# Set 2. Acidity

## Multiple Choice Questions

1. Calcium hydroxide is a strong base. Calculate the $[Ca^{2+}]$ and $[OH^-]$ for a solution prepared by dissolving 0.60 g of $Ca(OH)_2$ in enough water to make 1.50 L of solution.

   (a) $5.4 \times 10^{-3}$; $1.1 \times 10^{-2}$ (b) $5.4 \times 10^{-3}$; $5.4 \times 10^{-3}$
   (c) $5.4 \times 10^{-3}$; $9.1 \times 10^{-13}$ (d) $8.1 \times 10^{-3}$; $8.1 \times 10^{-3}$
   (e) $8.1 \times 10^{-3}$; $1.6 \times 10^{-2}$

2. What is the conjugate acid of $CH_3NH_2$?

   (a) $CH_3NH^-$ (b) $CH_3NH^+$ (c) $CH_3NH_3^+$ (d) $CH_3NH_2^-$
   (e) $CH_3NH^{2+}$

3. What is the conjugate acid of $H_2SO_4$?

   (a) $SO_4^{2-}$ (b) $H_2SO_4^-$ (c) $H_3SO_4^+$ (d) $HSO_4$
   (e) $HSO_4^+$

4. What is the pH of human muscle fluid with a hydronium ion concentration of $1.6 \times 10^{-7}$ mol $L^{-1}$?

   (a) 7.20 (b) 7.16 (c) 7.80 (d) 6.60
   (e) 6.20

5. The hydrogen ion concentration of the oceans is about $7.1 \times 10^{-9}$ mol $L^{-1}$. What is the pH of ocean water?

   (a) 5.85 (b) 7.15 (c) 8.15 (d) 8.85
   (e) 9.71

6. A brand of vinegar has a hydroxide ion concentration of $1.3 \times 10^{-12}$ mol $L^{-1}$. What is the pH of the vinegar?

   (a) 3.11 (b) 2.89 (c) 3.88 (d) 2.11
   (e) 11

7. What is the pH of a 0.012 mol $L^{-1}$ solution of calcium hydroxide?

   (a) 12.00 (b) 1.62 (c) 11.10 (d) 12.38
   (e) 12.20

8. The pH of a 0.10 mol $L^{-1}$ solution containing 0.10 mol $L^{-1}$ $NH_4Cl$ is 9.20. What is the concentration of the $H_3O^+$ ions in it?

   (a) $2.0 \times 10^{-9}$ (b) $6.3 \times 10^{-10}$ (c) $1.7 \times 10^{-10}$ (d) $1.0 \times 10^{-1}$
   (e) $1.6 \times 10^{-5}$

9. The pH of milk of magnesia, ($Mg(OH)_2$) is 10.50. What is the concentration of $OH^-$ ions in it?

(a) $3.2 \times 10^{-3}$ (b) $5.0 \times 10^{-10}$ (c) $3.2 \times 10^{-11}$ (d) $5.0 \times 10^{-11}$
(e) $3.2 \times 10^{-4}$

10. In the reaction, $HCO_3^-(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + H_2CO_3(aq)$, the conjugate base is

(a) $HCO_3^-$ (b) $H_2O$ (c) $OH^-$ (d) $H_2CO_3$

(e) none of these

11. The $K_a$ values for HF and $HNO_2$ are $6.8 \times 10^{-4}$ and $4.5 \times 10^{-4}$ respectively. Therefore, it follows that HF is a ________ acid than $HNO_2$ and $F^-$ is a________ base than $NO_2^-$.

(a) weaker, stronger (b) weaker, weaker

(c) stronger, weaker (d) stronger, stronger

12. Which of the following metals can react with a base and produce hydrogen?

1. Magnesium 2. Chromium 3. Lead 4. Sodium 5. Tin

(a) 1,2,3 (b) 2,3 (c) 4,5 (d) 3,4 (e) 1,3,5

13. What is the concentration of $OH^-$ ions in a neutral solution of water at 37°C where $K_w$ is $2.5 \times 10^{-14}$?

