TITRATION

Titration is a way to identify unknown concentrations of acids or bases.

In titration reactions, you neutralize an unknown acid with a known concentration of base. By knowing the amount of moles of base added, you can determine the moles and molarity of acid.

It is possible to determine the unknown concentration of a base by doing the same with a known concentration of acid.

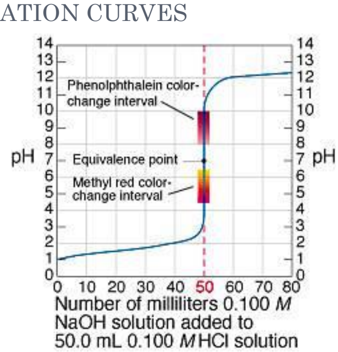
TITRATION CURVES

Titration curves are designed to graphically represent and determine the equivalence point during a titration.

Equivalence point - point at which the $[H_3O^+]$ is equal to the $[OH^-]$

End point of an indicator - pH where the indicator changes color.

TITRATION CURVES

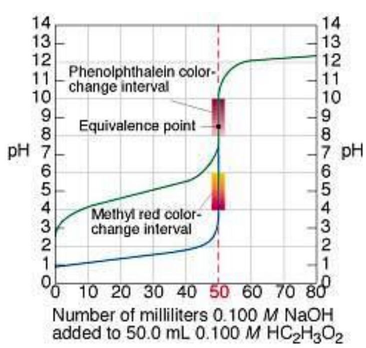


STRONG ACID/BASE TITRATIONS

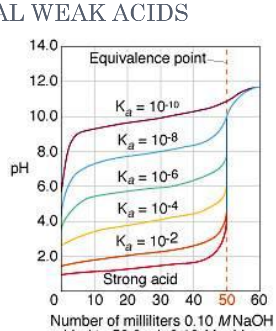
25.0 mL of a 0.20 M solution of hydroiodic acid is titrated with a 0.40 M solution of potassium hydroxide. Calculate the pH of the solution when the following amounts of titrant are added:

a) 0.0 mL
b) 10.0 mL
c) 12.5 mL
d) 15 mL

TITRATION CURVE - WEAK ACID



TITRATION CURVES OF SEVERAL WEAK ACIDS

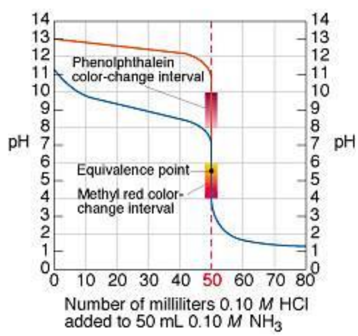


WEAK ACID/STRONG BASE TITRATION

A 30.0 mL sample of 0.12 M acetic acid (K_a = 1.8 x 10^{-5}) is titrated with 0.36 M sodium hydroxide. Determine the pH of the solution when the following amounts of titrant are added:

a) 0.0 mL
b) 9.0 mL
c) 10.0 mL
d) 11.0 mL

TITRATION CURVE OF A WEAK BASE



WEAK BASE/STRONG ACID TITRATION

- A 25 mL solution of 0.54 M hydrazine, H_2NNH_2 (K_b = 1.3 x 10^{-6}), a weak base, is titrated with 0.27 M HBr. Determine the pH when the following amounts of acid have been added to the base:
- a) 0.0 mL
- b) 48 mL
- c) 50 mL
- d) 52 mL

COMMON ION EFFECT

a) 500 mL of a 0.75 M HF solution ($K_a = 6.8 \times 10^{-4}$) is made. What is the pH of the solution?

b) 0.80 moles of NaF, a strong electrolyte, is added to the solution. What is the new pH of the solution?

COMMON-ION EFFECT

The addition of a salt ($NaC_2H_3O_2$) to a weak acid solution ($HC_2H_3O_2$) will lower the $[H^+]$ concentration due to Le Chatelier's Principle:

$$HC_2H_3O_2\ (aq) \Longleftrightarrow H^+\ (aq) + C_2H_3O_2^-\ (aq)$$

The result is a higher pH with fewer $[H^+]$ ions in water.

