Unit 7, Lesson 10: Answers to Titration Calculations

1. Define: titration, primary standard, standardization, equivalence point, endpoint, pipette and burette.

2. The following acid solutions were titrated with 0.150 mol/L sodium hydroxide. Write the neutralization equation for each reaction and calculate the concentrations of the acid solutions.

a) 25.00 mL of HCl requiring 16.50 mL of base solution

$n_{base} = C \times V$

$= 0.150 \text{ mol/L} \times 0.01650 \text{ L}$

$= 0.002475 \text{ mol NaOH}$

= moles of acid at equivalence point

$C_{acid} = n / V$

$= 0.002475 \text{ mol} / 0.02500 \text{ L}$

$= 0.0990 \text{ mol/L}$ (3 sig digs)

or

$C_aV_a = C_bV_b$

$$C_a = \frac{C_bV_b}{V_a}$$

$$= \frac{0.150 \text{ mol/L} \times 0.01650 \text{ L}}{0.02500 \text{ L}}$$

$= 0.0990 \text{ mol/L}$ (3 sig digs)

b) 25.00 mL of sulfuric acid solution requiring 42.00 mL of base solution

Because sulfuric acid is a diprotic acid, and both protons will be titrated, we must take into account the mole ratios (stoichiometry) of the reaction.

Reaction: $H_2SO_4 \text{ (aq)} + 2 \text{ NaOH (aq)} \rightarrow 2 \text{ } H_2O \text{ (l)} + Na_2SO_4 \text{ (aq)}$

$n_{base} = C \times V$

$= 0.150 \text{ mol/L} \times 0.04200 \text{ L}$

$= 0.00630 \text{ mol NaOH}$

1 mole of H_2SO_4 needs 2 moles NaOH

so, moles of acid = ½ x moles NaOH

= 0.00315 moles H_2SO_4

$C_{acid} = n / V$

$= 0.00315 \text{ mol} / 0.02500 \text{ L}$

$= 0.126 \text{ mol/L}$ (3 sig digs)

or

Because one mole of acid requires 2 moles of base, adjust the equation:

$C_aV_a = \frac{1}{2} C_bV_b$

$$C_a = \frac{\frac{1}{2} C_bV_b}{V_a}$$

$$= \frac{\frac{1}{2} \times 0.150 \text{ mol/L} \times 0.04200 \text{ L}}{0.02500 \text{ L}}$$

$= 0.0126 \text{ mol/L}$ (3 sig digs)

c) 10.00 mL of vinegar, containing acetic acid, CH_3COOH requiring 55.0 mL of base solution

$n_{base} = C \times V$

$= 0.150 \text{ mol/L} \times 0.0550 \text{ L}$

$= 0.00825 \text{ mol NaOH}$

= moles of acid at equivalence point

$C_{acid} = n / V$

$= 0.00825 \text{ mol} / 0.01000 \text{ L}$

$= 0.825 \text{ mol/L}$ (3 sig digs)

or

$C_aV_a = C_bV_b$

$C_a = \dfrac{C_bV_b}{V_a}$

$= \dfrac{0.150 \text{ mol/L} \times 0.0550 \text{ L}}{0.01000 \text{ L}}$

$= 0.825 \text{ mol/L}$ (3 sig digs)

3. What volume of 0.350 M ammonium hydroxide will be required to titrate 50.0 mL of 0.275 M HCl? The reaction is:

$$NH_4OH\ (aq) + HCl\ (aq) \rightarrow H_2O\ (l) + NH_4Cl\ (aq)$$

$n_{acid} = C \times V$

$= 0.275 \text{ mol/L} \times 0.0500 \text{ L}$

$= 0.01375 \text{ mol NaOH}$

= moles of base at equivalence point

$V_{base} = n / C$

$= 0.01375 \text{ mol} / 0.350 \text{ L}$

$= 0.0393 \text{ mol/L}$ (3 sig digs)

or 39.3 mL of acid

or

$C_aV_a = C_bV_b$

$C_b = \dfrac{C_aV_a}{V_b}$

$= \dfrac{0.275 \text{ mol/L} \times 0.0500 \text{ L}}{0.350 \text{ L}}$

$= 0.0393 \text{ mol/L}$ (3 sig digs)

or 39.3 mL of acid

4. A solution of HNO_3 of unknown concentration was titrated with 0.948 M KOH. 21.32 mL of the base was required to neutralize a 10.0 mL sample of acid. Find the concentration of the acid. The reaction is:

$$HNO_3\ (aq) + KOH\ (aq) \rightarrow H_2O\ (l) + KNO_3\ (aq)$$

$n_{base} = C \times V$

$= 0.948 \text{ mol/L} \times 0.02132 \text{ L}$

$= 0.020211 \text{ mol NaOH}$

= moles of acid at equivalence point

$C_{acid} = n / V$

$= 0.020211 \text{ mol} / 0.01000 \text{ L}$

$= 2.02 \text{ mol/L}$ (3 sig digs)

or

$C_aV_a = C_bV_b$

$C_a = \dfrac{C_bV_b}{V_a}$

$= \dfrac{0.948 \text{ mol/L} \times 0.02132 \text{ L}}{0.01000 \text{ L}}$

$= 2.02 \text{ mol/L}$ (3 sig digs)

5. 25.8 mL of 0.328 M sodium hydroxide solution are required to titrate 50.0 mL of sulfuric acid. Calculate the concentration of the acid. The reaction is:

$$2\ NaOH\ (aq)\ +\ H_2SO_4\ (aq)\ \rightarrow\ 2\ H_2O\ (l)\ +\ Na_2SO_4\ (aq)$$

