The Copperbelt University

CH 110 Tutorial Sheet 3 – September 15, 2015

Reactions in Aqueous Solutions

Question 1

a) Explain the meaning of the following concepts:

i) Solute

ii) Solvent

iii) Solution

Solutions:

*A **solute** is a puresubstance present in small amounts compared to the solvent in a solution**.***

*A **solvent** is the pure substance of a solution that is present in greatest abundance.*

*A **solution**is defined as homogeneous mixture of two or more pure substances.*

b) Identify a solute and a solvent when sodium chloride and water are mixed.

Solution:

***Water** is the solvent and **sodium chloride** is the solute.*

c) From the following list of substances, identify substances that can give you a strong electrolyte and those that can give you a weak electrolyte when in solution. Justify your answer.

Solution:

*i. Sodium chloride – **strong electrolyte, completely ionized in solution or soluble***

*ii. Ammonium chloride – **strong electrolyte, soluble***

*iii. Sulfuric acid – **strong electrolyte, completely ionized in solution***

*iv. Acetic acid – **weak electrolyte, partially ionized in solution***

*v. Ammonium hydroxide – **weak electrolyte, partially ionized in solution***

d) Both aqueous lead (II) bromide and molten Lead are good conductors of electricity, but they are not both electrolytes.

i. Explain this observation.

Solutions:

Lead (II) bromides are soluble and get solvated – they give ionic solutions. Ions conduct electricity

Metals conduct electricitybecause of the movement of electrons in them

ii. Which of the two substances is an electrolyte?

Solution:

Lead bromides is an electrolyte

e) When asked what causes electrolytic solutions to conduct electricity, a student responds that it is due to the movement of electrons through the solution. Is the student correct? If not, what should be the correct response?

Solution:

No. electrolytic solutions conduct electricity because of movement of ions in them.

Question 2

a) Calculate the molarity of the following solutions:

i. A solution of calcium chloride made by dissolving 27g of calcium chloride to make 500ml of solution.

Solution:

$$Molarity = \frac{n}{V} = \left(\frac{m}{M_r}\right) \times \left(\frac{1}{V}\right)$$

where$\mathrm{M_r} = \mathrm{M_{Ca}} + 2 \times \mathrm{M_{Cl}} = 40.08 + 70.90 = 110.98 \text{ g}$

m = 27 g and V = 0.5 L

$$Molarity = \frac{n}{V} = \left(\frac{27}{111}\right) \times \left(\frac{1}{0.5}\right) = \frac{54}{111} = \mathbf{0.49\ M}$$

ii. A solution with 34.6g of sodium chloride dissolved in 125ml of water.

Solution:

$$Molarity = \frac{n}{V} = \left(\frac{m}{M_r}\right) \times \left(\frac{1}{V}\right)$$

$whereM_r = M_{Na} + M_{Cl} = 22.99 + 35.45 = 58.44\ g$

m = 27 g and V = 0.5 L

$$Molarity = \frac{n}{V} = \left(\frac{34.6}{58.44}\right) \times \left(\frac{1}{0.125}\right) = \frac{8 \times 34.6}{58.44} = \mathbf{4.74\ M}$$

b) Explain how you can prepare the following solutions:

i. 100.0ml of $K_2Cr_2O_4$ with a concentration of 1.0M.

NB: Compound should be corrected to $KCrO_4$ or $K_2Cr_2O_7$. I take the latter so that the typo is on 4 in the formula

Solution:

$$C_M = \frac{n}{V}$$

where C_M is the molarity concentration, n is the number of moles and V is the volume of the solution. Since the volume is 100.0 we need to dissolve n moles in 100 ml of water. But the number of moles is

$$n = C_M \times V = 0.1\ L \times 1.0\ M = 0.1\ mol$$

We also know that $n = \frac{m}{M_r}$ *or* $m = n \times M_r$

Where $M_r = 2 \times M_K + 2 \times M_{Cr} + 7 \times M_O = 2 \times (40.08 + 52.00) + 7 \times 16.00$

$$= 2 \times (92.08) + 112.00 = 184.16 + 112.00 = \mathbf{296.16\ g/mol}$$

Thus $m = 0.1\ mol \times 296.16 \frac{g}{mol} = \mathbf{29.616\ g}$

The solution would be prepared by weighing 29.6 g of $K_2Cr_2O_7$ and dissolving it in 100.0 ml of water.

