★ Unit 9 ★

Acid-Base Chemistry

Complete Question Bank

11th Grade Chemistry | Booklet Order: MCQs → Short Questions → Long Questions

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Multiple Choice Questions	Short Questions & Numericals	Long Questions + Solved Examples

What's inside

- **Section 1 — Multiple Choice Questions:** exercise MCQs, SLO-based MCQs by topic, and the full answer key.
- **Section 2 — Short Questions & Numericals:** exercise short questions, SLO-based short questions by topic, descriptive-question pointers, and worked numericals.
- **Section 3 — Long Questions:** complete descriptive answers with definitions, derivations, tables, titration curves and solved sample problems.

SECTION 1

Multiple Choice Questions • Unit 9: Acid-Base Chemistry

MULTIPLE CHOICE QUESTIONS (EXERCISE)

Select the correct answer.

1. Given the following reaction:

$$NH_{3(g)} + H_2O_{(l)} \rightleftharpoons NH_{4(aq)}^{+} + OH^-_{(aq)}$$

(a) NH_3 is the acid, H_2O is the base
(b) NH_3 is the base, H_2O is the acid
(c) NH_4$^+$ is the base, OH^- is the acid
(d) H_2O is the base, OH^- is the acid

2. The pH of 10^{-3} mol dm^{-3} of an aqueous solution of H_2SO_4 is:

(a) 3.0
(b) 2.7
(c) 2.0
(d) 1.5

3. The solubility product of AgCl is 2.0×10^{-10} mol^2 dm^{-6}. The maximum concentration of Ag^+ ions in the solution is:

(a) 2.0×10^{-10} mol dm^{-3}
(b) 1.41×10^{-5} mol dm^{-3}
(c) 1.0×10^{-10} mol dm^{-3}
(d) 4.0×10^{-20} mol dm^{-3}

4. Which indicator is typically used for titrations involving strong acids and strong bases?

(a) Methyl red
(b) Phenolphthalein
(c) Bromothymol blue
(d) Litmus solution

5. Which of the following is the conjugate base of water?

(a) $OH^-_{(aq)}$
(b) $H^+_{(aq)}$
(c) $H_2O_{(\ell)}$
(d) $H_3O^+_{(aq)}$

6. Which of the following is a Lewis acid but not a Bronsted-Lowry acid?

(a) HCℓ
(b) NH_3
(c) AℓCℓ$_3$
(d) H_2O

7. In an acid-base titration, the equivalence point is reached when:

(a) pH of the solution is 7.0
(b) The indicator changes color
(c) Equal volumes of acid and base have been added
(d) The reaction stops

8. If the concentration of Cℓ$^-$ ion in a solution is increased, the solubility of silver chloride (AgCℓ) will:

(a) Decrease
(b) Increase
(c) Remain unchanged
(d) Become zero

9. Which of the following pairs of substances can act as a conjugate acid-base pair according to the Bronsted-Lowry theory?

(a) HCℓ and NaOH
(b) NH_3 and NH_4^+
(c) H_2O and H_2SO_4
(d) H_2O and CH_4

10. If the pH of a solution is 11, what is the $[OH^-]$ concentration in the solution?

(a) 1×10^{-3} M (b) 1×10^{-11} M
(c) 1×10^{-2} M (d) 1×10^{-14} M

11. Which of the following pairs forms a buffer solution?

(a) HCℓ and NaCℓ (b) CH_3COONa and CH_3COOH
(c) NaOH and HCℓ (d) NH_3 and Na_2SO_4

SLO BASED MULTIPLE CHOICE QUESTIONS

Bronsted lowery concept

12. An acid and base react to give a _____ acid and base.

(a) Amphoteric (b) Salt
(c) Conjugate (d) Neutral

13. The conjugate base of $HC\ell O_4$ is:

(a) $C\ell O_4^-$ (b) $C\ell O_4^{2-}$
(c) $HC\ell O_3$ (d) $H_2C\ell O_4^+$

14. The definition of amphoteric is:

(a) Only acid (b) Only base
(c) Both acid and base (d) Neither

15. Which is amphoteric?

(a) NaOH (b) HCℓ
(c) H_2O (d) NH_3

16. The Bronsted-Lowry definition identifies acids as:

(a) Electron pair acceptors (b) Proton donors
(c) Proton acceptors (d) Electron pair donors

17. In the reaction $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$, H_2O acts as a:

(a) Acid (b) Base
(c) Neutral (d) Salt

18. Which one is a monoprotic acid

(a) H_3PO_4 (b) H_2SO_4
(c) HCℓ (d) H_2CO_3

19. The conjugate acid of NH_3 is:

(a) NH_2^- (b) NH_4^+
(c) NO_2 (d) N_2H_4

20. The conjugate base of H_2CO_3 is

(a) HCO_3^- (b) CO_3^{2-}
(c) CO_2 (d) H_2CO_3

21. Which of these is NOT a property of acids?

(a) Conduct electricity (b) Turn litmus red
(c) Sour taste (d) Feel slippery

22. Litmus turns which color in a basic solution?

(a) Orange (b) Red
(c) Blue (d) Colorless

23. Which acid is present in vinegar?

(a) HNO_3 (b) CH_3COOH
(c) H_2SO_4 (d) HCℓ

24. Which of the following is a tribasic acid?

(a) H_2SO_4 (b) HNO_3
(c) HCℓ (d) H_3PO_4

25. Which of the following is a diprotic acid?

(a) HCℓ (b) H_2SO_4
(c) CH_3COOH (d) HNO_3

26. Which species is both a Bronsted acid and base?

(a) H_2O (b) Na^+
(c) OH^- (d) $C\ell^-$

27. NaOH is not considered a Bronsted-Lowry acid because:

(a) It's not soluble (b) Doesn't donate H^+
(c) Doesn't produce H^+ (d) It's neutral

28. $HC\ell + NH_3 \rightarrow NH_4C\ell$ shows:

(a) Redox reaction (b) Acid-base reaction
(c) Bronsted acid-base (d) Precipitation

Lewis concept

29. Which one is a Lewis acid?

(a) BF_3 (b) OH^-
(c) $C\ell^-$ (d) NH_3

30. Which one is a Lewis base?

(a) H^+ (b) NH_3
(c) HCℓ (d) BF_3

Ionic product of H_2O

31. Water's ion product K_w at 25 °C is:

(a) 1×10^{-2} (b) 1×10^{-12}
(c) 1×10^{-14} (d) 1×10^{-16}

pH and pOH

32. Which has the highest pH?

(a) 0.01 M NaOH (b) 0.1 M NaOH
(c) 0.1 M HCℓ (d) 0.01 M HCℓ

33. pH of 1×10^{-4} M HCℓ is:

(a) 3.5 (b) 4
(c) 4 (d) 7

34. The pOH of 1×10^{-3} M HCℓ is:

(a) 3 (b) 11
(c) 11 (d) 1

35. A solution at pH < 7 has:

(a) $[H^+] < 10^{-7}$ (b) $[H^+] > 10^{-7}$

(c) $[OH^-] > 10^{-7}$ (d) $[OH^-] < 10^{-7}$

36. If $[H^+] = 1\times10^{-5}$ M, then pH =
(a) 4 (b) 3
(c) 5 (d) 9

37. Which dialyzing solution has highest pH?
(a) Vinegar (b) Lemon juice
(c) Ammonia solution (d) Coffee

38. A 0.01 M solution of a strong acid has a pH of:
(a) 3 (b) 2
(c) 5 (d) 4

39. The pH scale ranges typically from:
(a) 1–10 (b) 0–14
(c) −1 to 1 (d) 7–14

40. The pH of distilled water at 25 °C is:
(a) 8 (b) 6
(c) 7 (d) 5

41. pH of 10^{-9} M HCℓ is:
(a) 5 (b) 9
(c) Slightly below 7 (d) Exactly 9

42. KOH is an example of a:
(a) Weak acid (b) Strong base
(c) Neutral salt (d) Weak base

43. Which has the lowest pH?
(a) 0.1 M CH_3COOH (b) 0.1 M HCℓ
(c) 1 M HCℓ (d) 1 M NH_3

44. The pH of a buffer doesn't change much when:
(a) Diluted with water (b) Acid or base is added
(c) Heated (d) Temp changes

45. Acid rain is mainly due to:
(a) CO_2 (b) SO_2 & NO_2
(c) H_2 (d) CH_4

46. Which is more acidic: pH 3 or pH 5?
(a) Both same (b) pH 3
(c) pH 5 (d) Cannot say

47. The logarithmic nature of pH means:
(a) Each pH unit = 2x H^+ (b) 10x H^+ per unit
(c) H^+ drops linearly (d) No change in H^+

48. pH of blood is around:
(a) 6.0 (b) 5.4
(c) 7.4 (d) 8.0

49. A solution with equal $[H^+]$ and $[OH^-]$ is:
(a) Basic (b) Acidic
(c) Neutral (d) Salt

50. Which of the following does not affect pH?
(a) Concentration (b) Temperature
(c) Color of solution (d) Strength

Ionization constant of acids

51. Strong acids completely dissociate in water because:
(a) High solubility (b) High K_a
(c) Forms complex (d) Low K_a

52. Which of the following is a weak base?
(a) KOH (b) NaOH
(c) NH_3 (d) $Ca(OH)_2$

53. If K_a of acetic acid is 1.8×10^{-5}, it is:
(a) Strong acid (b) Weak acid
(c) Base (d) Neutral

54. The strength of a base is measured by its:
(a) K_a (b) K_b
(c) pK_w (d) K_w

55. A base that ionizes completely in aqueous solution is called:
(a) Weak base (b) Strong base
(c) Conjugate acid (d) Neutral

56. An acid with low K_a value is:
(a) Strong (b) Weak
(c) Base (d) Neutral

57. Strong bases have:
(a) Low K_b (b) High K_b
(c) Low pH (d) High K_a

58. Which is a weak acid?
(a) HNO_3 (b) HF
(c) H_2SO_4 (d) HCℓ

59. The K_a of a weak acid is 10^{-5}. Its pK_a is:
(a) 3 (b) 2
(c) 5 (d) 10

Common Ion Effect

60. Which one of the following compound is added in purification of NaCℓ in common ion effect.
(a) H_2SO_4 (b) HCℓ
(c) HNO_3 (d) HF

61. The suppression of ionization of a weak electrolyte by adding common ion is:
(a) Le-Chateleir's effect (b) Common-ion effect
(c) Auf Bau effect (d) Hund's rule

Buffer Solution

62. A buffer solution is:

(a) Strong acid + strong base (b) Weak acid + its salt
(c) Weak base + its salt (d) Weak base + salt of strong acid