$n_{base} = C \times V$

$= 0.328\ mol/L \times 0.0258\ L$

$= 0.008462\ mol\ NaOH$

1 mole of H_2SO_4 needs 2 moles NaOH

so, moles of acid = ½ x moles NaOH

$= 0.004231\ moles\ H_2SO_4$

$C_{acid} = n / V$

$= 0.004231\ mol / 0.05000\ L$

$= 0.0846\ mol/L$ (3 sig digs)

or

Because one mole of acid requires 2 moles of base, adjust the equation:

$C_aV_a = \frac{1}{2} C_bV_b$

$$C_a = \frac{\frac{1}{2} C_bV_b}{V_a}$$

$$= \frac{\frac{1}{2} \times 0.328\ mol/L \times 0.0258\ L}{0.05000\ L}$$

$= 0.0846\ mol/L$ (3 sig digs)

6. 50.0 g of solid NaOH are dissolved in 2.00 L of water. In a titration, a 25.0 mL sample of this solution exactly neutralizes 32.6 mL of hydrochloric acid. What is the concentration of the acid? The reaction is:

$$NaOH\ (aq)\ +\ HCl\ (aq)\ \rightarrow\ H_2O\ (l)\ +\ NaCl\ (aq)$$

First, we have to calculate the concentration of the sodium hydroxide solution:

$n_{NaOH} = mass / MM_{NaOH}$

$$= \frac{50.0\ g}{40.00\ g/mol}$$

$= 1.25\ mol\ NaOH$

$C_{NaOH} = n / V$

$$= \frac{1.25\ mol}{2.00\ L}$$

$= 0.625\ mol/L$

Then, use the calculated concentration of the NaOH solution to do the titration calculation:

$n_{base} = C \times V$

$= 0.625\ mol/L \times 0.0250\ L$

$= 0.015625\ mol\ NaOH$

= moles of acid at equivalence point

$C_{acid} = n / V$

$= 0.015625\ mol / 0.0326\ L$

$= 0.479\ mol/L$ (3 sig digs)

or

$C_aV_a = C_bV_b$

$$C_a = \frac{C_bV_b}{V_a}$$

$$= \frac{0.625\ mol/L \times 0.0250\ L}{0.0326\ L}$$

$= 0.479\ mol/L$ (3 sig digs)

7. A sample of powdered vitamin C (ascorbic acid, $HC_6H_7O_6$) is dissolved in water and titrated with a 0.150 M solution of sodium hydroxide. A total of 21.50 mL of base solution was required to change the colour of the indicator. How many grams of ascorbic acid did the tablet contain, assuming the presence of one acidic hydrogen atom per molecule? The reaction can be written:

$$NaOH\ (aq) + HC_6H_7O_6\ (s) \rightarrow H_2O\ (l) + NaC_6H_7O_6\ (aq)$$

$n_{base} = C \times V$

$= 0.150$ mol/L x 0.02150 L

$= 0.003225$ mol NaOH

= moles of acid at equivalence point

$mass_{acid} = n \times MM$

$= 0.003225$ mol x 176.14 g/mol

$= 0.568$ g of ascorbic acid (3 sig digs)

or 568 mg of ascorbic acid

8. The compound acetylsalicylic acid (ASA), $HC_9H_7O_4$, is found in many pain relievers, such as Aspirin. An ASA product was analyzed by dissolving a tablet weighing 0.250 g in water and titrating with 0.030M KOH. The titration required 29.40 mL of base. What was the percentage by weight of acetylsalicylic acid in the tablet?

$$KOH\ (aq) + HC_9H_7O_4\ (aq) \rightarrow H_2O\ (l) + KC_9H_7O_4\ (aq)$$

$n_{base} = C \times V$

$= 0.030$ mol/L x 0.02940 L

$= 0.000882$ mol KOH

= moles of acid at equivalence point

$mass_{acid} = n \times MM$

$= 0.000882$ mol x 180.17 g/mol

$= 0.159$ g of ASA (3 sig digs)

Most medicines contain the active ingredient (such as ASA) along with other ingredients to stabilize and bind the drug into a tablet. To calculate the percentage by weight of the medicine in the tablet, divide the mass of the drug by the mass of the whole tablet and multiply by 100:

$$\text{\% by mass ASA} = \frac{\text{mass of ASA}}{\text{mass of tablet}} \times 100$$

$$= \frac{0.159\text{ g}}{0.250\text{ g}} \times 100$$

$$= 63.6\text{ \% of the tablet is ASA}$$

9. Maleic acid is a solid, diprotic acid with the chemical formula $H_2C_4O_4$. It was used to standardize a sodium hydroxide solution. It required 32.1 mL of the base to titrate 0.187 g of maleic acid. What is the concentration of the base?

$$2\ NaOH\ (aq) + H_2C_4O_4\ (s) \rightarrow 2\ H_2O\ (l) + Na_2C_4O_4\ (aq)$$

$$n_{acid} = \text{mass} / MM_{acid}$$

$$= \frac{0.187\ g}{114.06\ g/mol}$$

$$= 0.001639\ \text{mol maleic acid}$$

Maleic acid is a diprotic acid, so one mole of acid can neutralize 2 moles of base.

So, the moles of base = 2 x the moles of acid

= 2 x 0.001639 mol

= 0.003279 mol of NaOH

$$C_{NaOH} = n / V$$

$$= \frac{0.003279\ mol}{0.0321\ L}$$

$$= 0.102\ \text{mol/L NaOH solution}$$