ii. 100.0ml of NaCl solution with a concentration of 0.5M

Solution:

$$C_M = \frac{n}{V}$$

where C_M is the molarity concentration, n is the number of moles and V is the volume of the solution. Since the volume is 100.0 we need to dissolve n moles in 100 ml of water. But the number of moles is

$$n = C_M \times V = 0.1\ L \times 0.5\ M = 0.05\ mol$$

We also know that $n = \frac{m}{M_r}$ *or* $m = n \times M_r$

Where $M_r = M_{Na} + M_{Cl} = 22.99 + 35.45 = 58.44\ g/mol$

Thus $m = 0.05\ mol \times 58.44 \frac{g}{mol} = \mathbf{2.922\ g}$

The solution would be prepared by weighing 2.9 g of NaCl and dissolving it in 100.0 ml of water.

iii. 250.0ml of copper (II)sulfate dihydrate solution with a concentration of 2.0M

Solution:

$$C_M = \frac{n}{V}$$

where C_M is the molarity concentration, n is the number of moles and V is the volume of the solution. Since the volume is 250.0 we need to dissolve n moles in 250 ml of water. But the number of moles is

$$n = C_M \times V = 0.250\ L \times 2.0\ M = 0.50\ mol$$

We also know that $n = \frac{m}{M_r}$ *or* $m = n \times M_r$

Where

$$M_r = M_{Cu} + M_{SO_4^{2-}} + 2 \times M_{H_2O} = 63.55 + 96.07 + 36.032 = 195.65\ g/mol$$

Thus $m = 0.5\ mol \times 195.65 \frac{g}{mol} = \mathbf{97.8\ g}$

The solution would be prepared by weighing 97.8 g of $CuSO_4.H_2O$ and dissolving it in 250.0 ml of water.

c) Calculate the number of moles of the following:

i. 10g of sodium hydroxide

Solution:

We know that the number of moles $n = \frac{m}{M_r}$

Where

$$M_r = M_{Na} + M_{OH^-} = 22.99 + 17.008 = 39.998\ g/mol$$

$Therefore\ n = \frac{m}{M_r} = \frac{10\ g}{40\ g/mol} = \mathbf{0.250\ mol}$

ii. 20 ml of 0.1M sulfuric acid

Solution:

We know that the number of moles $C_M = \frac{n}{V}$

Where

$$C_M = 0.1\ M\ and\ V = 0.02\ L$$

$Therefore\ n = 0.1M \times 0.02L = \mathbf{0.002\ mol\ or\ 2\ millimole}$

Question 3

a) Calculate the concentration of 250ml HCl made by diluting 18.5ml of 2.3M HCl?

Solution:

$M_1V_1 = M_2V_2$ or $M_2 = \frac{M_1V_1}{V_2} = \frac{18.5\ ml \times 2.3M}{250\ ml} = \mathbf{0.17\ M}$

b) A stock solution containing Mn^{2+} ions was prepared by dissolving 1.584g pure manganese metal in nitric acid and diluting to a final volume of 1.000L. The following solutions were then prepared by dilution:

For solution A, 50.00mL of stock solution was diluted to 1000.0mL

For solution B, 10.00mL of stock solution A was diluted to 250.0mL

For solution C, 10.00mL of solution B was diluted to 500.0mL.

Calculate the concentrations of the stock solution, solutions A, B, and C.

Solution:

We first get the concentration of the stock solution using the equation

$$C_{Stock}\ or\ M_{stock} = \frac{n}{V}$$

$Where\ n = \frac{m}{M_r}\ when\ M_r = 54.94 \frac{g}{mol}, m = 1.584\ g\ and\ V = 1.000L$

$Thus\ n = \frac{m}{M_r} = \frac{1.584\ g}{54.94\ g/mol} = 0.0288\ mol\ and\ M_{stock} = \frac{n}{V} = \frac{0.0288}{1.000} = \mathbf{0.0288M}$

Using the concentrations of the new solutions are derived from the equation

$$V_{Stock\ aliquot} \times M_{Stock} = V_{new} \times M_{new}$$

Making the new concentration the subject of the formula we have

$$M_{new} = \frac{V_{Stock\ aliquot} \times M_{Stock}}{V_{new}}$$

We make these calculations using in table below with volumes in mL and concentrations in M