63. The pK_a value for HCOOH is?

(a) 4.74 (b) 4.78
(c) 4.24 (d) 3.78

Solubility Product

64. The solubility product of $PbSO_4$ is:

(a) 1.6×10^{-8} (b) 1.4×10^{-2}
(c) 1.8×10^{-6} (d) 1.6×10^{-2}

65. Which one of the following is not the application of solubility products?

(a) Determination of solubility from K_{sp} (b) Common ion effect
(c) Predicting precipitation (d) Dissociation of bonds

Salt Hydrolysis

66. A salt from strong acid + weak base gives:

(a) Neutral solution (b) Acidic solution
(c) Precipitate (d) Basic solution

67. CH_3COONa is a salt of:

(a) Strong acid + strong base (b) Weak acid + strong base
(c) None (d) Weak base + weak acid

68. Strong acid + strong base neutralization gives:

(a) Weak salt (b) Neutral salt
(c) None (d) Acidic salt

69. A salt that hydrolyzes in water is from:

(a) Strong acid & base (b) Weak acid or base
(c) None (d) Both strong

Acid-Base Indicators

70. A substance that changes color with pH is a:

(a) Base (b) Indicator
(c) Acid (d) Catalyst

71. Methyl orange in base turns:

(a) Orange (b) Yellow
(c) Red (d) Blue

72. In a titration, the equivalence point is when:

(a) No acid left (b) pH = 7
(c) Moles of acid = moles of base (d) Buffer forms

ANSWER KEY

1.	2.	3.	4.	5.	6.	7.	8.
B	B	B	B	A	C	C	A
9.	10.	11.	12.	13.	14.	15.	16.
B	A	B	C	C	C	C	B
17.	18.	19.	20.	21.	22.	23.	24.
B	C	B	A	D	C	B	D
25.	26.	27.	28.	29.	30.	31.	32.
B	A	C	C	B	B	C	B
33.	34.	35.	36.	37.	38.	39.	40.
C	C	B	C	C	B	B	C
41.	42.	43.	44.	45.	46.	47.	48.
C	B	C	C	B	B	B	B
49.	50.	51.	52.	53.	54.	55.	56.
C	C	B	B	B	B	B	B
57.	58.	59.	60.	61.	62.	63.	64.
B	B	C	B	B	D	D	A
65.	66.	67.	68.	69.	70.	71.	72.
D	B	A	B	B	B	C	C

SECTION 2

Short Questions & Numericals • Unit 9: Acid-Base Chemistry

SHORT QUESTIONS ANSWERS (EXERCISE)

Attempt the following short-answer questions.

1. Define the following with an example for each:

(i) Ionization constant (ii) Solubility product (iii) Common ion effect (iv) Acid-base Indicator

Ans.

(i) Ionization constant: It is the extent of ionization of acids and bases which is much less than 100%. It is the quantitative measure of the strength of an acid or base. 'K_a' is ionization constant of acids and 'K_b' is ionization constant of bases.

$$HA + H_2O \rightleftharpoons H_3O^+ + A^- \qquad K_a = [H_3O^+][A^-]/[HA]$$

(ii) Solubility product: It is the product of concentration of ions raised to the exponent equal to the coefficient of the balanced equation. It is represented by 'K_{sp}'.

Example:

$$PbSO_4 \rightleftharpoons Pb^{2+} + SO_4^{2-} \qquad K_{sp} = [Pb^{2+}][SO_4^{2-}]$$

(iii) Common ion effect: The suppression of ionization of a weak electrolyte by adding a common ion to it is called common ion effect.

Example:

When HCℓ gas is passed through the saturated solution of NaCℓ, due to formation of common $C\ell^-$ ions, concentration is increased and NaCℓ is crystallized out.

$$NaC\ell_{(s)} \rightleftharpoons Na^+_{(aq)} + C\ell^-_{(aq)} \qquad HC\ell_{(g)} \rightleftharpoons H^+_{(aq)} + C\ell^-_{(aq)}$$

(iv) Acid-base Indicator: An indicator is a substance that changes colour to mark a titration's end point. Acid-base indicators exhibit one colour in acid and another in base. They are mostly weak organic acids or weak organic bases.

Example:

Phenolphthalein, methyl orange etc.

2. Differentiate between

(i) Hydrolysis and dissolution (ii) Acidic and basic buffer solutions

Ans.

(i) Hydrolysis: The process in which a substance dissolves in water, dissociates into ions and interacts with water molecules is called hydrolysis.

- It is a chemical process.
- Formation of new substances occurs.

Example:

Hydrolysis of ester.

Dissolution: The process in which a substance dissolves in a solvent and forms a homogeneous mixture.

- It is a physical process.
- Formation of new substance does not occur.

Example:

Dissolution of sugar in water.

(ii) Acidic buffer: The buffer formed by mixing a weak acid and salt of a strong base is an acidic buffer.

Example:

Mixture of CH_3COOH and CH_3COONa

Basic buffer: The buffer formed by mixing a weak base and salt of a strong acid is a basic buffer.

Example:

Mixture of $NH_4C\ell$ and NH_4OH

3. Explain the concept of conjugate acid-base pairs. How are they related in terms of proton transfer?

Conjugate acid: A specie formed when a base accepts a proton from the acid is called conjugate acid.

In this reaction, $C\ell^-$ and H_3O^+ are conjugate acid-base pairs.

$$HC\ell + H_2O \rightleftharpoons H_3O^+ + C\ell^-$$

acid *base* *conj. acid* *conj. base*

4. What is the relationship between the strength of an acid and the strength of its conjugate base?

Ans. There is an inverse relationship between the strength of an acid and strength of its conjugate base. The strong acid produces a weak conjugate base. Similarly, the strong base produces a weak conjugate acid.

Example:

HCℓ is a strong acid and its conjugate base $C\ell^-$ is a weak base.

5. For the following three reactions, identify the reactants that are Arrhenius bases, Bronsted-Lowry bases, or Lewis bases. State which type of base each reactant is. Explain your answers.

(i) $NaOH_{(s)} \rightleftharpoons Na^+_{(aq)} + OH^-_{(aq)}$

(ii) $HF_{(aq)} + H_2O_{(\ell)} \rightleftharpoons F^-_{(aq)} + H_3O^+_{(aq)}$

(iii) $H^+_{(aq)} + NH_{3(aq)} \rightleftharpoons NH_4^+{}_{(aq)}$

Ans.

(i) In this reaction, NaOH dissociates in aqueous solution to produce OH^- ion. Hence, NaOH is an Arrhenius base.

(ii) In this reaction, H_2O behaves as Bronsted-Lowry base because it accepts a proton (H^+) from HF.

(iii) In this reaction, NH_3 behaves as a Lewis base because it donates a pair of electrons to H^+.

6. An amphoteric substance can behave as either an acid or a base. Identify whether water behaves as an acid or a base in each of the following reactions.

(i) $H_2O + HC\ell \rightleftharpoons H_3O^+ + C\ell^-$

(ii) $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

(iii) $HNO_3 + H_2O \rightleftharpoons H_3O^+ + NO_3^-$

(iv) $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

Ans.

(i) H_2O behaves as a base because it accepts a proton (H^+) from HCℓ.

(ii) H_2O behaves as an acid because it donates a proton (H^+) to NH_3.

(iii) H_2O behaves as a base because it accepts a proton (H^+) from HNO_3.

(iv) H_2O behaves as a base because it accepts a proton (H^+) from CH_3COOH.

7. Which salt would you expect to dissolve more readily in acidic solution: Barium carbonate or Copper sulfide? Explain.

($K_{sp(BaCO3)} = 1.1 \times 10^{-10}$, $K_{sp(CuS)} = 8 \times 10^{-34}$)

Ans. Barium carbonate ($BaCO_3$) dissolves more readily than CuS in acidic solution. When $BaCO_3$ dissociates,

$$BaCO_3 \rightleftharpoons Ba^{2+} + CO_3^{2-}$$

its conjugate base CO_3^{2-} reacts with H^+ ions and forms a weak acid H_2CO_3.

$$CO_3^{2-} + 2H^+ \rightleftharpoons H_2CO_3$$

Hence, due to formation of H_2CO_3, solubility of $BaCO_3$ increases. While in CuS,

$$CuS \rightleftharpoons Cu^{2+} + S^{2-}$$

its conjugate base does not react with H^+ ions. Hence, its solubility is less than $BaCO_3$. Moreover, the K_{sp} of $BaCO_3$ is more than the K_{sp} of CuS. So, $BaCO_3$ is more soluble than CuS.

8. Why does common ion effect decrease solubility of a less soluble salt?

Ans. The addition of a common ion to the solution of a less soluble electrolyte suppresses (decreases) its ionization and the concentration of unionized species increases, which may come out as a precipitate. That's why solubility of a salt decreases due to common ion effect.

9. State the basic principle of solubility product. Mention factors affecting solubility product.

Ans. The solubility product is the product of the concentrations of ions raised to an exponent equal to the co-efficient of the balanced equation. K_{sp} is the measure of dissociation of sparingly soluble salt. It is the product of molar solubilities of two ions at equilibrium stage.

Factors affecting K_{sp}:

(i) Temperature — K_{sp} is temperature dependent, value is usually very small at room temperature.

(ii) Common ion effect — The solubility of slightly soluble salts decreases due to common ions.

10. Prove by equations what happens when Na_2CrO_4 is added to a saturated solution of $PbCrO_4$.

Ans. When the solution of Na_2CrO_4 and $PbCrO_4$ are mixed, the concentration of $CrO_4^{2-}{}_{(aq)}$ is increased according to Le-Chatelier's principle. The solubility of $PbCrO_4$ decreases as it precipitates out. This is due to common ion effect of CrO_4^{2-} ion.

$$PbCrO_{4(aq)} \rightleftharpoons Pb^{2+}{}_{(aq)} + CrO_4^{2-}{}_{(aq)}$$

$$Na_2CrO_{4(aq)} \rightleftharpoons 2Na^+ + CrO_4^{2-}$$

11. According to the Lewis acid-base concept, boron trifluoride (BF_3) can act as an acid. Is this statement correct?

Ans. Yes, according to the Lewis acid-base concept, Boron trifluoride (BF_3) can act as an acid. Because in BF_3, central atom 'B' has incomplete octet or deficit of 2 electrons and it has ability to accept an electron pair.

12. If the concentration of hydrogen ions in a solution is 1×10^{-5} M, what is the pH of the solution?

Ans. Solution:

$$[H^+] = 1 \times 10^{-5}$$

$$pH = -\log (1 \times 10^{-5})$$

pH = 5

SLO BASED SHORT QUESTION ANSWERS

Bronsted-Lowry concept

13. What are microscopic characteristics of acids and bases?

Ans. Acids

- They have sour taste.
- They change litmus from blue to red.
- They change the colour of dyes.

