

Continue



Titration WS  
 Name Key Period \_\_\_\_\_ Date \_\_\_\_\_  
 1. The point where equal moles of 2 chemicals have been combined.  
 2. Define moles: moles or concentration  
 3. A solution has a concentration of 0.1M. What is the concentration in a factor of moles and liters? 0.1 mol/L  
 4. Answer 10 mL of 0.1M HCl is mixed with 20 mL of 0.1M NaOH. Calculate the molarity of the solution.  

$$\frac{25 \text{ mol}}{100 \text{ mL}} \times \frac{10 \text{ mL} + 20 \text{ mL}}{100 \text{ mL}} = \frac{0.25 \text{ mol}}{30 \text{ mL}} = 0.00833 \text{ mol/L} = 0.00833 \text{ M}$$
  
 5.  $\text{LiOH} + \text{HNO}_3 \rightarrow \text{LiNO}_3 + \text{H}_2\text{O}$   

$$\frac{30 \text{ mL}}{100 \text{ mL}} \times \frac{0.1 \text{ mol/L}}{100 \text{ mL}} = \frac{0.03 \text{ mol}}{100 \text{ mL}} = 0.0003 \text{ mol/L} = 0.0003 \text{ M}$$
  
 6. A solution is prepared by adding 40 mL of 0.1M HCl to 100 mL of 0.1M NaOH. Calculate the molarity of the solution.  

$$2 \text{HCl} + \text{Ca(OH)}_2 \rightarrow \text{CaCl}_2 + 2 \text{H}_2\text{O}$$
  

$$\frac{30 \text{ mL}}{100 \text{ mL}} \times \frac{0.1 \text{ mol/L}}{100 \text{ mL}} = \frac{0.03 \text{ mol}}{100 \text{ mL}} = 0.0003 \text{ mol/L} = 0.0003 \text{ M}$$
  
 7. How do you know when the endpoint of a titration has been reached?  
 (Color change)  
 1) Color change occurs (ind/cator changes color)  
 2) pH changes drastically and suddenly  
 8. A weak acid (HA) is titrated with a weak base (BOH). Draw a titration curve for this on the axes below.  
 The pH at the equivalence point is below 7 because the strong acid is stronger than the weak base.  

$$\frac{30 \text{ mL}}{100 \text{ mL}} \times \frac{0.1 \text{ mol/L}}{100 \text{ mL}} = \frac{0.03 \text{ mol}}{100 \text{ mL}} = 0.0003 \text{ mol/L} = 0.0003 \text{ M}$$
  

$$\frac{30 \text{ mL}}{100 \text{ mL}} \times \frac{0.1 \text{ mol/L}}{100 \text{ mL}} = \frac{0.03 \text{ mol}}{100 \text{ mL}} = 0.0003 \text{ mol/L} = 0.0003 \text{ M}$$
  

$$\text{H}_2\text{SO}_4 + 2\text{KOH} \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$$

Name \_\_\_\_\_ Hr \_\_\_\_\_

Chemistry Aqueous Acids and Bases - Titration

- 49 mL of 0.200 M HCl is mixed with 50 mL of 0.200 M NaOH to reach the endpoint.
  - moles HCl =
  - moles NaOH =
  - excess moles NaOH =
  - [H<sup>+</sup>]
  - [OH<sup>-</sup>]
  - pOH =
  - pH =
- 86.30 mL of an HCl solution was required to neutralize 31.75 mL of 0.150 M NaOH. Determine the molarity of the HCl.
- 63.15 mL of calcium hydroxide is required to titrate 18.9 mL of a 0.200 M H<sub>3</sub>PO<sub>4</sub> solution. What is the molarity of the basic solution?
- How many mL of 0.160 M HClO<sub>4</sub> are needed to titrate 35.0 mL of 0.215 M LiOH?
- 25.0 mL of 1.00 M HCl are required to titrate a Drano solution (active ingredient NaOH). How many moles of NaOH are present in the solution?
- Ten grams of vinegar (containing acetic acid, HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub>, is titrated with 65.40 mL of 0.150 M NaOH.
  - How many moles of acetic acid are present in ten grams of vinegar?
  - How many grams of acetic acid are present in ten grams of vinegar?