Solution	$V_{Stock\ aliquot}$	M_{Stock}	V_{new}	$M_{new} = \frac{V_{Stock\ aliquot} \times M_{Stock}}{V_{new}}$
A	50.0	0.0288	1000.0	$\frac{50\ mL \times 0.0288\ M}{1000\ mL} = \mathbf{1.44 \times 10^{-3} M}$

B	10.0	0.0288	250.0	$\frac{10\ mL \times \mathbf{1.44 \times 10^{-3} M}}{250\ mL} = \mathbf{5.8 \times 10^{-5} M}$
C	10.0	1.15×10^{-3}	500.0	$\frac{10\ mL \times \mathbf{5.8 \times 10^{-5} M}}{500\ mL} = \mathbf{1.16 \times 10^{-6} M}$

Question 4

a) Predict which of the following substances are likely to be soluble in water:

Solutions:

Use the table in the notes to get answers

i) *Aluminum nitrate -* ***Soluble***

ii) *Lead (II) sulfide -* ***Insoluble***

iii) *Magnesium hydroxide -* ***Insoluble***

iv) *Iron (III) phosphate -* ***Insoluble***

v) *Zinc chloride -* ***Soluble***

vi) *Ammonium carbonate –* ***Soluble***

vii) *Chromium (III) hydroxide –* ***Insoluble***

b) When the following solutions are mixed together, what precipitate (if any) will form?

Solutions:

i) *Aqueous Iron (II) sulfate and aqueous potassium chloride –* ***No precipitate***

ii) *Aqueous aluminum nitrate and aqueous barium hydroxide –* ***Al(OH)₃***

iii) *Aqueous calcium chloride and aqueous sodium sulfate –* ***$CaSO_4$***

iv) *Aqueous potassium sulfide and aqueous nickel (II) nitrate –* ***NiS***

v) *Aqueous carbonate and aqueous magnesium iodide –* ***$MgCO_3$***

c) For each of the reactions in (b) above, write a:

i) balanced formula equation

Solutions:

1. $FeSO_4(aq) + 2KCl(aq) \longrightarrow FeCl_2(aq) + K_2SO_4(aq)$
2. $2Al(NO_3)_3(aq) + 3Ba(OH)_2(aq) \longrightarrow 3Ba(NO_3)_2(aq) + 2Al(OH)_3(s)$
3. $CaCl_2(aq) + Na_2SO_4(aq) \longrightarrow CaSO_4(s) + 2NaCl(aq)$
4. $K_2S(aq) + Ni(NO_3)_2(aq) \longrightarrow 2KNO_3(aq) + NS(s)$
5. $CO_3^{2-}(aq) + MgI_2(aq) \longrightarrow MgCO_3(s) + 2I^-(aq)$

ii) complete ionic equation

Solutions:

1. $Fe^{2+}(aq) + SO_4^{2-}(aq) + 2K^+(aq) + 2Cl^-(aq) \longrightarrow Fe^{2+}(aq) + 2Cl^-(aq) + 2K^+(aq) + SO_4^{2-}(aq)$
2. $2Al^{3+}(aq) + 6NO_3^-(aq) + 3Ba^{2+}(aq) + 6OH^-(aq) \longrightarrow 3Ba^{2+}(aq) + 6NO_3^-(aq) + 2Al(OH)_3(s)$
3. $Ca^{2+}(aq) + 2Cl^-(aq) + 2Na^+(aq) + SO_4^{2-}(aq) \longrightarrow CaSO_4(s) + 2Na^+(aq) + 2Cl^-(aq)$

4. $2K^{+}(aq) + S^{2-}(aq) + Ni^{2+}(aq) + 2NO_3^{-}(aq) \longrightarrow 2K^{+}(aq) + 2NO_3^{-}(aq) + NS(s)$
5. $CO_3^{2-}(aq) + Mg^{2+}(aq) + \ 2I^{-}(aq) \longrightarrow MgCO_3(s) + 2I^{-}(aq)$

iii) net ionic equation

Solutions:

1. $Fe^{2+}(aq) + SO_4^{2-}(aq) + 2K^{+}(aq) + 2Cl^{-}(aq) \longrightarrow Fe^{2+}(aq) + 2Cl^{-}(aq) + 2K^{+}(aq) + SO_4^{2-}(aq)$
2. $2Al^{3+}(aq) + 6OH^{-}(aq) \longrightarrow 2Al(OH)_3(s)$
3. $Ca^{2+}(aq) + \ SO_4^{2-}(aq) \longrightarrow CaSO_4(s)$
4. $S^{2-}(aq) + Ni^{2+}(aq) \longrightarrow NS(s)$
5. $CO_3^{2-}(aq) + Mg^{2+}(aq) \longrightarrow MgCO_3(s)$

d) Separate samples of a solution of an unknown salt are treated with dilute solutions of HBr, H_2SO_4, and NaOH. A precipitate forms in all three cases. Which of the following cations could the solution contain: K^{+}, Pb^{2+}, Ba^{2+}?