Base

- They have bitter taste.
- They change litmus from red to blue.
- They change colour of indicators like phenolphthalein from colourless to pink.

14. What is the Bronsted-Lowry concept of acids and bases?

Ans. In Bronsted-Lowry theory, an acid is a proton (H^+) donor, and a base is a proton acceptor. This concept is broader than Arrhenius'.

Example:

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

15. What is a conjugate acid-base pair?

Ans. A conjugate acid-base pair differs by one proton. When an acid donates a proton, it forms its conjugate base, and vice versa.

Example:

$HC\ell$ / $C\ell^-$ and NH_4^+ / NH_3 are conjugate pairs.

16. What is meant by amphoteric substances?

Ans. Amphoteric substances can act as both acids and bases depending on the environment.

Example:

Water (H_2O) can accept or donate a proton.

17. What is meant by monoprotic and polyprotic acids?

Ans. Monoprotic acids donate one proton per molecule (e.g., $HC\ell$), while polyprotic acids donate more than one (e.g., H_2SO_4, H_3PO_4).

18. What happens when $HC\ell$ reacts with NH_3?

Ans. $HC\ell$ donates H^+ to NH_3:

$$HC\ell + NH_3 \rightarrow NH_4C\ell$$

It's a classic Bronsted acid-base reaction.

19. What is meant by conjugate acid and conjugate base?

Ans. When a base accepts a proton, it becomes a conjugate acid. When an acid donates a proton, it forms a conjugate base.

Example:

$NH_3 \rightarrow NH_4^+$; $H_2CO_3 \rightarrow HCO_3^-$

20. Why are acids sour and bases bitter?

Ans. Due to their chemical nature: acids release H^+ which activate sour taste receptors, while bases release OH^- which give a slippery, bitter feel.

Lewis concept

21. Define Lewis acid and Lewis base.

Ans. A Lewis acid is an electron pair acceptor, and a Lewis base is an electron pair donor.

Example:

NH_3 donates an electron pair to BF_3, which accepts it.

Ionic product of water

22. What happens to water's ionization with temperature?

Ans. Ionization of water increases with temperature, so K_w and $[H^+]$, $[OH^-]$ increase, though pH may decrease slightly.

23. Define ionization constant of water (K_w).

Ans. $K_w = [H^+][OH^-] = 1 \times 10^{-14}$ mol^2/dm^6 at 25 °C. It reflects self-ionization of water.

pH and pOH

24. What is pH and how is it calculated?

Ans. pH is a measure of hydrogen ion concentration. It is calculated as:

$$pH = -\log [H^+]$$

25. What is the pH of a neutral solution at 25 °C?

Ans. A neutral solution has $[H^+] = 1 \times 10^{-7}$ M, so pH = 7. Water is neutral at 25 °C.

26. What is pOH and how is it related to pH?

Ans. $pOH = -\log[OH^-]$. It is related to pH by the equation:

$$pH + pOH = 14 \quad \text{(at 25 °C)}$$

27. What is the pH of a strong acid like 0.01 M HCℓ?

Ans. Since HCℓ is a strong acid, it ionizes completely:

$$[H^+] = 0.01 \text{ M} \qquad pH = -\log(0.01) = 2$$

Ionization Constant of Acids (K_a)

28. How is strength of acid measured?

Ans. By acid dissociation constant K_a. Stronger base has higher K_a and produces more H^+ ions. The K_a is the ratio of concentration of acid ionized to the concentration of acid initially added.

29. How is pK_a related to acid strength?

Ans. $pK_a = -\log K_a$

Lower pK_a means stronger acid. It's a logarithmic measure of acid strength.

Common ion effect

30. What is common ion effect? Give example.

Common Ion Effect: The suppression of ionization of a weak electrolyte by adding a common ion to it is called common ion effect.

Purification of sodium chloride: NaCℓ is purified by passing hydrogen chloride gas through brine (saturated solution of NaCℓ). Sodium chloride is fully ionized in the solution.

$$NaC\ell_{(s)} \rightleftharpoons Na^+_{(aq)} + C\ell^-_{(aq)}$$

Buffer solution

31. What is meant by a buffer solution?

Ans. A buffer resists pH changes when small amounts of acid or base are added. It usually consists of a weak acid and its salt or weak base and its salt.

32. Give an example of an acidic buffer.

Ans. A mixture of CH_3COOH and CH_3COONa is an acidic buffer that maintains pH around 4.75.

33. What are the applications of buffers?

Ans. Buffers are used in blood, pharmaceuticals, fermentation, and laboratory titrations to maintain stable pH.

Solubility product

34. What is solubility product? Give example.

Solubility Product: The solubility product is the product of the concentrations of ions raised to an exponent equal to the co-efficient of the balanced equation.

Example:

The K_{sp} of $PbSO_4$ is 1.6×10^{-16}.

Salt hydrolysis

35. Why NaCℓ cannot be hydrolyzed in water?

Ans. In salts of strong acids and strong bases like sodium chloride (NaCℓ), the conjugate base of a strong acid ($C\ell^-$ from HCℓ) is very weak and does not significantly react with water. The conjugate acid of a strong base (Na^+ from NaOH) is also very weak and does not significantly react with water.

$$NaC\ell_{(s)} \xrightarrow{-H_2O-} Na^+_{(aq)} + C\ell^-_{(aq)}$$

36. What is salt hydrolysis?

Salt hydrolysis: The breakdown of a salt (made up of weak acid or base) by reacting with water, resulting in the bond breaking in that salt which changes the pH of water, is called hydrolysis.

Acid-base indicators

37. What is meant by titration?

Ans. Titration is a technique used to determine the concentration of an unknown acid or base using a solution of known concentration.

38. What is equivalence point in titration?

Ans. It is the point where moles of acid equal moles of base in titration. The pH at this point depends on the nature of acid/base.

39. What is the role of indicators in titration?

Ans. Indicators change colour near the equivalence point to help detect the end of titration.

Example:

Phenolphthalein turns pink in base.

DESCRIPTIVE QUESTIONS

3. Describe the Bronsted-Lowry theory of acids and bases. Provide examples of conjugate acid-pairs and explain clearly their relationship.

Ans. See Q.2 from the Long Questions section.

4. Define the Lewis theory of acids and bases. How does this theory differ from the Bronsted-Lowry theory? Give examples of Lewis acids and bases that do not involve proton transfer.

Ans. See Q.3 from the Long Questions section.

5. Discuss applications and implications of the common ion effect in various fields.

Ans. See Q.9 from the Long Questions section.

6. What is the solubility product for sparingly soluble salts? Give its two applications.

Ans. See Q.12 from the Long Questions section.

7. Describe the general shape of a titration curve for a strong acid titrated with a strong base. How can you identify the equivalence point on this titration curve?

Ans. See Q.15 from the Long Questions section.

NUMERICAL PROBLEMS

8. A buffer solution has a pH of 5.0. It is made from a weak acid HA with a pK_a of 4.8. What is the ratio of the concentration of the conjugate base $[A^-]$ to the concentration of the weak acid [HA] in this buffer?

Ans. Given: pH = 5.0, $pK_a = 4.8$. To find: $[A^-]/[HA]$

Solution: By the use of Henderson's equation

$$pH = pK_a + \log([A^-]/[HA])$$

$$5 = 4.8 + \log([A^-]/[HA]) \qquad \log([A^-]/[HA]) = 0.2$$

Taking antilog on both sides:

$[A^-]/[HA]$ = antilog 0.2 = 1.5848

9. Calculate the solubility of a sparingly soluble salt lead(II) iodide (PbI_2) in water. It has $K_{sp} = 1.4 \times 10^{-8}$.

Ans. Solution:

Let solubility 'S' in terms of mol dm^{-3}

$$PbI_2 \rightleftharpoons Pb^{2+} + 2I^-$$

Initial stage: 0 + 0 Equilibrium stage: S + 2S

$$K_{sp} = [Pb^{2+}][I^-]^2 = (S)(2S)^2 = 4S^3$$

$$4S^3 = 1.4 \times 10^{-8} \qquad S^3 = 3.5 \times 10^{-9}$$

$S = 1.52 \times 10^{-3}$ mol dm^{-3}

10. The molar solubility of silver chromate (Ag_2CrO_4) in pure water at 298 K is 6.5×10^{-5} mol dm^{-3}. Calculate the K_{sp} of silver chromate at this temperature.

Ans. Solution: Solubility of $Ag_2CrO_4 = S = 6.5 \times 10^{-5}$ mol dm^{-3}

$$Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^{2-}$$

Equilibrium stage: $2S = 1.3 \times 10^{-4}$ $S = 6.5 \times 10^{-5}$

$$K_{sp} = [Ag^+]^2[CrO_4^{2-}] = (1.3 \times 10^{-4})^2 (6.5 \times 10^{-5})$$

$K_{sp} = 1.1 \times 10^{-12}$

SECTION 3

Long Questions • Unit 9: Acid-Base Chemistry

DESCRIPTIVE QUESTIONS

1. Discuss the general properties of acids, bases and indicators.

Ans. General properties of acids, bases and indicators

- Scientists have recognized that acids and bases have distinct characteristic properties, since the earliest days of experimental chemistry.
- Acids have a sour taste while bases are bitter.
- Acids and bases change the colour of certain dyes called indicators, such as litmus and phenolphthalein.
- Acids change litmus from blue to red and phenolphthalein from pink to colourless.
- Bases change litmus from red to blue and phenolphthalein from colourless to pink. From these colour changes, acids and bases neutralize each other.
- During neutralization, acids and bases react with each other to produce ionic substances called salts.

Interesting Information!

S.Q. How acidity of stomach is treated?

Ans. Acidity in our stomach is due to excess of HCℓ. It is treated by taking mild bases such as baking soda ($NaHCO_3$), milk of magnesia $Mg(OH)_2$ and aluminum hydroxide $A\ell(OH)_3$.

Bronsted-Lowry Concept

2. Explain the Bronsted-Lowry concept of acid and base.

Ans. Bronsted-Lowry Concept

Introduction: In 1923, the Danish chemist J.N. Bronsted and the English chemist T.M. Lowry independently expanded the Arrhenius theory of acid-base.

Need of Bronsted-Lowry concept

The limitations of Arrhenius theory are that it describes the reaction of an acid and base only in the aqueous medium. There are many reactions that occur in solvents other than water or in the absence of any solvent.

Bronsted-Lowry pointed out that acid-base reactions can be seen as proton-transfer reactions and that acids and bases can be defined in terms of this proton (H^+) transfer.