**KEY**

- e) What colour would 0.10 M CH<sub>3</sub>COOH be in this indicator?  
 (Show how you got [H<sub>3</sub>O<sup>+</sup>]) (2)  

$$[\text{H}_3\text{O}^+] = \sqrt{C_0 K_a} = \sqrt{0.10 (1.8 \times 10^{-5})}$$
  

$$[\text{H}_3\text{O}^+] = 1.34 \times 10^{-3} \text{ M}$$
  

$$\text{pH} = 2.87$$
  
 Answer green (2)
7. An indicator HInd turns yellow in 0.10 M HCl and blue in 0.10 M NaOH.
- Write the equation describing the equilibrium in HInd. (1)  

$$\text{HInd}_{\text{yellow}} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{Ind}^-_{\text{blue}}$$
 (1)
  - What colour is HInd? (1) yellow What colour is Ind<sup>-</sup>? (1) blue (1)
  - HInd is green in the range pH = 5.4 to pH = 6.2. Determine the K<sub>a</sub> of HInd. (1)  

$$\text{p}K_a = \text{pH} = 5.8$$
  

$$K_a = 2 \times 10^{-6}$$
 (1)
  - When a few drops of HInd are added to a weak acid HA<sub>1</sub>, the colour is yellow. Which is the stronger acid, HInd or HA<sub>1</sub>? (1)  

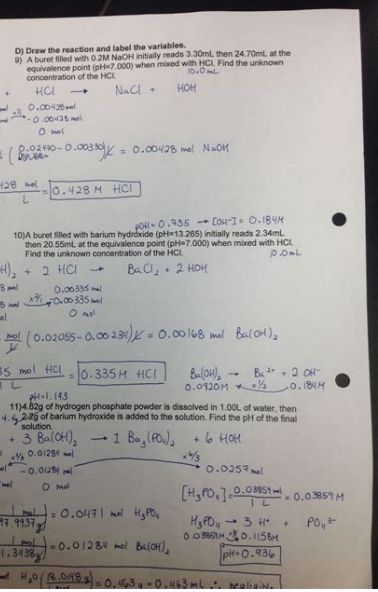
$$\text{HA}_1 + \text{Ind}^- \rightleftharpoons \text{HInd} + \text{A}_1^-$$
  
 SrA yellow wrA Answer HA<sub>1</sub> is the stronger acid. (1)
  - When a few drops of HInd are added to a weak acid HA<sub>2</sub>, the colour is blue. Which is the stronger acid, HInd or HA<sub>2</sub>? (1)  

$$\text{HA}_2 + \text{Ind}^- \rightleftharpoons \text{HInd} + \text{A}_2^-$$
  
 SrA blue wrA Answer HInd is the stronger acid. (1)
  - Which acid is stronger, HA<sub>1</sub>, or HA<sub>2</sub>? (1) HA<sub>1</sub> (1)
  - List the acids HInd, HA<sub>1</sub>, and HA<sub>2</sub> in order of strength from strongest to weakest. (1) (1)  
HA<sub>1</sub> > HInd > HA<sub>2</sub>
  - List the bases Ind<sup>-</sup>, A<sub>1</sub><sup>-</sup>, and A<sub>2</sub><sup>-</sup> in order of strength from strongest to weakest. (1) (1)  
A<sub>2</sub><sup>-</sup> > Ind<sup>-</sup> > A<sub>1</sub><sup>-</sup>

### Density Change in Distillation

- The arrangement of apparatus used to prepare halogenoalkane in methanolamine and water respectively.
  - What is observed when a piece of magnesium ribbon is placed in beaker A and B?
  - State the reason for your answer to 5b.
  - Name the particles present in beaker A.
  - Name the particles present in beaker B.
- Magnesium ribbon is removed. Water is added to the solution in beaker A and the mixture is stirred slowly. When colour change ceases to occur, effluencement occurs.
  - Name the gas and suggest a suitable test to identify the gas.
  - State the role of water in the reaction that caused the evolution of the gas.
  - Write an ionic equation for the reaction involving the evolution of the gas.
  - What is the conclusion that can be made from the observation?
- What is the minimum volume of 2.0 mol dm<sup>-3</sup> sodium hydroxide that is required to react completely with 1g of calcium carbonate? (A.A. 101.09, C. 12.01, O. 16.00)
- In a titration process, 5.1 mol dm<sup>-3</sup> sulphuric acid from a burette is added slowly to 20cm<sup>3</sup> of 0.2 mol dm<sup>-3</sup> sodium carbonate solution in a conical flask with methyl orange indicator until neutralization occurs. What is the total volume in the conical flask at the end point of titration?