(a) $1.5 \times 10^{-8}$ (b) $2.6 \times 10^{-8}$ (c) $2.6 \times 10^{-7}$

(d) $1.6 \times 10^{-7}$ (e) $1.0 \times 10^{-7}$

14. What is the concentration of a $Ba(OH)_2$ solution that has a pH 9.30?

(a) $1.00 \times 10^{-5}$ mol $L^{-1}$ (b) $2.00 \times 10^{-5}$ mol $L^{-1}$

(c) $2.50 \times 10^{-10}$ mol $L^{-1}$ (d) $5.01 \times 10^{-10}$ mol $L^{-1}$

15. How much water must be added to 200 mL of a 0.010 M solution of HCl to raise the pH to 2.5?

(a) 232 mL (b) 332 mL (c) 432 mL

(d) 532 mL (e) 632 mL

## Longer Questions

1. What is the relationship between the strength of an acid and its numerical value of $K_a$?

2. How will the following ions react with water? Give the equation, where applicable:
   $I^-$, $NO_2^-$, $K^+$, $HSO_4^-$, $CO_3^{2-}$?

3. Predict whether aqueous solutions of the following salts would be acidic, alkaline, or close to neutral: KCl, $NH_4NO_2$, Na(HCOO).

4. Predict whether aqueous solutions of the following salts would be acidic, alkaline or close to neutral: $NH_4CN$, KI, $LiCH_3COO$.

5. Identify the acids and bases according to the Brønsted-Lowry theory of acids and bases in the following
   (a) $HNO_{3}(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NO_3^-(aq)$

   (b) $HCl(g) + NH_3(g) \rightleftharpoons NH_4^+(aq) + Cl^-(aq)$

   (c) $H_2SO_4(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HSO_4^-(aq)$

   (d) $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

6. Write equations to show how a solution of $CO_2$ becomes acidic and $Na_2O$ in water becomes basic.

7. (a) What is meant by hydrolysis?

List three ions that hydrolyse to give

(b) an acidic solution

(c) a basic solution.

8. If 32.0 mL of a 0.1M $H_2SO_4$ is required to precipitate all the barium ions from a solution of $BaCl_2$, what mass of $BaCl_2$ was present in the solution?

9. Calculate the pH of the following solutions

(a) 0.10 M HBr

(b) 0.00001 M $HNO_3$

(c) 0.010 M KOH

(d) 0.01 M $Ba(OH)_2$

(e) 0.0172 M HCl

10. Calculate the final pH when the following are added to 1.00 L of 0.100 M HCl solution. (Assume there is no change in volume).

(a) 0.010 mol of KOH

(b) 0.050 mol of $Ba(OH)_2$

(c) 0.100 mol of HCl

11. Determine the pH of the solution formed by mixing equal volumes of the two solutions in each case.

(a) 0.10 M HCl and 0.10 M NaOH

(b) 0.20 M HCl and 0.10 M NaOH

(c) 0.40 M HCl and 0.20 M NaOH

(d) 0.10 M HCl and 0.10 M $Ba(OH)_2$

12. A 1.50 litre water solution contains 1.7 micrograms of pure HCl (1 microgram = $10^{-6}$ grams).

(i) Calculate the concentration of HCl in ppm.

(ii) Calculate the concentration of hydrogen ions in the solution.

(iii) The calculated pH of this solution appears to be 7.50 which indicates a basic solution of an acid - but this cannot be correct. What is wrong with the assumption here?

13. If 4.90 g of sulfuric acid is added to water so that the final volume is 100 mL, what will be the pH of the solution, if 1 mole of $H_2SO_4$ gives 1.3 moles of $H_3O^+$?

14. If 25.0 mL of 0.200 M sodium hydroxide solution is added to 30.0 mL of 0.175 mol $L^{-1}$ sulfuric acid, what is the pH of the mixture?

15. 200 mL of 0.0500 mol $L^{-1}$ barium hydroxide solution is mixed with 400 mL of 0.200 M nitric acid. The mixture is then diluted with water so that the final volume is 6.00 L.

What is the pH of the final solution?

16. A 10.0 mL sample of 0.005 mol $L^{-1}$ calcium hydroxide is diluted with water to make 1.0 L.

(a) What is the pH of the original solution?

(b) How does the pH change due to the dilution?

(c) What mass of calcium hydroxide is present in the undiluted and the diluted solution?

(d) What volume of $CO_2$ at 25°C and 110 kPa pressure must be bubbled through the solution above in order to precipitate all the $Ba^{2+}$ ions?

17. Explain using equations, why aqueous solutions of sodium carbonate, sodium sulfide and sodium ethanoate all have a pH value greater than 7.