Definition of Acid: An acid is the specie donating a proton in a proton-transfer reaction.

Definition of Base: A base is a specie that accepts a proton from the other species in a proton transfer reaction.

When HCℓ dissolves in water, an H^+ ion (a proton) is transferred from HCℓ to water, where it becomes attached to a lone pair of electrons on the O atom and forms H_3O^+. In effect, HCℓ (the acid) has donated the H^+ and H_2O (the base) has accepted it.

$$HC\ell_{(g)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + C\ell^-_{(aq)}$$

Acid Base Conjugate acid Conjugate base

Conjugate Acid: A specie formed after a Bronsted base accepts a proton from the acid is called the conjugate acid.

Example:

The hydronium ion (H_3O^+) is the conjugate acid of water.

Conjugate Base: A specie formed when an acid donates a proton to a base is called the conjugate base.

Example:

$C\ell^-$ is conjugate base of HCℓ. HCℓ-$C\ell^-$ and H_2O-H_3O^+ are conjugate acid-base pairs.

When ammonia dissolves in water, proton transfer also occurs. An H^+ from H_2O attaches to the N atom's lone pair and NH_4^+ and OH^- are formed. With fewer H^+, the water molecule becomes OH^- ion:

$$NH_{3(g)} + H_2O_{(\ell)} \rightleftharpoons OH^-_{(aq)} + NH_{4(aq)}{}^+$$

Base Acid Conjugate base Conjugate acid

H_2O (the acid) has donated the proton and NH_3 (the base) has accepted it. NH_3-NH_4^+ and H_2O-OH^- are conjugate acid-base pairs.

Definition of Amphoteric: An amphoteric substance is a species that can act as either an acid and a base (it can lose or gain a proton), depending on the other reactant. H_2O is amphoteric in nature because it acts as a base in one case and as an acid in the other.

Application of Bronsted-Lowry theory

It can be applied to acids in solvents other than water or even solventless reactions. The reaction between gaseous ammonia and HCℓ gives solid $NH_4C\ell$.

$$HC\ell_{(g)} + NH_{3(g)} \rightarrow NH_4C\ell_{(s)}$$

HCℓ is an acid because in the reaction it donates a proton to the NH_3 molecule. The NH_3 acts as a base, even though hydroxide ion OH^- is not present, and accepts a proton from the HCℓ molecule.

Table 9.1 — Conjugate acid-base pairs of common species

Acid	Base		Conjugate Acid		Conjugate Base
HNO_3	H_2O	⇌	H_3O^+	+	NO_3^-
H_2SO_4	H_2O	⇌	H_3O^+	+	HSO_4^-
H_2O	CO_3^{2-}	⇌	HCO_3^-	+	OH^-

Lewis Concept of Acids and Bases

3. Define and explain the Lewis concept of acids and bases.

Ans. Lewis concept of acids and bases

Introduction: In 1923, G.N. Lewis proposed a generalized definition of acid-base behaviour in which acids and bases are identified by their ability to accept or to donate a pair of electrons and form a coordinate covalent bond.

Lewis Acid: A Lewis acid is any species (molecule or ion) that can accept a pair of electrons.

Lewis Base: A Lewis base is any species (molecule or ion) that can donate a pair of electrons.

Reaction of NH_3 and Ag^+

A Lewis acid-base reaction occurs when a base donates a pair of electrons to an acid. Two ammonia molecules act as Lewis bases, donating a pair of electrons to a positively charged silver ion (the Lewis acid).

Reaction: NH_3 (Lewis base) + Ag^+ (Lewis acid)

$$H_3N: + Ag^+ \rightarrow [H_3N \rightarrow Ag \leftarrow NH_3]^+$$

(each NH_3 donates its lone pair to form a co-ordinate bond with Ag^+ — the adduct carries the sum of charges, +1)

Reaction of BF_3 and F^-

The boron atom in boron trifluoride, BF_3, has only six electrons in its valence shell. The boron atom has an incomplete octet, so it can behave as an electron pair acceptor. As a result, BF_3 is a very good Lewis acid and reacts with many Lewis bases; a fluoride ion in the Lewis base donates one of its lone pairs:

Reaction: BF_3 (Lewis acid) + F^- (Lewis base)

$$:F: + B(F)_3 \rightarrow [F \rightarrow BF_3]^-$$

Quick Check 9.1

(a) Identify the conjugate acid-base pairs in the following reactions:

(i) $HCN_{(g)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + CN^-_{(aq)}$

Ans. the conjugate acid-base pairs are HCN-CN^- and H_2O-H_3O^+.

(ii) $CH_3COOH_{(\ell)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$

Ans. the conjugate acid-base pairs are CH_3COOH–CH_3OO^- and H_2O–H_3O^+.

(iii) $H_2S_{(g)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + HS^-_{(aq)}$

Ans. the conjugate acid-base pairs are H_2S-HS^- and H_2O-H_3O^+.

(iv) $HC\ell_{(g)} + HCO_{3(aq)}^- \rightleftharpoons H_2CO_{3(aq)} + C\ell^-_{(aq)}$

Ans. the conjugate acid-base pairs are $HC\ell$-$C\ell^-$ and HCO_3^--H_2CO_3.

(b) Identify the Lewis acid and Lewis base in the reaction between (i) SO_3 and O^{2-} (ii) $HC\ell$ and H_2O (iii) BF_3 and NH_3. Also write the balanced equation for each and explain.

(i) $SO_3 + O^{2-} \rightarrow SO_4^{2-}$ In this reaction, SO_3 is Lewis acid and O^{2-} is Lewis base.

(ii) $HC\ell_{(g)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + C\ell^-_{(aq)}$ In this reaction, $HC\ell$ is Lewis acid and H_2O is Lewis base.

(iii) $H_3N: + BF_3 \rightarrow [H_3N \rightarrow BF_3]$ In this reaction, NH_3 is Lewis base and BF_3 is Lewis acid.

Ionic Product of Water

4. How ionic-product of water is measured? Explain what is the effect of temperature on K_w.

Definition: The product of concentration of H^+ and OH^- ions in water is called ionic product of water or dissociation constant of water (K_w).

Explanation

- Pure water is a poor conductor of electricity but its conductance is measurable.
- Water undergoes self-ionization reversibly as follows,

$$H_2O_{(\ell)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

Net reaction $H_2O_{(\ell)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$

The equilibrium constant for this reaction can be written as follows:

$$K_c = [H^+][OH^-]/[H_2O] = 1.8 \times 10^{-16} \text{ mol dm}^{-3}$$

- The concentration of H_2O, $[H_2O]$ in pure water may be calculated to be 1000 g dm^{-3} divided by 18 g mol^{-1} giving 55.5 mol dm^{-3}.

- Water is present in very large excess and very few of its molecules undergo ionization, so its concentration remains effectively constant.
- Constant concentration of water is taken on the left hand side and multiplied with K_c to get another constant called K_w.

$$K_c[H_2O] = [H^+][OH^-] \qquad 1.8 \times 10^{-16} \times 55.5 = 1.01 \times 10^{-14} = [H^+][OH^-]$$

This value, 1.01×10^{-14} is called K_w of water at 25 °C.

$$K_w = [H^+][OH^-] = 10^{-14} \text{ at } 25\ °C$$

Effect of temperature on K_w: The effect of temperature on K_w is shown in the table below. The value of K_w increases almost 75 times when temperature is increased from 0 °C to 100 °C.

Table 9.2 — K_w at various temperatures

Temp. (°C)	K_w
0	0.11×10^{-14}
10	0.30×10^{-14}
25	1.0×10^{-14}
40	3.00×10^{-14}
100	7.5×10^{-14}

For neutral water: $[H^+] = [OH^-]$ $[H^+]^2 = 10^{-14}$ (at 25 °C) $[H^+] = [OH^-] = 10^{-7}$ mol dm^{-3}

Effect of addition of acid and base on K_w

Whenever some quantity of acid or base is added to water, then K_w remains the same, but $[H^+]$ and $[OH^-]$ are no more equal.

Addition of acid to water: $[H^+] > [OH^-]$

Addition of base to water: $[OH^-] > [H^+]$

During both of these additions, the values of K_w will remain the same, i.e. 10^{-14} at 25 °C.

pH and pOH

5. Explain concept of pH and pOH.

pH — Definition: The negative logarithm of molar concentration of H^+ ions in a mixture is called pH.

pOH — Definition: The negative logarithm of molar concentration of OH^- ions in a mixture is called pOH.

Explanation

- In all aqueous solutions, the concentration of H^+ and OH^- are too low to be conveniently expressed and used in calculations.
- In 1909, Sorenson, a Danish biochemist, introduced the term pH and pOH.

$$pH = -\log[H^+] \qquad pOH = -\log[OH^-]$$

For neutral water, $pH = -\log[10^{-7}] = 7$, $pOH = -\log[10^{-7}] = 7$, $pH + pOH = 14$

- The value of pH normally varies between 0 → 14 at 25 °C. The pH can have a negative value or be greater than 14.

Interesting Information!

S.Q. What pH values other than 0–14 are possible?

Ans. Solutions of negative pH and having values more than 14 are also known.

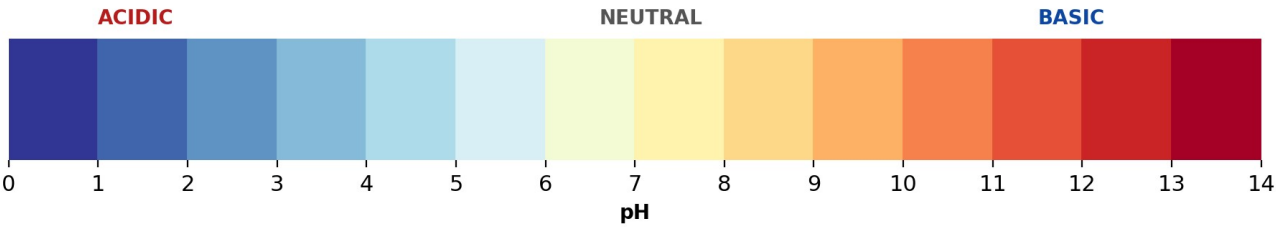


Fig: pH values of some common substances (acidic ↔ neutral ↔ basic)

Quick Check 9.2

(a) A solution is prepared by mixing equal volumes of two solutions: one with a pH of 4.0 and another with a pH of 10.0. Calculate the K_w for this mixture at 25 °C.