Solution: *Pb^{2+}*

e) Name the spectator ions in any reactions that may be involved when each of the following pairs of solutions are mixed:

i) Aqueous sodium carbonate and aqueous magnesium sulfate.

Solution: *Na^{+} and SO_4^{2-}*

ii) Aqueous lead (II) nitrate and aqueous sodium sulfide.

Solution: *Na^{+} and NO_3^{-}*

iii) Aqueous ammonium phosphate and aqueous calcium chloride.

Solution: *NH_4^{+} and Cl^{-}*

f) A 100.0ml aliquot of 0.200M aqueous potassium hydroxide is mixed with 100.0ml of 0.200M aqueous magnesium nitrate.

i) Write a balanced chemical equation for any reaction that occurs.

Solution:

$$\mathbf{2KOH(aq) + Mg(NO_3)_2(aq) \longrightarrow Mg(OH)_2(s) + 2KNO_3(aq)}$$

ii) What precipitate is formed from this reaction?

Solution:

$$\mathbf{Mg(OH)_2(s)}$$

iii) Calculate the mass of the precipitate formed.

Solution:

Mole ratios: 2 moles of KOH produce a mole of 1 of $Mg(OH)_2$

Moles of KOH used up in this reaction is

$$M_{KOH} \times V_{KOH} = 0.1L \times 0.2M = 0.020\ mol$$

Therefore 0.010 mol$Mg(OH)_2$ is formed. The mass formed is given by the formula $m = n \times M_r$ since

$M_r = M_{Mg} + 2 \times M_{OH} = 24.31 + 2 \times 17.008 = 58.33\ g/mol,$

$$m = n \times M_r = 0.010\ mol \times 58.33 \frac{g}{mol} = \mathbf{0.58\ g}$$

iv) Calculate the concentration of each ion remaining in solution after precipitation is complete.

Solution:

The moles of the magnesium nitrate is remaining in solution is 0.01 mol, that is 0.01 mol of Mg^{2+} ion and 0.02 mol of NO_3^-. A similar amount was used to make Potassium nitrate. The total volume of the solution is 0.2 L. Thus the concentrations are

$$\boldsymbol{NO_3^-} = \frac{\mathbf{0.04\ mol}}{\mathbf{0.2\ L}} = \mathbf{0.20\ M}\ and\ \boldsymbol{Mg^{2+}} = \frac{\mathbf{0.01\ mol}}{\mathbf{0.2\ L}} = \mathbf{0.05M};$$

$$\boldsymbol{K^+} = \frac{\mathbf{0.02\ mol}}{\mathbf{0.2\ L}} = \mathbf{0.10\ M}$$

Question 5

a) Calculate the oxidation state for the underlined atom in each of the following:

Solution:

i) K<u>Mn</u>$O_4 \Longrightarrow 1 + x - 8 = 0; \boldsymbol{x = 7}$

ii) Sr<u>Cr_2</u>$O_7 \Longrightarrow 2 + 2x - 14 = 0; \boldsymbol{x = 6}$

iii) <u>Cr_2</u>$O_7^{2-} \Longrightarrow 2x - 14 = -2; \boldsymbol{x = 6}$

iv) <u>Pb</u>$SO_3 \Longrightarrow 2x - 14 = -2; \boldsymbol{x = 6}$

v) Na<u>Bi</u>$O_3 \Longrightarrow 1 + x - 6 = 0; \boldsymbol{x = 5}$

vi) $(NH_4)_2$<u>Ce</u>$(SO_4)_3 \Longrightarrow 2 + x - 6 = 0; \boldsymbol{x = 4}$

vii) <u>Cr_2</u>$O_3 \Longrightarrow 2x - 6 = 0; \boldsymbol{x = 3}$

viii) Na_4<u>Fe</u>$(OH)_6 \Longrightarrow 4 + x - 6 = 0; \boldsymbol{x = 2}$

ix) <u>N</u>$H_4^+ \Longrightarrow x + 4 = 1; \boldsymbol{x = -3}$

x) <u>P</u>$O_4^{3-} \Longrightarrow x - 8 = -3; \boldsymbol{x = 5}$

b) Specify which of the following are Redox reactions. For the Redox reactions, identify the reducing agent, the oxidizing agent, the substance oxidized, and the substance reduced.