18. A solution has a pH of 6. If you dilute it to 100 times its original volume its pH only changes to 8. Explain this.

19. Ammonium carbonate dissolves in water as follows:

$(NH_4)_2CO_3(s) \rightarrow 2NH_4^+(aq) + CO_3^{2-}(aq)$ (1)

Following dissolution, two further reactions occur as follows:

$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$ (2)

$CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$ (3)

(a) In what way do reactions 2 and 3 differ from 1? (What kind of reactions?)

(b) Indicators show that an aqueous solution $(NH_4)_2CO_3$ is basic, and pH tests confirm this. $NH_4^+$ ions are acidic and $CO_3^{2-}$ ions are basic in solution.

Explain how a salt can produce a solution that is basic.

20. Suggest a reason why $H_2PO_4^-$ is a weaker acid than $H_3PO_4$.

but for $H_2SO_4$ result is about $0.1 \times 1.3 = 1.3$, so less $H^+$ and higher pH.

6. Answer is D.
7. Acidic compounds are $CO_2$, $SO_3$, $OCl_2$ and $NO_2$ (non-metal oxides).
8. $Ba(OH)_2$ is a strong base with 3 ions per formula unit.
9. $OH^-$, $HCO_3^-$, $NH_3$ (all can gain or lose a proton).
10. (i) $CO_2$ dissolves in water to produce $H_2CO_3$ which reacts with water:
    $H_2CO_3 + H_2O \rightarrow HCO_3^- + H_3O^+$ (acidic pH < 7). When boiled, this reaction reverses the produce neutral water again and pH 7.
    (ii) The K value for water goes up because the reaction is endothermic so more $H^+$ ions.
    (iii) Although there are more $H^+$ ions if the pH is smaller, there is still an equal number of OH- ions so it is still neutral. N.B. The pH of water is 7 only at 25°C.
11. (ii) and (v) are not Brønsted–Lowry reactions.
12. (i) The ammonia solution must have a much higher concentration than the HCl as it a weak base and a 0.05 M solution would not supply enough $OH^-$ ions to neutralise it.
    (ii) $NH_3$ ions would be most common as it "wants to" stay as a molecule.
13. A Brønsted–Lowry acid is a species that donates a $H^+$ ion. To react with water, a proton must be exchanged which means the reactant must be a Brønsted–Lowry acid or base.
14. $NH_3 + H^+ \rightarrow NH_4^+$
    $H_2O + H^+ \rightarrow H_3O^+$
    $HCO_3^- + H^+ \rightarrow H_2CO_3$
    $CO_3^{2-} + H^+ \rightarrow HCO_3^-$
    $CN^- + H^+ \rightarrow HCN$
    $OH^- + H^+ \rightarrow H_2O$
15. $HCO_3^- + H_2O \rightarrow H_3O^+ + CO_3^{2-}$
    $HCO_3^- + H_3O^+ \rightarrow H_2CO_3 + H_2O$
    $CO_3^{2-}$ is the conjugate base of $HCO_3^-$ and $H_2CO_3$ is the conjugate acid of $HCO_3^-$.
16. Acids: $NH_4^+$, $H_2O$, HBr Bases: $SO_3^{2-}$, $AlCl_3$, $Cl^-$
17. (i) HCl, HCN, $HS^-$, $NH_4^+$
    (ii) $O^{2-}$, $HSe^-$, $NH_2^-$, $CO_3^{2-}$

## Set 2 Acidity

### Multiple Choice Answers

1. a, 2. c, 3. c, 4. d, 5. c, 6. d, 7. d, 8. b, 9. e, 10. a, 11. c, 12. b, 13. d. 14. a, 15. c.