Ans. For the first solution: pH = 4.0, $[H^+] = 10^{-4}$ M, pOH = 10, $[OH^-] = 10^{-10}$ M. For the second: pH = 10, pOH = 4, $[H^+] = 10^{-10}$ M, $[OH^-] = 10^{-4}$ M.

$$[H^+]_{mix} = (10^{-4}+10^{-10})/2 = 5\times10^{-5}\ M \qquad [OH^-]_{mix} = (10^{-14}+10^{-4})/2 = 5\times10^{-5}\ M$$

$$K_w = [H^+]_{mix}[OH^-]_{mix} = (5\times10^{-5})(5\times10^{-5}) = 2.5 \times 10^{-9}$$

(b) At a specific temperature, K_w is 1.0×10^{-14}. The concentration of OH^- in a solution is 2.5×10^{-8} M. Calculate $[H^+]$ and pH.

$$[H^+] = K_w/[OH^-] = (1.0\times10^{-14})/(2.5\times10^{-8}) = 4.0 \times 10^{-7}\ M$$

$$pH = -\log(4.0 \times 10^{-7}) = 6.40$$

(c) A copper-plate etching solution is prepared by diluting concentrated HNO_3 to 0.30 M. Calculate $[H^+]$, pH, $[OH^-]$ and pOH at 25 °C.

$$HNO_3 \rightleftharpoons H^+ + NO_3^- \qquad [H^+] = 0.30\ M \qquad pH = -\log(0.30) = 0.5229$$

$$pOH = 14 - 0.5229 = 13.4771 \qquad [OH^-] = K_w/[H^+] = (1.0\times10^{-14})/0.3 = 3.33 \times 10^{-14}\ M$$

Ionization Constant of Acids (K_a)

6. Discuss the ionization constant of acids. How it affects the acid strength?

Definition: The value of K_a, called the dissociation constant of an acid, is the quantitative measure of the strength of the acid.

Explanation

Acids and bases, when dissolved in water, may or may not be completely dissociated. Many acids are weak electrolytes and ionize to an extent which is much less than 100%.

Derivation: Suppose we have an acid HA dissolved in water,

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

K_c for the reversible reaction will be written as follows:

$$K_c = [H_3O^+][A^-]/[HA][H_2O] \qquad K_c[H_2O] = [H_3O^+][A^-]/[HA]$$

Let $K_c[H_2O] = K_a$. Hence,

$$K_a = [H_3O^+][A^-]/[HA]$$

- This equation can be used to calculate K_a for any acidic solution if we know the pH or $[H^+]$ of that solution and the initial concentration of the dissolved acid [HA].
- This can also be used to calculate the equilibrium concentration of H_3O^+ and A^- produced if we know the initial concentration of acid HA and its K_a value.

Effect of K_a on strength of acid

(i) $K_a < 10^{-3}$ acid is weak.

(ii) $K_a = 1$ to 10^{-3} acids are moderately strong.

(iii) $K_a > 1$ acid is strong.

Table 9.3 — The values of K_a for some acids

Acid	Dissociation	K_a	Relative strength
HCℓ	$HC\ell \rightleftharpoons H^+ + C\ell^-$	Very large (10^{-6})	Very strong
HNO_3	$HNO_3 \rightleftharpoons H^+ + NO_3^-$	Very large (10^{-3})	Very strong
H_2SO_4	$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$	Large (10^{-3})	Very strong
HSO_4^-	$HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$	1.3×10^{-2}	Strong
HF	$HF \rightleftharpoons H^+ + F^-$	6.7×10^{-5}	Weak
CH_3COOH	$CH_3COOH \rightleftharpoons H^+ + CH_3OO^-$	1.85×10^{-5}	Weak
H_2CO_3	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^{-1}$	4.4×10^{-7}	Weak
H_2S	$H_2S \rightleftharpoons H^+ + HS^{-1}$	1.0×10^{-7}	Weak

Acid	Dissociation	K_a	Relative strength
NH_4^{1+}	$NH_4^{1+} \rightleftharpoons H^+ + NH_3$	5.7×10^{-10}	Weak
HCO_3^-	$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.7×10^{-11}	Weak
H_2O	$H_2O \rightleftharpoons H^+ + OH^-$	1.8×10^{-16}	Very weak

7. How is H_3O^+ ions calculated from K_a and molar concentration of weak acid?

Ans. Calculating H_3O^+ ions from K_a

- The weak acids do not completely dissociate in water.
- The concentration of hydronium ions is not equal to the initial concentration of the acid.
- The equilibrium expression is used to calculate the H_3O^+ concentration.

For a weak acid HA, the dissociation in water can be represented as:

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)} \qquad H^+_{(aq)} + H_2O \rightleftharpoons H_3O^+_{(aq)}$$

Let the initial concentration of the weak acid HA be C mol dm^{-3}. At equilibrium: $[H_3O^+] = x$; $[A^-] = x$ and $[HA] = C - x$. The expression for the acid dissociation constant (K_a) is:

$$K_a = [H_3O^+][A^-]/[HA] = x \cdot x/(C-x) = x^2/(C-x)$$

For weak acids, x is usually very small compared to C, so $C-x \approx C$. Therefore, the expression simplifies to:

$$K_a \approx x^2/C \qquad x^2 \approx K_a \times C \qquad x \approx \sqrt{(K_a \times C)}$$

Thus, the concentration of H_3O^+ is given by $x \approx \sqrt{(K_a \times C)}$

Quick Check 9.3

(a) The pH of a 0.10 M solution of formic acid, HCOOH, at 25 °C is 2.38. Calculate K_a for formic acid at this temperature.

Ans. pH = 2.38, $[H^+] = 10^{-2.38} = 4.17 \times 10^{-3}$ M. Initial concentration [HCOOH] = 0.10 M.

$$HCOOH \rightleftharpoons H^+ + HCOO^- \qquad K_a = (x)(x)/(0.1-x) = (4.17\times10^{-3})^2/(0.1-4.17\times10^{-3})$$

$$\mathbf{K_a = 1.74\times10^{-3}/0.09583 = 1.81 \times 10^{-4}}$$

(b) Calculate $[H_3O^+]$ in a 0.1 M solution of nitrous acid (HNO_2), given $K_a = 4 \times 10^{-4}$.

$$[H_3O^+] \approx \sqrt{(K_a \times C)} \approx \sqrt{(4\times10^{-4} \times 0.1)} = 2 \times 10^{-2} \text{ M}$$

(c) A vinegar sample has 0.837 M CH_3COOH. Its $[H_3O^+] = 3.86 \times 10^{-3}$ mol dm^{-3}. Calculate K_a for acetic acid.

$$K_a = [H_3O^+]^2/C = (3.86\times10^{-3})^2/0.837$$

$$\mathbf{K_a = 1.79 \times 10^{-5}} \quad \textbf{Answer}$$

Common Ion Effect

8. Define and explain common-ion effect. Give examples.

Definition: The suppression of ionization of a weak electrolyte by adding a common ion to it is called common ion effect.

Explanation

Purification of sodium chloride: NaCℓ is purified by passing hydrogen chloride gas through brine (saturated solution of NaCℓ). Sodium chloride is fully ionized in the solution. Equilibrium constant expression for this process can be written as follows:

$$NaCℓ_{(s)} \rightleftharpoons Na^+_{(aq)} + Cℓ^-_{(aq)} \qquad K_c = [Na^+][Cℓ^-_{(aq)}]/[NaCℓ]$$

HCℓ also ionizes in solution:

$$HCℓ_{(g)} \rightleftharpoons H^+_{(aq)} + Cℓ^-_{(aq)}$$

- On passing HCℓ gas, concentration of $Cℓ^-$ ion is increased.
- NaCℓ crystallizes out of the solution to maintain the constant value of the equilibrium constant. This type of effect is called the common ion effect.
- The addition of a common ion to the solution of a less soluble electrolyte suppresses its ionization and the concentration of unionized species increases, which may come out as a precipitate.

$$Na^+_{(aq)} + Cℓ^-_{(aq)} \rightleftharpoons NaCℓ_{(s)}$$

More examples of common ion effect

(i) Solubility of $KC\ell O_3$ — The solubility of less soluble salt $KC\ell O_3$ in water is suppressed by the addition of a more soluble salt $KC\ell$ by common ion effect. K^+ is a common ion. The ionization of $KC\ell O_3$ is suppressed and it settles down as precipitate.

$$KC\ell O_{3(s)} \rightleftharpoons K^+_{(aq)} + C\ell O_{3(aq)}^- \qquad KC\ell_{(s)} \rightleftharpoons K^+_{(aq)} + C\ell^-_{(aq)}$$

(ii) Solubility of H_2S — The dissociation of a weak acid H_2S in water can be suppressed by the addition of stronger acid $HC\ell$. H^+ is a common ion. H_2S becomes less dissociated in acidic solution. In this way, low concentration of S^{2-} ion is produced, which helps to precipitate radicals of the second group basic radicals during salt analysis.

$$H_2S_{(g)} \rightleftharpoons 2H^+_{(aq)} + S^{2-}_{(aq)} \qquad HC\ell_{(aq)} \rightleftharpoons H^+_{(aq)} + C\ell^-_{(aq)}$$

(iii) Solubility of $NH_4C\ell$ — An addition of $NH_4C\ell$ in NH_3 solution suppresses the concentration of OH^- due to the presence of a large excess of NH_4^+ from $NH_4C\ell$. This combination is used as a group reagent in third group basic radicals for salt analysis.

$$NH_4C\ell_{(aq)} \rightleftharpoons NH_{4(aq)}^+ + C\ell^-_{(aq)} \qquad NH_3 + H_2O_{(aq)} \rightleftharpoons NH_{4(aq)}^+ + OH^-_{(aq)}$$

Common ion effect finds extensive applications in the qualitative analysis and the preparation of buffers.

9. Discuss applications and implications of the common ion effect in various fields.

Ans. Applications of common ion effect

(i) Purification of $NaC\ell$ — $NaC\ell$ is purified by passing hydrogen chloride gas through the saturated brine. Sodium chloride is fully ionized in the solution. On passing $HC\ell$ gas, concentration of $C\ell^-$ ions is increased. Therefore, $NaC\ell$ crystallizes out of the solution to maintain the constant value of equilibrium constant.

(ii) Precipitation of $KC\ell O_3$ — The solubility of a less soluble salt $KC\ell O_3$ in water is suppressed by the addition of more soluble salt $KC\ell$ by common ion effect. K^+ is a common ion. The ionization of $KC\ell O_3$ is suppressed and it settles down as precipitate.

(iii) Precipitation of 2^{nd} group basic radicals — The dissociation of a weak acid H_2S in water can be suppressed by the addition of stronger acid $HC\ell$. H^+ is a common ion. H_2S becomes less dissociated in acidic solution. In this way, low concentration of S^{-2} is developed. This low concentration of S^{2-} ions helps to do the precipitation of radicals of second group basic radicals during salt analysis.