The common-ion effect says that the dissociation of a weak electrolyte is lowered by the addition of a strong electrolyte with a common ion in it.

BUFFERS

A buffered solution is a solution that resists a change in pH. A buffer contains an acidic species which will neutralize additional $[OH^-]$ ions and a basic species which will neutralize additional $[H^+]$ ions that are added.

These requirements are fulfilled by a weak acid-base conjugate pair, or by mixing a weak acid or a weak base with a strong electrolyte salt with ions that produce the conjugate.

A buffer is an application of the common ion effect.

BUFFERS

A buffer is made by mixing a solution so that it has a concentration of 0.25 M nitrous acid ($K_a = 4.5 \times 10^{-4}$) and 0.30 M sodium nitrite. What is the pH of the solution?

BUFFER CHARACTERISTICS

Buffers have varying pHs due to the K_a.

To calculate the pH of a buffer:

$$pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

Buffer capacity is the amount of acid or base a buffer can neutralize before the pH changes a lot.

The greater the amounts of conjugate acid-base pair in solution, the more resistance to pH change.

BUFFER PROBLEM

A buffer is made by mixing 0.40 moles of NH_3 with 0.50 moles of NH_4Cl to a solution of 100 mL.

a) What is the pH of the buffer?

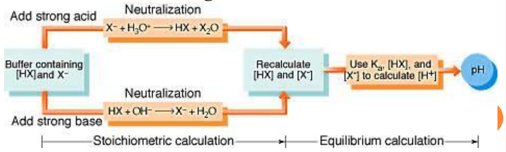
b) If 20 mL of 0.10 M HCl are added to the buffer, what will be the new pH?

c) If 20 mL of 0.10 M NaOH are added to the buffer, what will be the new pH?

ADDITION OF ACID OR BASE TO BUFFER

When small amounts of strong acid is added to a buffer, the weak base component of the buffer consumes all of the acid to keep the pH from changing.

The following strategy can be used to determine the new pH of a buffer after addition of strong acid or base:



DISSOLUTION & PRECIPITATION

Dissolution is the process in which an ionic solid dissolves in a polar liquid to form ions.

During dissolution, the liquid becomes saturated with ions that are uniformly distributed.

Equation: $FeCl_3$ (s) <==> Fe^{+3} (aq) + 3 Cl^- (aq)

Precipitation is the process in which ions leave a solution and regenerate an ionic solid.

Equation: Ag^+ (aq) + Br^- (aq) <==> AgBr (s)

Dissolution and precipitation are opposite processes.

SOLUBILITY PRODUCT

Ionic compounds have differing degrees of solubility. Each compound has its own solubility product (K_{sp}).

Equation: $Fe(OH)_2$ (s) <==> Fe^{+2} (aq) + 2 OH^- (aq)

The solubility product expression is determined in the same way as the equilibrium constant (K_c).

Since each substance has it's own K_{sp}, it is possible to determine how much solid will dissolve in a solution.

Remember: Solids and Liquids are not included in K values, only gases and aqueous solutions.

Examples

1) Chromium (III) hydroxide is a very insoluble salt. In 10.0 L of water, only 0.0161 mg of the solid will dissolve. What is the solubility product for chromium (III) hydroxide?
2) What concentration of ions will be present in a saturated solution of PbF_2 ($K_{sp} = 3.6 \times 10^{-8}$)

Ion Product (Will Precipitate Form)

A precipitate will form if the ion product (Q) of the concentrations of ions is larger than the Ksp.

In this case, the solution is supersaturated with ions, and some must precipitate to bring the concentration down.

Example: Will a precipitate of chromium (III) hydroxide form if a 1.4×10^{-5} M solution of $CrCl_3$ is placed in a solution with a pH of 7.4?

What will happen if the Q is less than the Ksp? What type of solution is this an example of?

What if Q = Ksp? What kind of solution?

Common Ion Effect & Solubility

Solubility can also be affected by the common ion effect, in which common ions may lower solubility.

Example:

Determine the solubility of $Cr(OH)_3$ in

a) Pure water

b) A solution with a pH of 10.