### Answers to Longer Questions

1. Acid strength increases with increasing $K_a$, acid strength decreases with decreasing $K_a$.
2. $I^-$ ion is the conjugate base of a strong acid. It will simply be hydrated and stay as $I^-$.
   $NO_2^-$ ion is the conjugate base of a weak acid, $HNO_2$, and would react to form a base.
   $NO_2^-(aq) + H_2O(l) \rightleftharpoons HNO_2(aq) + OH^-(aq)$.
   $K^+$ ion does not react with water except to get hydrated because the ions stay as $K^+$.
   $CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$
   $HSO_4^- + H_2O \rightleftharpoons SO_4^{2-} + H_3O^+$
3. KCl is the salt of the cation $K^+$, which is a weaker acid than water and the anion $Cl^-$, which is a weaker base than water. Because neither ion reacts with water to any extent, a KCl solution will be neutral.
   $NH_4NO_2$ is the salt of a cation and anion both of which can react with water. The solution will be nearly neutral. The pH will depend on the relative degree of ionisation of the anion and cation in water.
   Na(HCOO) is the salt of a cation that does not react with water and an anion that is stronger base than water (because it is the anion of a weak acid). The solution will therefore be basic.
4. $NH_4CN$: close to neutral, because both the parent acid and base are weaker than water
   KI: close to neutral, because both the parent acid and base are stronger than water.
   $LiCH_3COO$: alkaline since the parent base $Li(OH)_2$ is a stronger base than water.
5. (a) Acids: $HNO_3$, $H_3O^+$ Bases: $H_2O$, $NO_3^-$
   (b) Acids: HCl, $NH_4^+$ Bases: $NH_3$, $Cl^-$
   (c) Acids: $H_2SO_4$, $H_3O^+$ Bases: $H_2O$, $HSO_4^-$
   (d) Acids: $H_2O$, $H_3O^+$ Bases: $H_2O$, $OH^-$
6. $CO_2 + H_2O \rightarrow H_2CO_3$, and, $H_2CO_3 + H_2O \rightarrow H_3O^+ + HCO_3^-$
   $Na_2O + H_2O \rightarrow 2Na^+ + 2OH^-$
7. (a) Hydrolysis is a reaction in which a salt reacts with water to produce a $H^+$ ($H_3O^+$) or an $OH^-$ ion. Acidic salts give $H_3O^+$ ions and basic salts give $OH^-$ ions with water.
   (b) Cations of the following acidic salts give $H_3O^+$ ions:
   $Fe^{3+} + 2H_2O \rightarrow Fe(OH)^{2+} + H_3O^+$
   $NH_4^+ + H_2O \rightarrow NH_3 + H_3O^+$
   $HSO_4^- + H_2O \rightarrow SO_4^{2-} + H_3O^+$
   (c) Anions of the following basic salts give $OH^-$ ions:
   $NH_2^- + H_2O \rightarrow NH_3 + OH^-$
   $CO_3^{2-} + H_2O \rightarrow HCO_3^- + OH^-$
   $PO_4^{3-} + H_2O \rightarrow HPO_4^{2-} + OH^-$
8. $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$
   1 mol of $H_2SO_4$ reacts with 1 mole of $BaCl_2$.
   $\rightarrow$ mass of $BaCl_2 = [0.1 \times (137.3 + 70.9) \times 0.032] = 0.667$ g.

9. All are strong acids and completely dissociated.
(a) $[H^+] = 1 \times 10^{-1}$ M $\therefore$ pH = 1.0
(b) $[H^+] = 1 \times 10^{-5}$ M $\therefore$ pH = 5.0
(c) $[OH^-] = 1 \times 10^{-2}$ M $\therefore$ $[H^+] = 1 \times 10^{-12}$ M $\therefore$ pH = 12.0
(d) $[OH^-] = 2 \times 10^{-2}$ M $\therefore$ $[H^+] = 5 \times 10^{-13}$ M $\rightarrow$ pH = 12.3
(e) $[H^+] = 1.72 \times 10^{-2}$ M $\therefore [H^+] = 1 \times 10^{-1.76}$ M $\therefore$ pH = 1.76

10. 1.0 L of 0.1 M HCl contains 0.1 mol of HCl.
(a) HCl left after reaction with 0.010 mol of KOH = 0.09 mol
$[H^+] = 0.09$ mol $L^{-1} = 9.0 \times 10^{-2}$M,
$\therefore$ pH = $-\log 9 \times 10^{-2} = 1.05$
(b) $[OH^-] = 2 \times 0.05 = 0.1$ mol $\therefore$ HCl left after reaction is none $\therefore$pH = 7.
(c) $[OH^-] = 0.1$ mol $\therefore$ HCl left after reaction is none $\therefore$ pH = 7