(iv) Preparation of buffers — Common ion effect has its extensive applications in the qualitative analysis and preparation of buffer solutions.

Buffer Solution

10. What are buffer solutions? Explain its types.

Definition: The solutions which resist the change in their pH when a small amount of an acid or a base is added to them, are called buffer solutions. They have a specific constant value of pH and their pH values do not change on dilution and on keeping for a long time.

Types of Buffer

Buffer solutions are mostly prepared by mixing two substances.

(i) Acidic Buffers

- By mixing a weak acid and its salt with a strong base. Such solutions give acidic buffers with pH less than 7.
- Mixture of acetic acid and sodium acetate is one of the best examples of acidic buffers.

(ii) Basic Buffers

- By mixing a weak base and its salt with a strong acid. Such solutions will give basic buffers with pH more than 7.
- Mixture of NH_4OH and $NH_4C\ell$ is one of the best examples of basic buffers.

Mechanism of action of Buffers

- The example of an acidic buffer consisting of CH_3COOH and CH_3COONa. CH_3COOH being a weak electrolyte undergoes very little dissociation.
- When CH_3COONa, which is a strong electrolyte, is added to CH_3COOH solution, then the dissociation of CH_3COOH is suppressed, due to common ion effect of CH_3COO^-.

$$CH_3COOH_{(aq)} + H_2O_{(\ell)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)} \qquad CH_3COONa_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + Na^+_{(aq)}$$

- If one goes on adding CH_3COONa in CH_3COOH solution, then the added concentrations of CH_3COO^- decrease the dissociation of CH_3COOH and the pH of solution increases. The greater the concentration of acetic acid as compared to CH_3COONa, lesser is the pH of solution.

Table 9.4 — Effect of addition of acetate ions on the pH of acetic acid solution

$[CH_3COOH]$ (mol dm^{-3})	$[CH_3COO^-]$ (mol dm^{-3})	% Dissociation	pH
0.10	0.00	1.3	2.89

$[CH_3COOH]$ (mol dm^{-3})	$[CH_3COO^-]$ (mol dm^{-3})	% Dissociation	pH
0.10	0.05	0.036	4.44
0.10	0.10	0.018	4.74
0.10	0.15	0.012	4.92

- A buffer mentioned above is a large reservoir of CH_3COOH and CH_3COO^- components.
- When an acid or H_3O^+ ions are added to this buffer, they will react with CH_3COO^- to give back acetic acid and hence the pH of the solution will almost remain unchanged, because CH_3COOH being a weak acid will prefer to remain un-dissociated.
- The buffer solution consisting of $NH_4C\ell$ and NH_4OH, can resist the change of pH and pOH, when acid or base is added from outside.
- When a base or OH^- ions are added in it, they will react with H_3O^+ to give back H_2O and the pH of the solution again will remain almost unchanged.

11. Derive an expression of Henderson's equation.

Ans. Calculating the pH of a Buffer

Consider a weak acid HA and its salt NaA with a strong base say NaOH. The reversible reactions for dissociation of HA and NaA are as follows:

$$HA \rightleftharpoons H^+ + A^- \qquad NaA \rightleftharpoons Na^+ + A^-$$

pH of a buffer solution can be calculated by using Henderson equation as given below:

$$pH = pK_a - \log([Acid]/[Salt])$$

Interchanging the numerator and denominator, the sign of log changes.

$$pH = pK_a + \log([Salt]/[Acid])$$

The relationship is called Henderson's equation.

- The equation shows that two factors evidently govern the pH of a buffer solution.
- First the pK_a of the acid used and the second is the ratio of the concentrations of the salt and the acid.
- The best buffer is prepared by taking equal concentration of salt and acid.
- pH is controlled by pK_a of the acid.

Example: For the buffer solution of acetic acid and sodium acetate, the pH can be calculated as follows.

$$[CH_3COOH] = [CH_3COONa]$$

Then,

$$pH = pK_a + \log([CH_3COONa]/[CH_3COOH]) = pK_a + \log(1) = pK_a + 0 = pK_a = 4.74$$

It means that the pH of this buffer is just equal to the pK_a of the acid.

Interesting Information!

S.Q. How does buffer maintain pH of blood?

Ans. The body's blood buffering system, involving bicarbonates, helps maintain a stable pH — the composition of injections are buffered to ensure that no change in blood's pH occurs.

Quick Check 9.4

(a) Explain the impact of common ion effect on solubility.

Ans. The solubility of sparingly soluble salts decreases in the presence of common ions in a mixture. Common ion does not affect the solubility of completely soluble salt.

Solubility of $KC\ell O_3$: The solubility of less soluble salt $KC\ell O_3$ in water is suppressed by the addition of a more soluble salt $KC\ell$ by common ion effect.

$$KC\ell O_{3(s)} \rightleftharpoons K^+_{(aq)} + C\ell O_{3(aq)}^- \qquad KC\ell_{(s)} \rightleftharpoons K^+_{(aq)} + C\ell^-_{(aq)}$$

(b) To a saturated solution of $AgC\ell$, some $NaC\ell$ solution was added.

(i) When $NaC\ell$ is added to the saturated solution of $AgC\ell$, the concentration of Ag^+ decreases or becomes low at equilibrium.

$$NaC\ell_{(s)} \rightleftharpoons Na^+_{(aq)} + C\ell^-_{(aq)} \qquad AgC\ell_{(s)} \rightleftharpoons Ag^+_{(aq)} + C\ell^-_{(aq)}$$

(ii) $NaC\ell$ solution is a strong electrolyte and produces $C\ell^-$ ions more rapidly. Due to common ion ($C\ell^-$)

equilibrium shifts to the left. As a result, the concentration of Ag^+ ions becomes lowered and the solubility of AgCℓ decreases.

(c) How does a buffer maintain pH stability?

Ans. A Buffer maintains pH stability due to common ions. It maintains the pH and pOH when acid or base is added from outside. For example, buffer solution consisting of $NH_4C\ell$ and NH_4OH — when a base or OH^- ions are added, they will react with H_3O^+ to give back H_2O. The pH of solution will remain almost unchanged.

(d) Calculate the pH of a buffer consisting of 0.50 M HF and 0.45 M of a fluoride (F^-) salt before and after addition of 0.40 g NaOH to 1.0 dm^3 of the buffer (K_a of HF = 6.8×10^{-4}).

Ans. [Acid] = [HF] = 0.50 M, [Salt] = $[F^-]$ = 0.45 M, $K_a = 6.8\times10^{-4}$. Concentration of NaOH added = 0.40⁄40 = 0.01 M.

$pK_a = -\log(6.8\times10^{-4}) = 3.17$ $\quad pH = pK_a + \log([Salt]/[Acid]) = 3.17 + \log(0.45/0.50) = 3.17 - 0.046 = 3.12$

When 0.01 M NaOH is added to the buffer, acid is consumed while the concentration of salt (F^-) is increased.

$[HF] = 0.50 - 0.01 = 0.49$ M $\quad [F^-] = 0.45 + 0.01 = 0.46$ M

New $pH = pK_a + \log([Salt]/[Acid]) = 3.17 + \log(0.46/0.49) = 3.17 - 0.028 = 3.14$

(e) Calculate the pH of a buffer solution in which 0.11 molar CH_3COONa and 0.09 molar acetic acid solutions are present. K_a for CH_3COOH is 1.85×10^{-5}.

Ans. $[CH_3COONa]$ = [Salt] = 0.11 M, $[CH_3COOH]$ = [Acid] = 0.09 M, $K_a = 1.85\times10^{-5}$.

$pK_a = -\log(1.85\times10^{-5}) = 4.74$ $\quad pH = pK_a + \log([Salt]/[Acid]) = 4.74 + \log(0.11/0.09)$

pH = 4.74 + 0.087 = 4.83

Solubility Product

12. Define and explain solubility product. Give its application.

Definition: The solubility product is the product of the concentrations of ions raised to an exponent equal to the co-efficient of the balanced equation.

- The value of K_{sp} is a measure of the dissociation of sparingly soluble salt.
- The dissociation is completed at equilibrium stage for NaCℓ when it is mixed in water because it is a completely soluble salt.
- The dissociation is not complete at equilibrium stage in slightly soluble salts like $PbC\ell_2$, $PbSO_4$ etc.
- K_{sp} is only for slightly or sparingly soluble salts.

Slightly soluble ionic compounds

(i) Lead sulfate ($PbSO_4$): When it is shaken with water the solution contains Pb^{2+}, SO_4^{2-} and undissociated $PbSO_4$, it means that equilibrium exists between solid solute, $PbSO_4$ and the dissolved ions, SO_4^{2+} and Pb^{2+}.

$$PbSO_{4(s)} \text{ --}H_2O\text{--> } Pb^{2+}_{(aq)} + SO_{4(aq)}^{2-}$$

$$K_c = [Pb^{2+}][SO_4^{2-}]/[PbSO_4]$$

Being a sparingly soluble salt, the concentration of lead sulphate ($PbSO_4$) almost remains constant. Bring $[PbSO_4]$ on left hand side with K_c.

$K_c[PbSO_4] = [Pb^{2+}][SO_4^{2-}]$ $\quad$ If $K_c[PbSO_4] = K_{sp}$, $K_{sp} = [Pb^{2+}][SO_4^{2-}] = 1.6 \times 10^{-8}$ at 25 °C

K_{sp} is called the solubility product of $PbSO_4$. It is the product of molar solubilities of two ions at equilibrium stage.

(ii) Lead chloride ($PbC\ell_2$): It is a well-known sparingly soluble compound and it dissociates to a very small extent.

$$K_{sp} = [Pb^{2+}][C\ell^-]^2$$

Effect of temperature

K_{sp} is usually a very small quantity at room temperature. The value of K_{sp} is temperature dependent. It increases by increase in temperature. The following table shows the K_{sp} values of some slightly soluble ionic compounds.

Table 9.5 — K_{sp} values of some ionic compounds

Salt	Ion Product	K_{sp}	Salt	Ion Product	K_{sp}
AgBr	$[Ag^+][Br^-]$	5.0×10^{-13}	CuS	$[Cu^{2+}][S^{2-}]$	8×10^{-34}
AgCℓ	$[Ag^+][C\ell^-]$	1.8×10^{-10}	Fe_2S_3	$[Fe^{3+}]^2[S^{2-}]^3$	1.4×10^{-85}
$A\ell(OH)_3$	$[A\ell^+][OH^-]^3$	3×10^{-34}	$MgCO_3$	$[Mg^{2+}][CO_3^{2-}]$	3.5×10^{-8}

Salt	Ion Product	K_{sp}	Salt	Ion Product	K_{sp}
$BaCO_3$	$[Ba^{2+}][CO_3^{2-}]$	1.1×10^{-10}	MnS	$[Mn^{2+}][S^{2-}]$	3×10^{-11}
$CaCO_3$	$[Ca^{2+}][CO_3^{2-}]$	3.3×10^{-9}	$PbCrO_4$	$[Pb^{2+}][CrO_4^{2-}]$	2.3×10^{-13}
CaF_2	$[Ca^{2+}][F^-]^2$	3.2×10^{-11}	$PbSO_4$	$[Pb^{2+}][SO_4^{2-}]$	1.6×10^{-8}

Applications of solubility product

(a) Determination of Solubility from K_{sp}

- To calculate solubility from K_{sp}, the formula of the compound and K_{sp} value is required.
- The unknown molar solubility S is calculated and the concentration of the ions is determined.

