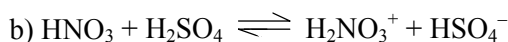


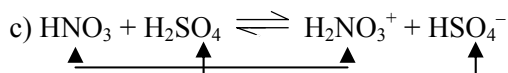
CHAPTER 8 ACIDS AND BASES

Page 211 Ex 8.1

1. C
2. C
3. B
4. C
5. D
6. a) HCl ; HPO_4^{2-} ; $\text{C}_5\text{H}_5\text{NH}^+$; $\text{H}_3\text{N}-\text{NH}_3^{2+}$; $^-\text{OOC}-\text{COOH}$.
 b) NO_3^- ; I^- ; SO_4^{2-} ; NH_3 ; HONH_2
 c) HSO_4^- and N_2H_5^+ .
 d) e.g. H_2O Conjugate base OH^- Conjugate acid H_3O^+
 or HCO_3^- Conjugate base CO_3^{2-} Conjugate acid H_2CO_3
7. a) i) An acid is a species that can donate hydrogen ions to another species.
 ii) A base is a substance that can accept a hydrogen ion from another species.

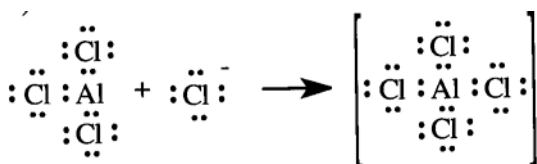


Sulfuric acid donates a proton to nitric acid, and thus behaves as an acid.



- d) Conjugate acid–base pair means that the two species are related by the gain/loss of only a single hydrogen ion (H^+).
8. a) Diprotic acids are acids in which each molecule of acid can donate two hydrogen ions.
 b) An amphiprotic species is one that can both accept and donate hydrogen ions, i.e. they can act as both an acid and a base. For example the hydrogencarbonate ion can act as an acid to form the carbonate ion (CO_3^{2-}), or it can act as a base and form carbonic acid (H_2CO_3).
 c) e.g. Water $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{OH}^-(\text{aq}) + \text{H}^+(\text{aq})$ and $\text{H}_2\text{O}(\text{l}) + \text{H}^+(\text{aq}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq})$
 or hydrogenphosphate $\text{HPO}_4^{2-}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{PO}_4^{3-}(\text{aq})$
 and $\text{HPO}_4^{2-}(\text{aq}) + \text{H}^+(\text{aq}) \rightleftharpoons \text{H}_2\text{PO}_4^-(\text{aq})$
9. a) A Lewis acid is a species that can accept a non-bonding electron pair to form a dative covalent bond.

b)



- c) A dative bond. Both of the electrons in the bond came from the chloride ion.
- d) i) Trigonal planar
 ii) Tetrahedral

CHAPTER 8 ACIDS AND BASES

10. a) Lewis base – the P in the PH_3 has a lone pair of electrons that it can donate.
 b) Lewis acid – the B in the BCl_3 has an incomplete valence shell, so can accept a pair of electrons.
 c) Lewis base – the S in the H_2S has a lone pair of electrons that it can donate.
 d) Lewis base – the S in the SF_4 has a lone pair of electrons that it can donate.
 e) Lewis acid – the Cu^{2+} has incomplete p and d orbitals, so it can accept a pair of electrons.

Page 212 Exercise 8.2

- D
- B
- $\text{Fe(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{FeSO}_4 + \text{H}_2(\text{g})$
 - $\text{PbCO}_3(\text{s}) + 2 \text{HNO}_3(\text{aq}) \rightarrow \text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
 - $\text{ZnO}(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - $\text{Ca}(\text{OH})_2(\text{s}) + 2 \text{HNO}_3(\text{aq}) \rightarrow \text{Ca}(\text{NO}_3)_2(\text{aq}) + 2 \text{H}_2\text{O}(\text{l})$
 - $2 \text{NaHCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2 \text{H}_2\text{O}(\text{l}) + 2 \text{CO}_2(\text{g})$
 - $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$

Page 214 Ex 8.3

- C
- A
- A
- $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$.