11. Assume 0.5 L of each solution is mixed.
(a) Since $[H^+] = [OH^-]$ in the mixture, the pH is 7.
(b) Before reaction, there are 0.20 × 0.50 (0.1 mol) of HCl, and 0.10 × 0.500 = (0.05 mol) of NaOH.
After the reaction 0.05 mole of HCl is left. $\therefore$ $[H^+] = 0.05/1.0 = 0.05$ mol $L^{-1}$
$\therefore$ pH= $-\log 0.05 = 1.30$
(c) Before reaction, there are 0.40 × 0.50 (0.2 mol) of HCl, and 0.20 × 0.500 = (0.10 mol) of NaOH.
After the reaction 0.10 mole of HCl is left. $\therefore$ $[H^+] = 0.10/1.0 = 0.1$ mol $L^{-1}$
$\therefore$ pH= $-\log 0.1 = 1.00$
(d) Before reaction, there are 0.10 × 0.50 (0.05 mol) of HCl, and 0.20 × 0.50 = (0.10 mol) of $OH^-$ ions.
After the reaction 0.05 mole of $OH^-$ ions left.
$\therefore$ $[OH^-] = 0.05/1.0 = 0.05$ mol $L^{-1}$
$\therefore$ pOH= $-\log 0.05 = 1.30$
so pH = 14 − 1.3 = 12.7.

12.
(i) $1.7 \times 10^{-6}$ g in 1500 g of solution
$= 1.13 \times 10^{-9}$ grams per gram
$1.13 \times 10^{-9}$ multiply by a million
$= 1.13 \times 10^{-3}$ ppm.

(ii) $n(HCl) = \dfrac{1.7 \times 10^{-6}}{36.45} = 4.664 \times 10^{-8}$ mol

$[H^+] = \dfrac{4.664 \times 10^{-8}}{1.5} = 3.109 \times 10^{-8}$ mol $L^{-1}$

(iii) Calculated pH = $-\log(3.109 \times 10^{-8})$ = 7.50 but, this concentration of $H^+$ ions is lower than that for neutral water and so the contribution from water cannot be disregarded.
Total concentration of $H^+ = 3.109 \times 10^{-8}$ (acid) + $1 \times 10^{-7}$ (water) = $1.31 \times 10^{-7}$ mol $L^{-1}$
pH = $-\log(1.31 \times 10^{-7}) = 6.88$ (acidic!)

13. $n(H_2SO_4) = (4.90/ 98.086) = 0.050$ mol,
$\therefore [H_2SO_4] = (0.050/0.100) = 0.50$ M
$\therefore [H^+] = 1.3 \times 0.5$ M = 0.65 M
$\therefore$pH = $-\log [0.65] = 0.187$

14. $2\,NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$
n (NaOH) = 0.025 × 0.20 = 0.0050 mol (limiting reagent)
$n(H_2SO_4)$ required to use up all the NaOH = ½ × 0.005 = 0.0025 mol
For $H_2SO_4$, $n(H_2SO_4) = 0.030 \times 0.175$ = 0.00525 mol (excess reagent)
$H_2SO_4$ is left over and is in excess by (0.00525 − 0.0025) = 0.00275 mol
$\therefore$ Final $[H_2SO_4] = \dfrac{0.00275}{0.05} = 0.055$ M
$\therefore$ Final $[H^+] = 2 \times 0.055 = 0.959$ M
$\therefore$ pH = $-\log [0.182] = 0.74$

15. $Ba(OH)_2 + 2HNO_3 \rightarrow Ba(NO_3)_2 + 2H_2O$
For $Ba(OH)_2$, n = 0.050 × 0.200 = 0.010 mol = (limiting reagent)
For $HNO_3$, n = 0.200 × 0.400 = 0.080 mol = (excess reagent)
$HNO_3$ is left over and is in excess.
$\therefore$ $H^+$ is in excess by 0.080 mol − 0.020 = 0.060 mol.
After dilution $[H^+]$= 0.06/6.0 = $1 \times 10^{-2}$ M
pH = 2.