WJEC Chemistry



Which indicator is suitable for a weak-acid strong-base titration and why. What does a strong base and weak acid make. Can you titrate a weak acid with a weak base.

Acetic acid is an example of a weak base. ELLA MARU STUDIO / Getty Images Definition: A weak base is a base that is partially dissociated in an aqueous solution. Examples: NH<sub>4</sub>OH Titration is an analytical chemistry technique used to find an unknown concentration of an analyte (the titrand) by reacting it with a known volume and concentration of a standard solution (called the titrant). Titrations are typically used for acid-base reactions and redox reactions. Here's an example problem determining the concentration of an analyte in an acid-base reaction: A 25 ml solution of 0.5 M NaOH is titrated until neutralized into a 50 ml sample of HCl. What was the concentration of the HCl? Step 1: Determine [OH<sup>-</sup>] Every mole of NaOH will have one mole of OH<sup>-</sup>. Therefore [OH<sup>-</sup>] = 0.5 M. Step 2: Determine the number of moles of OH<sup>-</sup>: Molarity = number of moles/volume Number of moles = Molarity x Volume Number of moles OH<sup>-</sup> = (0.5 M)(0.025 L)Number of moles OH<sup>-</sup> = 0.0125 mol Step 3: Determine the number of moles of H<sup>+</sup> When the base neutralizes the acid, the number of moles of H<sup>+</sup> = the number of moles of OH<sup>-</sup>. Therefore, the number of moles of H<sup>+</sup> = 0.0125 moles. Step 4: Determine the concentration of HCl Every mole of HCl will produce one mole of H<sup>+</sup>, therefore, the number of moles of HCl = number of moles of H<sup>+</sup>. Molarity = number of moles/volume Molarity of HCl = (0.0125 mol)/(0.05 L)Molarity of HCl = 0.25 M Answer The concentration of the HCl is 0.25 M. The above steps can be reduced to one equation: Macid/Vacid = Mbase/Vbase where Macid = concentration of the acid/Vacid = volume of the acid/Mbase = concentration of the base/Vbase = volume of the base This equation works for acid/base reactions where the mole ratio between acid and base is 1:1. If the ratio were different, as in Ca(OH)<sub>2</sub> and HCl, the ratio would be 1 mole acid to 2 moles base. The equation would now be: Macid/Vacid = 2Mbase/Vbase For the example problem, the ratio is 1:1: Macid/Vacid = Mbase/Vbase Macid(50 ml) = (0.5 M)(25 ml)Macid = 12.5 Mml/50 mlMacid = 0.25 M Different methods are used to determine the equivalence point of a titration. No matter which method is used, some error is introduced, so the concentration value is close to the true value, but not exact. For example, if a colored pH indicator is used, it might be difficult to detect the color change. Usually, the error here is to go past the equivalence point, giving a concentration value that is too high. Another potential source of error when an acid-base indicator is used is if water used to prepare the solutions contains ions that would change the pH of the solution. For example, if hard tap water is used, the starting solution would be more alkaline than if distilled deionized water had been the solvent. If a graph or titration curve is used to find the endpoint, the equivalence point is a curve rather than a sharp point. The endpoint is a sort of "best guess" based on the experimental data. The error can be minimized by using a calibrated pH meter to find the endpoint of an acid-base titration rather than a color change or extrapolation from a graph. Now consider the titration of a weak acid, HA, with NaOH titrant. The neutralization reaction is  $\text{HA} + \text{OH}^- \rightleftharpoons \text{A}^- + \text{H}_2\text{O}$ . There are four distinct regions in the course of such a titration with four different approaches to calculating the pH in each region. To illustrate this general approach, consider the titration of 10.0 mL of 0.0300 M benzoic acid (the analyte) with 0.0100 M NaOH(aq) solution (the titrant). Benzoic acid has the formula  $\text{C}_6\text{H}_5\text{CO}_2\text{H}$ , but for simplicity we will abbreviate it as  $\text{HBz}$ , and it has a  $K_a = 6.46 \times 10^{-5}$ . Calculate the pH at each of the following points in the titration: the initial point, before any titrant has been added; after adding 10.0 mL of 0.0100 M NaOH(aq); after adding 15.0 mL of 0.0100 M NaOH(aq); after adding 30.0 mL of 0.0100 M NaOH(aq); after adding 45.0 mL of 0.0100 M NaOH(aq). Once again, before carrying out any other calculations, determine (a) the number of millimoles of analyte initially present, and (b) the volume of titrant that must be added to reach the equivalence point. It will be useful to keep these numbers in mind as we proceed to calculate the pH values throughout the course of the titration. The number of millimoles of HBz initially present is  $n(\text{HBz}) = (0.0300, \text{M})(10.0, \text{mL}) = 0.300 \text{ mmol}$ . For our example titration the neutralization reaction is  $\text{HBz} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{NaBz}$ . At the equivalence point  $n(\text{HBz})_{\text{initial}} = n(\text{NaOH})_{\text{added}}$ . Therefore, the volume of NaOH(aq) solution needed to reach the equivalence point is  $V_{\text{NaOH}} = \frac{n_{\text{HBz}}}{M_{\text{NaOH}}} = \frac{0.300 \text{ mmol}}{0.0100 \text{ M}} = 30.0 \text{ mL}$ . Note that these are the same values we obtained in the example  $\text{HBr}/\text{NaOH}$  titration above, and we will be calculating the pH at the same of the same points throughout the course of the titration. You will want to compare the procedures and results in this weak acid - strong base example to those of the previous strong acid - strong base example. (i) The initial point, before any titrant has been added where we simply have a solution of 0.0300 M  $\text{HBz}$ . We must use  $K_a = 6.46 \times 10^{-5}$  to calculate  $[\text{H}_3\text{O}^+]$  and pH. But in this case,  $K_a \ll C_{\text{HBz}}$  (gg  $K_a$ ), so we can ignore hydrolysis of molecular  $\text{HBz}$  and use the approximate method to calculate  $[\text{H}_3\text{O}^+]$ , without needing to solve the quadratic equation.  $[\text{H}_3\text{O}^+] = \sqrt{C_{\text{HBz}} K_a} = \sqrt{(0.0300)(6.46 \times 10^{-5})} = 1.39 \times 10^{-3} \text{ M}$  therefore  $\text{pH} = 2.856$  (ii) After adding 10.0 mL of 0.0100 M NaOH(aq) The number of millimoles of added NaOH is  $n(\text{NaOH}) = (0.0100 \text{ M})(10.0 \text{ mL}) = 0.100 \text{ mmol}$ . We can summarize the numbers of millimoles of analyte and titrant before and after the addition by the following pseudo-ICE table: ICE Table  $\text{HBz} + \text{NaOH} \rightleftharpoons \text{Bz}^- + \text{H}_2\text{O}$  Initial  $0.300, 0.000, 0.000$  After Reaction  $0.200, 0.100, 0.000$  This is a buffer solution, so we really do not need to be concerned with the total volume of the solution at this point, because a buffer's pH is independent of volume. All we need to do is substitute the numbers of millimoles of HBz and Bz<sup>-</sup> into the  $K_a$  expression and solve for  $[\text{H}_3\text{O}^+]$ .  $K_a = \frac{[\text{H}_3\text{O}^+][\text{Bz}^-]}{[\text{HBz}]}$   $6.46 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+](0.100 \text{ mmol})}{(0.300 \text{ mmol})}$   $[\text{H}_3\text{O}^+] = 1.292 \times 10^{-4} \text{ M}$  therefore  $\text{pH} = 3.889$  (iii) After adding 15.0 mL of 0.0100 M NaOH(aq) The number of millimoles of added NaOH is  $n(\text{NaOH}) = (0.0100 \text{ M})(15.0 \text{ mL}) = 0.150 \text{ mmol}$ . We can summarize the numbers of millimoles of analyte and titrant before and after the addition by the following table: ICE Table  $\text{HBz} + \text{NaOH} \rightleftharpoons \text{Bz}^- + \text{H}_2\text{O}$  Initial  $0.300, 0.000$  After Reaction  $0.150, 0.150$  Note that we have added half the volume of NaOH(aq) solution needed to reach the equivalence point. This is called the half-titration point. As the table above shows, at the half-titration point, the solution contains equal numbers of millimoles of acid and conjugate base, making it an equimolar buffer solution. If we substitute into the  $K_a$  expression we obtain  $K_a = \frac{[\text{H}_3\text{O}^+][\text{Bz}^-]}{[\text{HBz}]}$   $6.46 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+](0.150 \text{ mmol})}{(0.150 \text{ mmol})}$   $[\text{H}_3\text{O}^+] = 6.46 \times 10^{-5} \text{ M}$  From which it follows  $\text{pH} = \text{p}K_a = -\log_{10}(6.46 \times 10^{-5}) = 4.190$ . This is a general result worth knowing: At the half-titration point in any weak acid strong base titration,  $\text{pH} = \text{p}K_a$  (iv) After adding 30.0 mL of 0.0100 M NaOH(aq) The number of millimoles of added NaOH is  $n(\text{NaOH}) = (0.0100 \text{ M})(30.0 \text{ mL}) = 0.300 \text{ mmol}$ . This is the equivalence point (see preliminary calculation above), because we have added as many millimoles of NaOH titrant as there were millimoles of HBz analyte in the sample. At this point all of the HBz has been converted to Bz<sup>-</sup>, the conjugate base ICE Table  $\text{HBz} + \text{NaOH} \rightleftharpoons \text{Bz}^- + \text{H}_2\text{O}$  Initial  $0.300, 0.000$  After Reaction  $0.000, 0.300$  We now have a solution of 0.300 mmol pure weak base  $\text{Bz}^-$  in a volume that is the sum of the initial volume plus the volume of added titrant; i.e.,  $V_{\text{total}} = V_{\text{HBz}} + V_{\text{NaOH}} = 10.0 \text{ mL} + 30.0 \text{ mL} = 40.0 \text{ mL}$ .  $[\text{Bz}^-] = \frac{0.300 \text{ mmol}}{40.0 \text{ mL}} = 7.50 \times 10^{-3} \text{ M}$ . The pH of the solution will be determined by the hydroxide ion produced by the hydrolysis of the weak base  $\text{Bz}^- + \text{H}_2\text{O} \rightleftharpoons \text{HBz} + \text{OH}^-$ . (Note that this is the back reaction to the neutralization reaction.) To calculate  $[\text{OH}^-]$ , we need the  $K_b$  of  $\text{Bz}^-$ , which can be obtained from the relationship  $K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.46 \times 10^{-5}} = 1.548 \times 10^{-10}$ . As this value indicates, Bz<sup>-</sup> is a very weak base, so we can ignore loss of the analytical concentration from hydrolysis and solve for  $[\text{OH}^-]$  without using the quadratic equation.  $[\text{OH}^-] = \sqrt{C_{\text{Bz}^-} K_b} = \sqrt{(7.50 \times 10^{-3})(1.548 \times 10^{-10})} = 1.077 \times 10^{-6} \text{ M}$  therefore  $\text{pOH} = 5.968$   $\text{pH} = 8.032$  (v) After adding 45.0 mL of 0.0100 M NaOH(aq) This is 15.0 mL beyond the equivalence point. The number of millimoles of added NaOH is  $n(\text{NaOH})_{\text{added}} = (0.0100 \text{ M})(45.0 \text{ mL}) = 0.450 \text{ mmol}$ . At this point our table of millimoles of analyte and titrant before and after the addition now has the following values: ICE Table  $\text{HBz} + \text{NaOH} \rightleftharpoons \text{Bz}^- + \text{H}_2\text{O}$  Initial  $0.300, 0.000$  After Reaction  $0.000, 0.450$  The solution contains 0.150 mmol of  $\text{NaOH}$  beyond what was needed to neutralize the 0.300 mmol of  $\text{HBz}$  initially present. This amount of excess  $\text{OH}^-$  is vastly greater than any amount supplied by the weak base  $\text{Bz}^-$ , and it alone governs the pOH and pH of the solution. In other words, the calculation in this region is identical to the previous strong acid - strong base case. As before, the total volume of the solution at this point is  $V_{\text{total}} = V_{\text{HBz}} + V_{\text{NaOH}} = 10.0 \text{ mL} + 45.0 \text{ mL} = 55.0 \text{ mL}$ . Therefore,  $[\text{OH}^-] = \frac{0.150 \text{ mmol}}{55.0 \text{ mL}} = 2.727 \times 10^{-3} \text{ M}$ .  $\text{pOH} = 2.564$   $\text{pH} = 11.436$ . Consider 25.0 mL of 0.100 M  $\text{HBz}$ , for which  $K_a = 1.00 \times 10^{-5}$ , titrated with 0.0500 M  $\text{NaOH}$  solution. How many milliliters of  $\text{NaOH}$  are needed to reach the equivalence point? How many millimoles of  $\text{HBz}$  are initially present in the sample? What is the initial pH, before adding any titrant? What is the pH after adding 10.0 mL of titrant? What is the pH after adding 25.0 mL of titrant? What is the pH after adding 50.0 mL of titrant? What is the pH after adding 75.0 mL of titrant?

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