Table 9.6 shows the relationship between the K_{sp} values and the solubility of some sparingly soluble compounds.

Table 9.6 — The relation between solubility and solubility products of some salts

Formula	No. of ions	K_{sp}	Solubility g dm^{-3}
$PbSO_4$	2	1.69×10^{-4}	1.3×10^{-4}
$Ca(OH)_2$	3	6.5×10^{-6}	1.175×10^{-2}
CaF_2	3	3.2×10^{-11}	2.0×10^{-4}
Ag_2CrO_4	3	2.6×10^{-12}	8.7×10^{-5}

(b) Common Ion Effect

- The presence of a common ion decreases the solubility of a slightly soluble ionic compound.
- To explain it, consider a saturated solution of $PbCrO_4$, which is a sparingly soluble ionic salt.

$$PbCrO_{4(aq)} \rightleftharpoons Pb^{2+}_{(aq)} + CrO_{4(aq)}^{2-}$$

Now add Na_2CrO_4 which is a soluble salt. CrO_4^{2-} is the common ion. It combines with Pb^{2+} to form more insoluble $PbCrO_4$. So the equilibrium is shifted to the left to keep K_{sp} constant.

(c) Predicting Precipitation

The solubility product can also help in predicting whether the precipitation of a salt will occur or not.

Example: The solubility product of $CaSO_4$ is 2×10^{-5}. If we add 10^{-2} mol dm^{-3} solution of Ca^{2+} to 10^{-2} mol dm^{-3} solution of SO_4^{2-} ions at 25 °C, the concentrations of each ionic species can be calculated as:

$$[Ca^{2+}] = [SO_4^{2-}] = 10^{-2}/2 = 5.0\times10^{-3} \text{ mol dm}^{-3}$$

$$[Ca^{2+}][SO_4^{2-}] = (5.0\times10^{-3})(5.0\times10^{-3}) = 2.5\times10^{-5} \text{ mol}^2 \text{ dm}^{-6} > K_{sp} \text{ of } CaSO_4$$

As the ionic product of concentrations is greater than K_{sp}, therefore $CaSO_4$ will precipitate out.

Ionic Product vs K_{sp} — predicting the state of a solution

Ionic Product	Type of Solution	Precipitation
$> K_{sp}$	Supersaturated	Yes
$= K_{sp}$	Saturated	No
$< K_{sp}$	Unsaturated	No

Quick Check 9.5

(a) The K_{sp} of silver chloride (AgCℓ) is 1.77×10^{-10} at 25 °C. A solution already contains 0.10 M AgCℓ. Find the solubility of AgCℓ if 0.10 M NaCℓ is added.

Ans. Let solubility of Ag^+ = S. $AgC\ell \rightleftharpoons Ag^+ + C\ell^-$. As NaCℓ is a strong electrolyte, it gives 0.10 M $C\ell^-$. Total $[C\ell^-]$ from AgCℓ and NaCℓ = S + 0.1.

Due to common ion effect, S is very small, so $[C\ell^-] \approx 0.1$ M.

$$K_{sp} = [Ag^+][C\ell^-] \qquad 1.77\times10^{-10} = [S][0.1] \qquad S = 1.77\times10^{-9} \text{ M}$$

(b) Predict whether $CaSO_4$ will be precipitated when 10^{-3} mol dm^{-3} of each of Ca^{2+} and SO_4^{2-} are mixed together.

$$\text{Ionic product} = [Ca^{2+}][SO_4^{2-}] = 10^{-3}\times10^{-3} = 10^{-6} \text{ mol}^2 \text{ dm}^{-6}$$

Ionic Product	Type of Solution	Precipitation

K_{sp} for $CaSO_4 = 2 \times 10^{-5}$. Since $10^{-6} < K_{sp}$, $CaSO_4$ will not be precipitated out.

(c) Solution of potassium carbonate is basic. (i) Explain why is the solution alkaline? (ii) Give the equation for the hydrolysis of the above salt.

Ans. (i) Solution of potassium carbonate is a derived basic salt because it is from weak acid (H_2CO_3) and strong base (KOH). So, it is alkaline in nature.

(ii) Hydrolysis of salt when it is dissolved in water, it dissociates:

$$K_2CO_3 \rightleftharpoons 2K^+ + CO_3^{2-} \qquad CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$$

CO_3^{2-} is conjugate base of H_2CO_3 (weak acid) and reacts with water to produce OH^-. K^+ is conjugate acid of KOH (strong base) and does not affect the pH. The solution is basic due to CO_3^{2-} ion as it produces OH^-.

Interesting Information!

S.Q. What is the role of calcium carbonate in the shell of Nautilus?

Ans. The shell of this nautilus is composed mainly of calcium carbonate. The nautilus adjusts conditions so shell material is formed when the concentration of calcium ions and carbonate ions in seawater are high enough to precipitate calcium carbonate.

Salt Hydrolysis

13. What is salt hydrolysis? How it explains the solutions of some salts are acidic, basic or neutral?

Definition: The breakdown of a salt (made up of weak acid or base) by reacting with water, resulting in the bond breaking in that salt which changes the pH of water, is called hydrolysis.

Explanation

When a salt dissolves in water, it dissociates into its constituent ions. These ions can interact with water, affecting the solution's pH depending on the nature of the acid and base from which it is derived. It is called salt hydrolysis.

(i) Salts of strong acids and strong bases

In salts of strong acids and strong bases like sodium chloride (NaCℓ).

- The conjugate base of a strong acid ($C\ell^-$ from HCℓ) is very weak and does not significantly react with water.
- The conjugate acid of a strong base (Na^+ from NaOH) is also very weak and does not significantly react with water.

$$NaC\ell_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + C\ell^-_{(aq)}$$

- Na^+ is the conjugate acid of NaOH (a strong base) and does not affect the pH.
- $C\ell^-$ is the conjugate base of HCℓ (a strong acid) and does not affect the pH.

The solution remains **neutral.**

(ii) Salts of strong acids and weak bases

In salts of strong acids and weak bases like ammonium chloride ($NH_4C\ell$).

- The conjugate base of the strong acid ($C\ell^-$ from HCℓ) does not react with water.
- The conjugate acid of the weak base (NH_4^+ from NH_3) reacts with water to produce H_3O^+ ions, making the solution acidic.

$$NH_4C\ell_{(s)} \rightleftharpoons NH_{4(aq)}^+ + C\ell^-_{(aq)} \qquad NH_{4(aq)}^+ + H_2O_{(\ell)} \rightleftharpoons NH_{3(g)} + H_3O^+_{(aq)}$$

The solution is **acidic** due to the ammonium (NH_4^+) ions.

(iii) Salts of weak acids and strong bases

In salts of weak acid and strong base like sodium acetate (CH_3COONa).

$$CH_3COONa_{(s)} + H_2O_{(\ell)} \rightleftharpoons CH_3COO^-_{(aq)} + Na^+_{(aq)}$$

- CH_3COO^- is the conjugate base of CH_3COOH (a weak acid) and reacts with water.

$$CH_3COO^-_{(aq)} + H_2O_{(\ell)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$$

- Na^+ is the conjugate acid of NaOH (a strong base) and does not affect the pH.

The solution is **basic** due to the acetate (CH_3COO^-) ions.

(iv) Salts of weak acids and weak bases

In salts of weak acids and weak bases like ammonium acetate ($CH_3COO^-\ NH_4^+$).

- The conjugate acid (NH_4^+) and the conjugate base (CH_3COO^-) both affect the pH.

$$CH_3COO^- NH_{4(aq)}^+ \rightleftharpoons NH_{4(aq)}^+ + CH_3COO^-_{(aq)}$$

- NH_4^+ hydrolyses to produce H_3O^+.

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

- CH_3COO^- hydrolyses to produce OH^-.

$$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH_{(aq)} + OH^-$$

The resultant pH of the solution depends on the relative strengths of the conjugate acid and conjugate base. The solution may be acidic, basic, or nearly neutral, depending on which reaction is more dominant.

Acid-Base Indicators

14. What are acid-base indicators? How is it selected for a particular reaction?

Definition: An indicator is a substance that changes colour to mark a titration's endpoint.

- Acid-base indicators exhibit one colour in acid and another in base.
- Most indicators used in acid-base titration are weak organic acids or weak organic bases.

Acid solution: In solution, a weak-acid indicator (HIn) can be represented by the equation below.

$$HIn \rightleftharpoons H^+ + In^-$$

In^- is the symbol of the anion part of the indicator. Because the reaction is reversible, both HIn and In^- are present. The colours displayed result from the fact that HIn and In^- have different colours.

Acidic solutions: In acidic solutions, any In^- ions that are present act as Bronsted bases and accept protons from the acid. The indicator is then present in largely unionized form, HIn.

Basic Solutions: In basic solutions, the OH^- ions from the base combine with the H^+ ions produced by the indicator.

- The indicator molecules further ionize to equalize the loss of H^+ ions.
- The indicator is thus present largely in the form of its anion, In^-. The solution now shows the base-indicating colour, which for litmus is blue.

Colour of indicators

Different indicators change colour at different pH values. The colour depends on the relative amount of HIn and In^- at a given pH.

Example: Methyl red changes from red to yellow between pH 4.4 and 6.2. At pH 4.4, the indicator exists mostly as HIn molecules, which appear red in the solution.

Selecting a suitable indicator

The two general criteria for an indicator to be used in a titration are:

- The pH at the end of the titration should be close to the indicator's neutral point.
- The indicator should indicate a sharp colour change near the equivalence point of the titration.

Each pH indicator changes colour over a defined range of pH, known as the indicator range. An indicator changes colour over a range of about 2 pH units.

Titration Curve and Equivalence Point

15. Explain the titration curve and equivalence point.

Definition of titration curve: A pH curve is a graph of the pH of the solution versus the volume of titrant added.

Definition of equivalence point: The equivalence point is the point at which the amount of titrant added is just enough to neutralize the analyte solution completely.