16. (a) $[OH^-] = 2 \times 0.00500 = 0.0100$ M
$= 1.00 \times 10^{-2}$M
$\therefore$ pOH = 2 pH = 14 − 2 = 12
$\therefore$ The pH of the original solution = 12.
(b) After making the final volume 1.0 L, the final concentration is calculated using the dilution relationship, $c_1V_1 = c_2V_2$.
0.010 L × 1.0 × $10^{-2}$ M = 1.99 L × $c_2$
$\therefore$ $c_2 = [(0.010 \times 1.0 \times 10^{-2})/ 1.0\ L]$
$= 1.0 \times 10^{-4}$ M
$\therefore$ $[OH^-] = 1.0 \times 10^{-4}$ M, and $[H^+]$
$= 1.0 \times 10^{-10}$ M
$\therefore$ pH = 10.
pH changes from 12 to 10.
(c) $n(Ca(OH)_2 = 0.00500 \times 0.010$
$= 5.00 \times 10^{-5}$ mol
$m\,[(Ca(OH)_2] = 5.00 \times 10^{-5} \times 74.09$
= 0.00370 g.
(d) $Ca(OH)_2 + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$
$n\,(CO_2) = n[(Ca(OH)_2] = 5.00 \times 10^{-5}$ mol
Using the relationship pV = nRT and V = nRT ÷ p, V = [(5.00 × $10^{-5}$ × 8.314 × 298)/ (110)] = 1.126 × $10^{-3}$ L = 1.126 mL.

17. $CO_3^{2-} + H_2O \rightarrow HCO_3^- + OH^-$ ($CO_3^{2-}$ is a stronger base than $H_2O$)
$S^{2-} + H_2O \rightarrow HS^- + OH^-$ ($S^{2-}$ is a stronger base than $H_2O$)

$CH_3COO^- + H_2O \rightarrow CH_3COOH + OH^-$
($CH_3COO^-$ *is a stronger base than* $H_2O$)

18. *pH of 6 means* $[H^+] = 1 \times 10^{-6}$ *mol* $L^{-1}$ *and a pH of 8 means* $[H^+] = 1 \times 10^{-8}$ *mol* $L^{-1}$ *which is 100 times less than the first concentration.*
19. (a) $(NH_4)_2CO_3(s) + water \rightarrow 2NH_4^+(aq) + CO_3^{2-}(aq)$
*This reaction (reaction 1) is an ionisation reaction.*
$NH_4^+(aq)$ *produces* $H_3O^+$ *ions (reaction 2). It is from an acidic salt. Anions of basic salts produce* $H_3O^+$ *ions.*
$CO_3^{2-}(aq)$ *produces* $OH^-$ *ions (reaction 3). It is from a basic salt. Cations of acidic salts produce* $OH^-$ *ions.*
*Reactions 2 and 3 are hydrolysis reactions.*
$NH_4^+$ *ion is a weak acid. It does not produce many* $H_3O^+$ *ions. Although the carbonate ion is in the minority, it has a higher K value so it will react with water to give an excess of* $OH^-$ *ions.*
20. *In a series, acids such as* $H_3PO_4$, $H_2PO_4^-$ *and* $HPO_4^{2-}$ *the extent of dissociation decreases as* $H^+$ *ions had to be removed from negative ions and consequently their strength decreases also. (*$H_3PO_4$ *is the relatively strongest one of the three).*

## Set 3 Acids/Base Reactions

### Multiple Choice Answers

*1. a, 2. e, 3. e, 4. e, 5. b, 6. c, 7. c, 8. e, 9. d, 10. e, 11. b, 12. b, 13. c, 14. c, 15. b, 16. c, 17. c, 18. b, 19. b, 20. c.*