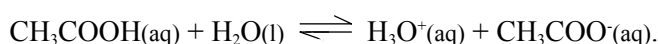
The presence of hydroxide ions make the solution alkaline or NH_4^+ is a weak acid and OH^- is a strong base, thus the solution is basic.

b) A base is a species that accepts a hydrogen ion, for example NH_3 accepts a hydrogen ion to form NH_4^+ . Weak means that the base is not completely converted to hydroxide ions in aqueous solution, so aqueous ammonia molecules are in equilibrium with ammonium and hydroxide ions.

c) A lower pH. The sodium hydroxide is totally converted to sodium and hydroxide ions, so $[\text{OH}^-] = 0.1 \text{ mol dm}^{-3}$ whereas in ammonia the ammonia is in equilibrium with the ammonium and hydroxide ions, so $[\text{OH}^-]$ is much less than 0.1 mol dm^{-3} . This means that the pH is lower in the ammonia solution.

- $\text{HCl}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$.

The acid is fully dissociated.



The acid is only partially dissociated.

b) The concentration of ions in the hydrochloric acid (strong acid) is much greater than that in the ethanoic acid (weak acid) so it has a much greater conductivity.

CHAPTER 8 ACIDS AND BASES

c) If the acids were added to magnesium ribbon (or any other reactive metal or metal carbonate) the rate of reaction, and hence of evolution of gas, would be much greater for the strong acid. Or if titrated with a strong base, NaOH, the equivalence point with the strong acid will be 7, whereas with the weak acid, it will be greater than 7 (approximately 9).

Page 216 Ex 8.4

- B
- B
- D
- $10^{-3} \text{ mol dm}^{-3}$
 - $10^{-11} \text{ mol dm}^{-3}$
 - 1 mol dm^{-3}
- 4
 - 2
 - 7
- The hydrochloric acid is full dissociated and so $[\text{H}^+] = 5 \cdot 10^{-4} \text{ mol dm}^{-3}$, hence the pH is between 3 and 4. The ethanoic acid is only partially dissociated and so, even though its concentration is much greater than that of the hydrochloric acid, the concentration of hydrogen ions and hence the effect on universal indicator, is similar.
- $n = \frac{m}{M} = \frac{0.63}{63} = 0.01, c = \frac{n}{v} = \frac{0.01}{0.1} = 0.1 \text{ mol dm}^{-3}$
 - 1
 - $10^{-13} \text{ mol dm}^{-3}$
 - $\text{HNO}_2 \rightleftharpoons \text{H}^+ + \text{NO}_2^-$
 - The pH would be greater than that of nitric acid, as nitrous acid is only partially dissociated, so the hydrogen ion concentration will be much lower than in nitric acid, giving a greater pH.
- Hydrochloric acid is a strong monobasic acid ($\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$), hence the $[\text{H}^+] = 0.01 \text{ mol dm}^{-3}$ and the pH is 2. Sulfuric acid is a strong dibasic acid ($\text{H}_2\text{SO}_4 \rightarrow 2\text{H}^+ + \text{SO}_4^{2-}$), hence the $[\text{H}^+] = 0.02 \text{ mol dm}^{-3}$ and the pH is less than 2. Ethanoic acid is a weak monobasic acid, and hence there is an equilibrium between the undissociated acid and the ions ($\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CH}_3\text{COO}^-$). This means that $[\text{H}^+]$ is much less than 0.01 mol dm^{-3} and hence the pH is much greater than 2.

CHAPTER 8 ACIDS AND BASES

Page 220,221 Ex 18.1

1. D
2. C
3. C
4. D
5. C
6. a) 1.14 b) 8.34 c) $1.48 \times 10^{-6} \text{ mol dm}^{-3}$ d) $0.00437 \text{ mol dm}^{-3}$ e) 13.13
7. a) $0.00100 \text{ mol dm}^{-3}$
b) $1.0 \times 10^{-11} \text{ mol dm}^{-3}$
c) Greater as each $\text{Ba}(\text{OH})_2$ forms two OH^- ions.
8. a) 7.21
b) $[\text{H}^+] = \sqrt{(0.02 \times 6.17 \times 10^{-8})} = 3.51 \times 10^{-5} \text{ mol dm}^{-3}$, therefore pH = 4.45
c) $1.62 \times 10^{-7} \text{ mol dm}^{-3}$
9. a) $K_a = \frac{[\text{H}^+].[\text{A}^-]}{[\text{HA}]} = \frac{(10^{-4.67})^2}{0.28} = 1.63 \times 10^{-9} \text{ mol dm}^{-3}$
b) Weaker ($pK_a = 8.79$); the greater pK_a , the weaker the acid.
c) $K_a = \frac{[\text{H}^+].[\text{A}^-]}{K_a} = \frac{(10^{-5})^2}{1.63 \times 10^{-9}} = 0.0613 \text{ mol dm}^{-3}$
10. a) $\text{HCN}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{CN}^-(\text{aq})$
b) $[\text{H}^+] = \sqrt{K_a \cdot [\text{HA}]} = \sqrt{(10^{-9.3} \cdot 0.01)} = 2.24 \times 10^{-6} \text{ mol dm}^{-3}$, therefore pH = 5.65.
This is much greater than that of a strong acid of this concentration (pH = 2).
c) Percentage = $\frac{100 \times [\text{H}^+]}{[\text{HA}]} = \frac{100 \times 2.24 \times 10^{-6}}{0.01} = 0.0224\%$

