).
- 2) **Buffers:** consist of similar amounts of weak acid + salt of conj. base or weak base + salt of conj. acid.
- A buffer will resist changing its pH when an acid or a base is added.
 - The buffer capacity is the amount of acid or base that can be added before the buffer is destroyed.
 - Buffer systems involve the common ion effect. Thus, problems can be solved with ICE or with the Henderson-Hasselbalch Equation as above.
 - Buffers usually have pH values close to the K_a of the weak acid in the buffer. This is because when looking at Henderson-Hasselbalch equation one sees that if $[HA] = [A^-]$, then $pH = pK_a$
- 3) **Acid-Base Titrations:** concept, curves, equivalence point, halfway to equivalence point, indicators
- a) Strong acid - strong base titrations: pH is neutral at equivalence point. No buffer ever formed.
 - b) Weak acid - strong base titrations:
 - halfway to equivalence point: buffer solution formed where $[HA] = [A^-]$. Thus, $pH = pK_a$
 - at equivalence point, $pH > 7$ because pH is determined by hydrolysis of A^- (HA is consumed.)
 - past the equivalence point, the unconsumed strong base determines the pH (hydrolysis of A^- is negligible.)
 - c) Weak base - strong acid titrations:
 - halfway to equivalence point: buffer solution formed where $[B] = [HB^+]$. Thus, $pOH = pK_b$
 - at equivalence point, $pH < 7$ because pH is determined by hydrolysis of HB^+ (B is consumed.)
 - past the equivalence point, the unconsumed strong acid determines the pH (hydrolysis of HB^+ is negligible.)
 - d) pH indicators in titrations:
 - To be useful in titrations, pH indicators must change color over the range of pH near the equivalence point (steep part of curve.)
 - pH indicators change color when $[HIn] = [In^-]$, so this is when $pH = pK_a$ of the indicator.
 - Thus, choose an indicator with a pK_a similar to the pH at the equivalence point of the titration. This ensures that the indicator will change color at the equivalence point of the titration.

4) Solubility Equilibria:

- a) Solubility product (K_{sp}): Equilibrium constant determined when a solid is dissolved into water.
- K_{sp} expressions represent solutions that are saturated because this is when a solution is at equilibrium. (Never include concentration of the solid in a K_{sp} expression!)
 - K_{sp} gives a general measure of solubility because when K_{sp} is greater there are more ions dissolved in the saturated solution.
- b) Relationship between solubility and K_{sp}
- Solubility is the maximum grams of solid that will dissolve in 1 L of solution. (g/L)
 - Molar solubility is the maximum moles of solid that will dissolve in 1 L of solution. (moles/L)
 - Make conversions between solubility (or molar solubility) and K_{sp} . (Use ICE.)
- c) Analysis of Precipitation Reactions: Two aqueous solutions are mixed together. What happens?
- a) To determine if a ppt forms:
- First, find concentrations of all ions in new total solution assuming no reaction takes place.
 - Second, based on “solubility rules” determine what solid is likely to precipitate and write solubility equilibrium equation.
 - Third, solve for “Q”. If $Q > K_{sp}$, then ppt forms.
- b) To determine final concentrations of ions that do react: (*Note: The final concentrations of all ions that do not react—i.e. spectator ions—have already been determined.*)
- First, since most questions will involve very insoluble solids, assume that ions react completely to precipitate all possible solid. (Use stoichiometry: limiting reactant, ICF.)
 - Second, have this system come to equilibrium and determine final ion concentrations (use ICE and K_{sp}).

5) pH and Solubility

- a) Decreasing pH, increases $[H^+]$, and thus increases solubility of hydroxides. Increasing pH, increases $[OH^-]$, and thus decreases solubility of hydroxides.
- b) For salts (MA) containing the conjugate base of a weak acid (A^-), decreasing pH, increases $[H^+]$ which reacts with A^- to form HA, increasing solubility of MA.

6) Complex Ion Solubility

- a) The reaction of a metal ion with ligands (a Lewis acid-base reaction) is given by the formation constant, K_f , to form the complex ion.
- b) Complex ion formation is coupled with K_{sp} of a slightly soluble salt of the metal ion to increase the solubility of the salt; for the overall reaction, $K = K_f K_{sp}$.
- c) Amphoteric oxides and hydroxides exhibit increased solubility at low pH (high $[H^+]$) due to acid-base reaction to produce H_2O and metal ion. Solubility is also increased at high pH (high $[OH^-]$) due to formation of soluble hydroxo complex.

7) Separation of Ions by Fractional Precipitation

- Use specific precipitating reagents, containing anions making specific metal ions insoluble, to selectively precipitate one metal in order to separate two metals, or to identify the presence of different classes of ion in solution.
- If the precipitating reagent will precipitate both metals, find the needed concentration of the precipitating agent to almost completely precipitate the one metal before the other metal begins to precipitate.

8) Qualitative Analysis—general concepts and techniques