Explanation of titration curve

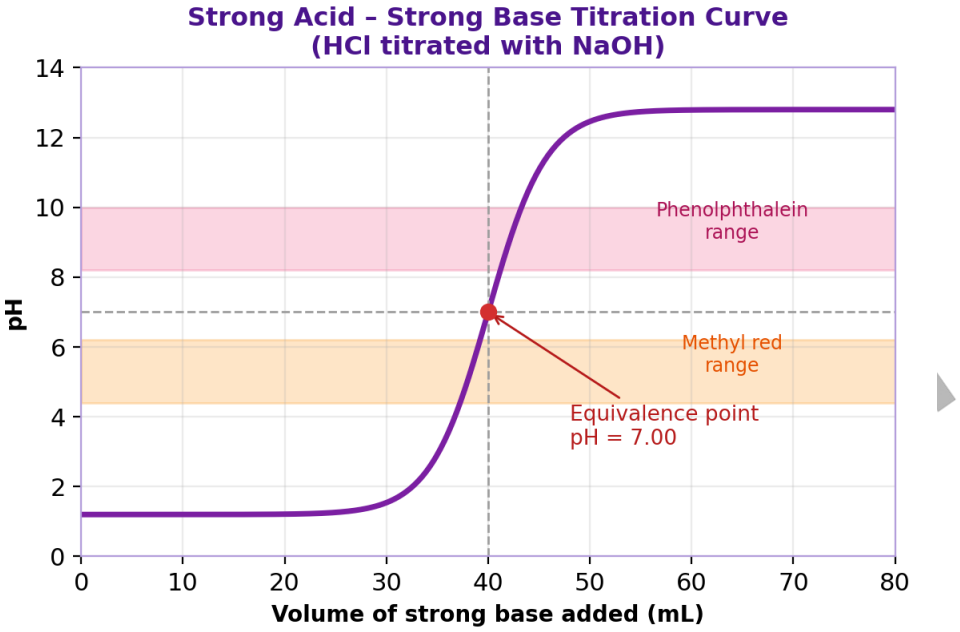
- Titration curves show how the pH of an acidic or basic solution changes as a basic or acidic solution is added to it.
- The titration curve helps to choose an indicator that will show when the titration is complete and we reach the equivalence point.
- The end point of the titration occurs when the indicator changes colour.
- We choose an indicator with an end point close to the equivalence point.

(i) Strong Acid–Strong Base Titration Curve

The titration of HCℓ with NaOH is an example of strong acid-strong base titration.

- The solution contains only the strong acid at the start.
- As the acid is strong, it completely dissociates leading to a high concentration of H^+ ions and the pH is very low.

- By the addition of NaOH, OH^- ions from NaOH begin to neutralize the H^+ ions from HCℓ and pH of the solution rises.
- The equivalence point is reached when the amount of OH^- added is stoichiometrically equal to the amount of H^+ originally present in the acid.
- The pH at the equivalence point in a strong acid-strong base titration is 7.
- The titration curve at this stage shows a steep rise in pH, changing quickly form acidic to neutral. The pH value at the endpoint changes about from 4.0 to 10.0.



Conclusion: A pH of 7.0 is at equivalence, so, phenolphthalein as indicator can be used as they show different colours in this range.

(ii) Strong Acid–Weak Base Titration Curve

The titration of NH_3 with HCℓ is an example of strong acid-weak base titration.

(a) Initial stage

- The solution contains only the weak base at start, dissociates partially, with a lower concentration of OH^- ions compared to a strong base.
- The initial pH will be greater than 7 but lower than the pH of a strong base.
- By the addition of HCℓ, the concentration of OH^- decreases, resulting in further decrease in pH.

(b) Change in pH

- Before reaching the equivalence point the solution is in a buffer-like region where the pH changes more gradually.
- The presence of the weak base and its conjugate acid (from the salt formed) creates a buffering effect, which helps to moderate the pH change as the base is added.

(c) Equivalence point

The pH at the equivalence point will be less than 7 (pH = 5.27) because the conjugate acid slightly dissociates, releasing H^+ ions making the solution acidic ($NH_4^+ \rightarrow NH_3 + H^+$).

(d) Titration curve

- After the equivalence point, the pH of the solution decreases rapidly. Because the strong acid dissociates completely in water, providing a high concentration of hydrogen ions (H^+), which significantly lowers the pH.
- Methyl orange has its colour change in this range, therefore it can be used as an indicator for strong acid-weak base titrations.

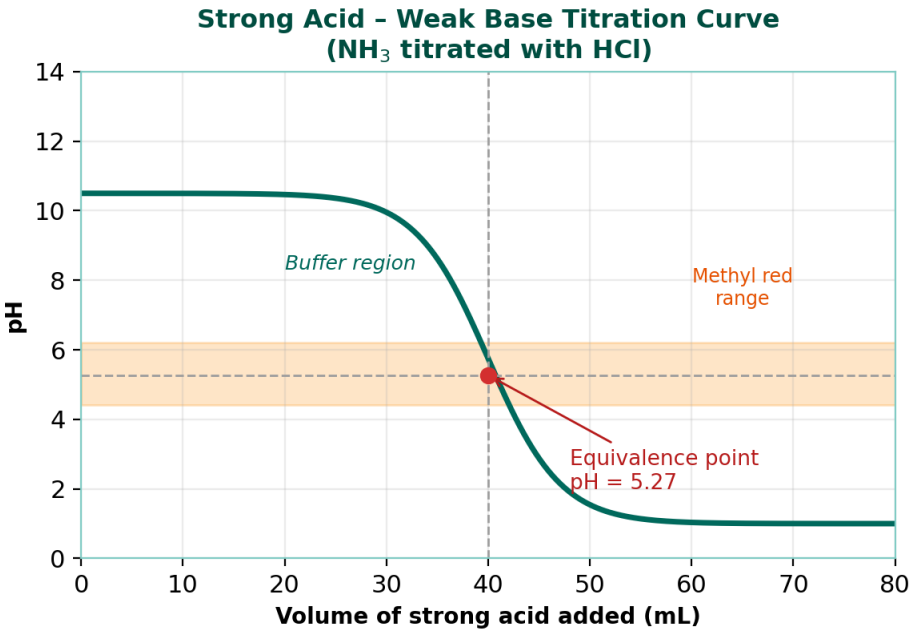


Table 9.7 — pH ranges of common indicators

Indicators	Acid colour	Base colour	pH Range	Type of Titration
Methyl orange	Orange	Yellow	3.2–4.5	Strong acid-strong base / Strong acid-weak base
Phenolphthalein	Colourless	Red	8.2–10.0	Weak acid-strong base

Quick Check 9.6

(a) Differentiate end point and equivalent point.

End point	Equivalence point
End point is the point at which titration is complete and indicator changes colour. End point is usually determined experimentally.	Equivalence point is the point at which the amount of titrant added is just enough to neutralize the analyte solution completely. Equivalent point is usually determined theoretically.

(b) Explain how an indicator changes its colour in acidic and basic solution.

Ans. Acid base indicator change colour due to shift in equilibrium of their acidic and basic form which have different colours. When indicator is added to a solution it can exist in either acidic or basic form, causing the indicator to display its colour.

(c) Suggest a suitable indicator for weak acid and strong base titration.

Ans. A suitable indicator for weak acid and strong base titration is phenolphthalein.

(d) Suggest a suitable indicator for weak acid and weak base titration.

Ans. A suitable indicator for weak acid and weak base titration is Bromothymol Blue.

SAMPLE PROBLEMS (SOLVED EXAMPLES)

Sample Problem 9.1

In a solution, the pH is 9.2. Determine the ionic product of water K_w at 25 °C.

Solution

Calculate the pOH from the pH: $pH + pOH = 14 \rightarrow pOH = 14 - 9.2 = 4.8$

Taking antilog of pH and pOH:

$$[H^+] = 10^{-pH} = 10^{-9.2} \approx 6.3 \times 10^{-10}\ M \qquad [OH^-] = 10^{-pOH} = 10^{-4.8} \approx 1.6 \times 10^{-5}\ M$$

$$K_w = [H^+][OH^-] = (6.3\times10^{-10})(1.6\times10^{-5}) = 1.01 \times 10^{-14}$$

Sample Problem 9.2

The ionic product of water at a certain temperature is $K_w = 1.0 \times 10^{-14}$ at 25 °C. If $[H^+] = 1.0 \times 10^{-7}$ M, calculate $[OH^-]$, pH and pOH.

Solution

$$K_w = [H^+][OH^-] \qquad 1.0\times10^{-14} = (1.0\times10^{-7})[OH^-] \qquad [OH^-] = 1.0 \times 10^{-7}\ M$$

$$pH = -\log(1.0\times10^{-7}) = 7 \qquad pOH = -\log(1.0\times10^{-7}) = 7$$

So the ionic product of water at the given temperature is $\approx 1.01 \times 10^{-14}$.

Sample Problem 9.3

Calculate the concentration of $[H_3O^+]$ in a 0.1 M (mol dm^{-3}) solution of acetic acid $[CH_3COOH]$, given that K_a is 1.8×10^{-5}.

Solution

$$[H_3O^+] \approx \sqrt{(K_a \times C)} \approx \sqrt{(1.8\times10^{-5} \times 0.1)} \approx \sqrt{(0.18\times10^{-6})} \approx 1.34 \times 10^{-3}\ M$$

The concentration $[H_3O^+]$ in the solution is approximately 1.34×10^{-3} mol dm^{-3}.

Sample Problem 9.4

$Ca(OH)_2$ is a sparingly soluble compound. Its solubility product is 6.5×10^{-6}. Calculate the solubility of $Ca(OH)_2$.

Solution

Let the solubility be represented by S in terms of mol dm^{-3}. The balanced equation is:

$$Ca(OH)_{2(aq)} \rightleftharpoons Ca^{2+}_{(aq)} + 2OH^-_{(aq)}$$

Initial stage: 0 + 0 Equilibrium stage: S + 2S

The concentration of OH^- is double that of Ca^{2+}, so:

$$K_{sp} = [Ca^{2+}][OH^-]^2 = S(2S)^2 = 4S^3$$

$$4S^3 = 6.5\times10^{-6} \qquad S^3 = 1.625\times10^{-6} \qquad S = (1.625)^{1/3} \times 10^{-2}$$

$$\mathbf{S = 1.18 \times 10^{-2}\ mol\ dm^{-3}}$$

Hence, at equilibrium stage 1.18×10^{-2} mol dm^{-3} of Ca^{2+} and $2\times1.18\times10^{-2} = 2.36\times10^{-2}$ mol dm^{-3} OH^- are present in the solution.