If diluted, the percentage increases.
d) pK_b of conjugate base = $14 - pK_a = 14 - 9.3 = 4.7$
 $[\text{OH}^-] = \sqrt{K_b \cdot [\text{A}^-]} = \sqrt{(10^{-4.7} \cdot 0.1)} = 1.41 \cdot 10^{-3} \text{ mol dm}^{-3}$,
therefore $[\text{H}^+] = 7.08 \cdot 10^{-12}$ & pH = 11.15

CHAPTER 8 ACIDS AND BASES

Page 223,224 Ex 18.2

1. B
2. D
3. B
4. C
5. a)

$$K_b = \frac{[BH^+][OH^-]}{[B]} = \frac{0.1 \times 10^{-4.7}}{0.1} = 10^{-4.7} = 2.00 \times 10^{-5} \text{ mol dm}^{-3}$$

b) The solution must be more acidic, so the equilibrium given above must be displaced to the left. This requires the addition of more ammonium chloride.

c)

$$[BH^+] = \frac{K_b[B]}{[OH^-]} = \frac{2.00 \times 10^{-5} \times 0.1}{1 \times 10^{-5}}$$

$$= 0.2 \text{ mol dm}^{-3}$$

d) Ethanoic acid and sodium ethanoate (or any similar strength weak acid and conjugate base).

6. a) A buffer solution is one that maintains an approximately constant pH when small amounts of acid or base are added to it.

b) $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$. When a small amount of acid is added, the hydrogen ions from the acid will react with the hydroxide ions in the equilibrium above. This will cause the equilibrium to shift to the right, forming more hydroxide ions until their pH is restored to almost the original value. If a small amount of alkali is added, then the hydroxide ions will react with ammonium ions, displacing the above equilibrium to the left until the hydroxide ion concentration falls to almost its initial value.

Page 225 Ex 18.3

1. C
2. B
3. C
4. a) Slightly acidic – salt of a weak base and strong acid.
 b) Neutral – salt of a strong base and a strong acid.
 c) Slightly acidic – salt of a small, highly charged cation.
 d) Slightly alkaline – salt of a strong base and a weak acid.

CHAPTER 8 ACIDS AND BASES

Page 227,228 Ex 18.4

1. A
2. B
3. C
4. a) Amount of HCl = $c \cdot V = 0.2 \cdot 0.015 = 3 \times 10^{-3}$ moles

HCl + NH₃ → NH₄Cl, therefore amount of ammonia = 3×10^{-3} moles

$$c = \frac{n}{V} = \frac{3 \times 10^{-3}}{0.020} = 0.15$$

0.15 mol dm⁻³

- b) i) $[\text{OH}^-] = \sqrt{K_b \cdot [\text{B}]} = \sqrt{(10^{-4.7} \cdot 0.15)} = 1.73 \times 10^{-3}$;
pOH = 2.76, so pH = 14 - 2.76 = 11.24
- ii) At the mid point $[\text{B}] = [\text{BH}^+]$, so pOH = pK_b = 4.7. pH = 14 - 4.7 = 9.3
- iii) $[\text{H}^+] = \sqrt{K_a \cdot [\text{HA}]} = \sqrt{(10^{-9.3} \cdot 0.15)} = 8.67 \cdot 10^{-6}$, therefore pH = 5.06

(c)



- d) The section of the curve around pH 9 is the buffering region, this is because the addition of a small amount of acid or alkali has little effect on the pH.
- e) Firstly the pH would remain at about pH 13 until almost the equivalence point. Secondly the volume of acid would be required to reach the equivalence point.