i) $2H^+(aq) + 2CrO_4^{2-}(aq) \rightarrow Cr_2O_7^{2-}(aq) + H_2O(l)$

Solution:

Not a redox reaction!

ii) $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Solution:

Redox reaction; H^+ is the oxidizing agent, Zn is oxidized, Zn is the reducing agent and hydrogen is reduced.

iii) $CH_4(g) + H_2O(g) \rightarrow CO(g) + 3H_2(g)$

Solution:

Redox reaction; Hydrogen in water is the oxidising agent, it oxidises carbon in methane. Carbon is the reducing agent it reduces hydrogen.

iv) $2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

Redox reaction; Cu oxidized, oxidizing agent is Ag^{+}, Cu is the reducing agent, it reduces Ag^{+}

c) Balance the following Redox reaction equations using oxidation numbers:

i) $Co^{3+}(aq) + Ni(s) \rightarrow Co^{2+}(aq) + Ni^{2+(}aq)$

Solution:

Step 1.Unbalanced equation is:$Co^{3+}(aq) + Ni(s) \rightarrow Co^{2+}(aq) + Ni^{2+(}aq)$

Step 2.We start by writing the oxidation numbers of each specie in the equation

$Co^{3+}(aq) + Ni(s) \rightarrow Co^{2+}(aq) + Ni^{2+(}aq)$

+3 0 +2 +2

Step 3. We show how electrons are gained and lost

2 e^{-} lost

$Co^{3+}(aq) + Ni(s) \rightarrow Co^{2+}(aq) + Ni^{2+(}aq)$

1 e^{-}gained

Step 4. Show coefficients needed to equalize the electrons gained and lost

2 e^{-} lost

$Co^{3+}(aq) + Ni(s) \rightarrow Co^{2+}(aq) + Ni^{2+(}aq)$

1 e^{-}gained MULTIPLY BY 2 to get equation below

$$\mathbf{2Co^{3+}(aq) + Ni(s) \longrightarrow 2Co^{2+}(aq) + Ni^{2+}(aq)}$$

Step 5. All elements are balanced therefore the above equation is the required balanced equation

ii) $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

Solution:

Step 1. Unbalanced equation is: $Zn(s) + H_2SO_4\,(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

Step 2. We start by writing the oxidation numbers of each specie in the equation

$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

0 +1+2 0

Step 3. We show how electrons are gained and lost

2 e^{-}gained

$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

2e^{-}lost

Step 4.There are no coefficients needed to equalize the electrons gained & lost

2 e^{-}gained

$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

2e^{-}lost

Step 5. All elements are balanced therefore the equation below is the required balanced equation

$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

iii) $Cl_2(g) + Al(s) \rightarrow Al^{3+}(aq) + Cl^-(aq)$

Solution:

Step 1. Unbalanced equation is: $Cl_2(g) + Al(s) \rightarrow Al^{3+}(aq) + Cl^-(aq)$

Step 2. We start by writing the oxidation numbers of each specie in the equation

$Cl_2(g) + Al(s) \rightarrow Al^{3+}(aq) + Cl^-(aq)$

0 0 +3 -1

Step 3. We show how electrons are gained and lost

3e⁻ lost

$Cl_2(g) + Al(s) \rightarrow Al^{3+}(aq) + 2Cl^-(aq)$

2e⁻ gained

Step 4.Show arecoefficients needed to equalize the electrons gained and lost

3e⁻ lost MULTIPY BY 2

$Cl_2(g) + Al(s) \rightarrow Al^{3+}(aq) + 2Cl^-(aq)$

2e⁻ gained MULTIPY BY 3 to give the equation below

$3Cl_2(g) + 2Al(s) \rightarrow 2Al^{3+}(aq) + 6Cl^-(aq)$

Step 5. All elements are balanced therefore the equation aboveis the required balanced equation.

iv) $H^+(aq) + MnO_4^-(aq) + Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq) + H_2O(l)$

Solution:

Step 1. Unbalanced equation is:

$H^+(aq) + MnO_4^-(aq) + Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq) + H_2O(l)$