### Acid Base Calculations

1. *No. of moles of* $Ca(OH)_2 = cV = 0.015 \times 0.02 = 3.0 \times 10^{-4}$ *mol*
$n(OH^-) = 2 \times 3.0 \times 10^{-4}$ *mol* $= 6.0 \times 10^{-4}$ *mol* ($Ca(OH)_2 \rightarrow 2OH^-$).
$n(H^+)$ *from* $HNO_3 = 0.010 \times 0.080 = 8.0 \times 10^{-4}$ *mol.*
$OH^-$ *is the limiting reagent. So* $6.0 \times 10^{-4}$ *moles of* $H_2O$ *is formed and* $n(H^+)$ *left will be* $8.0 \times 10^{-4} - 6.0 \times 10^{-4} = 2.0 \times 10^{-4}$ *mol.*
*Total volume of mixed solution = 20 + 80 = 100 mL so*
$[OH^-] = n/V = 2.0 \times 10^{-4}/0.10 = 2.0 \times 10^{-3}$ *mol* $L^{-1}$
$pH = 2.7$
2. $NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
$n(NaOH) = c \times V = 0.0250 \times 0.200$
$= 5.00 \times 10^{-3}$ *mol*
$n\ (HNO_3) = c \times V = 0.030 \times 0.175$
$= 5.25 \times 10^{-3}$ *mol*
*From the equation 1 mol of NaOH reacts with 1 mol of* $HNO_3$.
$\therefore 5.00 \times 10^{-3}$ *mol of NaOH will react with* $5100 \times 10^{-3}$ *mol* $HNO_3$.
$\rightarrow HNO_3$ *is in excess by* $0.25 \times 10^{-3}$ *mol and*
$[H^+] = n/V = (0.25 \times 10^{-3}/\ 0.055)$
$= 4.55 \times 10^{-3}$ *mol* $L^{-1}$
$\therefore pH = 2.35$
3. $Ba(OH)_2 + 2HNO_3 \rightarrow Ba(NO_3)_2 + 2H_2O(l)$
$n\ (Ba(OH)_2) = 0.0500 \times 0.200 = 0.0100$ *mol*
$n\ (HNO_3) = 0.200 \times 0.400\ = 0.0800$ *mol*
*From the equation, 1 mole of* $Ba(OH)_2$ *reacts with 2 mols of* $HNO_3$.
*Accordingly, 0.0100 mol of* $Ba(OH)_2$ *reacts with 0.0200 mol of* $HNO_3$.
$HNO_3$ *is in excess by 0.0600 mol.*
*The final concentration after diluting to 6.00 L is [0.0600 /6.00 L] = 0.01 M*
$[H^+] = 1.0 \times 10^{-2}$ *M and,* $pH = 2.0$
4.
(a) $pH = 2$
$\therefore [H^+] = 1.0 \times 10^{-2}$ *M; volume = 0.300 L*
$\therefore n\ (H^+)$ *needed* $= 1.0 \times 10^{-2} \times 0.300$
$= 3.00 \times 10^{-3}$ *mol*
*Hence, n (HCl) needed* $= 3.00 \times 10^{-3}$ *mol*
$\therefore m\ (HCl) = 3.00 \times 10^{-3} \times 36.46 = 0.109$ *g*
(b) $n\ (HCl) = 0.0730\ /\ 36.46 = 2.00 \times 10^{-3}$ *mol.*
$\therefore n\ (H^+) = 2.00 \times 10^{-3}$ *mol.*
$[H^+] = (2.00 \times 10^{-3}\ /2.00\ L) = 1.0 \times 10^{-3}$ *M*
$\therefore pH = 3.00$
5.
(a) $pH = 13$, *and hence* $[H^+] = 1.0 \times 10^{-3}$ *M and* $[OH^-] = 1.0 \times 10^{-1}$ *M*
$V = 0.600$ *L and,*
$n\ (NaOH) = c \times v = 1 \times 10^{-1} \times 0.600$
$= 6.00 \times 10^{-2}$ *mol*
$m\ (NaOH) = 6.00 \times 10^{-2} \times 40.0 = 2.40$ *g.*
(b) $m\ (NaOH) = 0.600$ *g and* $n\ (NaOH)$
$= 0.600/40.0 = 0.0150$ *mol*
$[NaOH] = n/V = 0.0150/\ 1.500 = 0.01$ *M, and* $[OH^-] = 1.0 \times 10^{-2}$ *M*
$\therefore [H^+] = 1.0 \times 10^{-12}$ *M, thus* $pH = 12.0$
6. $n\ (HCl) = 0.100 \times 0.020 = 2.00 \times 10^{-3}$ *mol*
*pH required is 3. Therefore,* $[H^+]$ *should be* $1.0 \times 10^{-3}$ *M.*
*Since the number of moles are the same before and after dilution,*
$0.100 \times 0.020 = 1.0 \times 10^{-3} \times V$
$V = [(0.100 \times 0.020)\ /1.0 \times 10^{-3}] = 2.0$ *L.*
$\therefore$ *Volume of water to be added*
$= 2.0\ L - 0.020\ L = 1.98\ L$
7. $pH = 4$ *and so* $[H^+] = 1.0 \times 10^{-4}$ *M.V = 1.00 L*
$n\ (HCl) = n(H^+) = c \times V = 1.0 \times 10^{-4} \times 1.00$
$= 1.0 \times 10^{-4}$ *mol*
$V\ (HCl\ @\ STP) = n \times 22.71$
$= 1.0 \times 10^{-4} \times 22.71 = 2.27$ *mL*
8. *All these hydrolyse to produce* $OH^-$ *ions.*