CHAPTER 8 ACIDS AND BASES

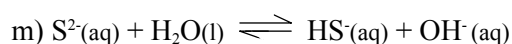
Page 229,230 Ex 18.5

1. B
2.
 - a) Methyl orange
 - b) Because the titration involves a weak base and a strong acid, the rapid change of pH at the end point will be from about 2 to 6. Methyl orange changes colour within this pH range.
 - c) From red to yellow.
 - d) The indicator is itself a weak acid, so enough alkali must be added to react with this as well as with the other acid present - this amount is assumed to be negligible.
 - e) Bromophenol blue. Because its pK_a is 4.0, the mid point of its color change will be at about pH 4, which is within the expected pH change at the equivalence point.
3.
 - a) Ethanoic acid is a weak acid, so only partially dissociated in aqueous solution.
 - b) Phenolphthalein.
 - c) The rapid change in pH at the equivalence point is only about 1 unit, too small for most indicators to completely change colour.
 - d) For Curve b, when the volume of alkali added is twice that required for neutralisation, $[NH_3] = [NH_4^+]$ and hence the pH at this point is equal to the pK_a .
4.
 - a)
 - i). The sudden drop in pH would occur at $\sim 12.5 \text{ cm}^3$ as H_2SO_4 is dibasic.
 - ii) The sudden drop in pH at $\sim 25 \text{ cm}^3$ would only go to $\sim \text{pH} = 6$ and this would gradually fall as excess ethanoic acid is added.
 - b) Phenolphthalein would be an appropriate indicator for all three as the rapid changes in pH at the equivalence point all pass through the range for phenolphthalein (10 to 8). Methyl orange, which changes between pH 4.5 and 3.0 would only change colour with the two strong acids, it would not change colour with ethanoic acid.
5.
 - a) HS^-
 - b) $H_2S_{(aq)} \rightleftharpoons HS^-_{(aq)} + H^+_{(aq)}$
 - c) $c = \frac{m}{MV} = \frac{3.4}{34} = 0.10 \text{ mol dm}^{-3}$
 - d) $[H^+] = \sqrt{K_a \cdot [HA]} = \sqrt{(9.55 \times 10^{-8} \cdot 0.1)} = 9.77 \times 10^{-5} \text{ mol dm}^{-3}$,
therefore $\text{pH} = 4.01$
 - e) A buffer solution is one that maintains an approximately constant pH when small amounts of acid or base are added to it.
 - f) The hydroxide ions would react with the hydrogen ions. As a result the equilibrium in b) will shift to the right, producing more hydrogen ions, until the pH is restored to almost its original level.
 - g) $0.196 \text{ mol dm}^{-3}$
 - h) $\text{pH} = 7.31$
 - i) $9.77 \cdot 10^{-3} \text{ mol dm}^{-3}$. The concentration is so low that it would soon be consumed if an acid were added.

j) The solution has a pH of 4.01 (see d), so with bromothymol blue it will be yellow (the “acid” colour), and with methyl yellow it will also be yellow (the ‘alkali’ colour).

k) Indicators are weak acids/bases in which the colours of the acidic and basic forms (HIn and In^- in the equilibrium $\text{HIn}_{(\text{aq})} \rightleftharpoons \text{H}^+_{(\text{aq})} + \text{In}^-_{(\text{aq})}$) have different colours. If an acid is added, the hydrogen ions drive the equilibrium to the left producing one colour. If an alkali is added, then this reacts with the hydrogen ions driving the equilibrium to the right, giving another colour. The pH at which this occurs will depend on the $\text{p}K_a$ value of the acid.

l) Bromothymol blue would be the more suitable, because it would change colour when the pH reached about 7, near to the expected equivalence point. Methyl yellow is already in the alkaline form and so would never change colour if titrated with an alkali.



n) The hydrogensulfide ion is amphiprotic, i.e. it can both gain and lose hydrogen ions.

o) $0.0891 \text{ mol dm}^{-3}$

p) $1.26 \cdot 10^{-13} \text{ mol dm}^{-3}$

q) Add either hydrogensulfide ions, or a dilute acid to convert sulfide ions to hydrogensulfide ions.