Step 2. We start by writing the oxidation numbers of each specie in the equation

$H^+(aq) + MnO_4^-(aq) + Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq) + H_2O(l)$

+7 +2 +2 +3

Step 3. We show how electrons are gained and lost

5e⁻ gained

$H^+(aq) + MnO_4^-(aq) + Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq) + H_2O(l)$

1e⁻ lost

Step 4.Show arecoefficients needed to equalize the electrons gained and lost

5e⁻ gained

$H^+(aq) + MnO_4^-(aq) + Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq) + H_2O(l)$

1e^{-} lost MULTIPLY BY 5

$H^+(aq)$ + $MnO_4^-(aq)$ + $5Fe^{2+}(aq)$ → $Mn^{2+}(aq)$ + $5Fe^{3+}(aq)$ + $H_2O(l)$

Step 5.Show coefficients needed to balance the remaining elements.

Balance O:

$H^+(aq)$ + $MnO_4^-(aq)$ + $5Fe^{2+}(aq)$ → $Mn^{2+}(aq)$ + $5Fe^{3+}(aq)$ + $4H_2O(l)$

Balance H:

$8H^+(aq)$ + $MnO_4^-(aq)$ + $5Fe^{2+}(aq)$ → $Mn^{2+}(aq)$ + $5Fe^{3+}(aq)$ + $4H_2O(l)$

All elements are now balanced.

Question 6

a) Write the balanced formula equation for the following acid-base reactions:

i) Potassium hydroxide (aqueous) and nitric acid

Solution:

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

ii) Barium hydroxide (aqueous) and hydrochloric acid.

Solution:

$$Ba(OH)_2(aq) + 2HCl(aq) \longrightarrow 2H_2O(l) + BaCl_2(aq)$$

iii) Perchloric acid (aqueous) and solid iron (III) hydroxide

Solution:

$$Fe(OH)_3(aq) + 3HClO_4(aq) \longrightarrow 3H_2O(l) + Fe(ClO_4)_3(aq)$$

b) What volume of 0.0200 M calcium hydroxide is needed to neutralize 35.00 ml of 0.0500 M nitric acid?

Solution:

The reaction is

$$Ca(OH)_2(aq) + 2HNO_3(aq) \longrightarrow 2H_2O(l) + Ca(NO_3)_2(aq)$$

A mole of calcium hydroxide neutralizes 2 moles nitric acid.

The moles of Nitric Acid is $n=M_1xV_1$=35 x 0.05 mmol = 3.5x0.5 = 1.75 mmol

The moles of Calcium Hydroxide required is 1.75/2=0.875 mmol $=M_2V_2$

Since M_2=0.0200 M, $V_2 = \frac{n}{M_2} = \frac{8.75\times10^{-4}}{2.00\times10^{-2}}L = 4.375 \times 10^{-2} L = 43.75\text{ ml}$

c) What volume of each of the following bases would react completely with 25.00 mL of 0.200 M HCl?

i) 0.100 M NaOH

ii) 0.050 M $Ba(OH)_2$

iii) 0.250 M KOH

Solution:

Moles of HCl = 25.00 mL x 0.2 M = 2.5 x 2 mmol = 5 mmol.

(i) The reaction of NaOH with HCl is

$NaOH+HCl \rightarrow H_2O+NaCl$

We require 5 mmol of solution NaOH, that is 5mmol= V_2 x 0.100 M

$$V_2 = \frac{n}{M_2} = \frac{5 \times 10^{-3}}{1.00 \times 10^{-1}} L = 5 \times 10^{-2} L = 50\text{ml}$$

(ii) The reaction of $Ba(OH)_2$ with HCl is

$Ba(OH)_2+2HCl \rightarrow 2H_2O+BaCl_2$

We require 2.5 mmol of solution $Ba(OH)_2$, i.e., 2.5 mmol= V_2 x 0.050 M

$$V_2 = \frac{n}{M_2} = \frac{2.5 \times 10^{-3}}{5.00 \times 10^{-2}} L = 5 \times 10^{-2} L = 50\text{ ml}$$

(iii) The reaction of KOH with HCl is

$KOH+2HCl \rightarrow H_2O+KCl$

We require 5 mmol of solution KOH, i.e., 5 mmol= V_2 x 0.250 M

$$V_2 = \frac{n}{M_2} = \frac{5 \times 10^{-3}}{2.50 \times 10^{-1}} L = 2 \times 10^{-2} L = 20\text{ ml}$$