

Class: 11th

subject: Chemistry

Unit 9: ACID-BASE

CHEMISTRY

❖ Important MCQs:

1. According to Bronsted-Lowry theory, an acid is a substance which:

- (a) Accepts electron pair
- (b) Donates proton
- (c) Produces OH^- in water
- (d) Produces salt only

2. Which of the following acts as a Bronsted base?

- (a) HCl
- (b) NH_3
- (c) HNO_3
- (d) H_2SO_4

3. In the reaction $\text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$, the conjugate base is:

- (a) HCl
- (b) H_2O
- (c) H_3O^+
- (d) Cl^-

4. Which pair represents conjugate acid-base pair?

(a) HCl / H₂O

(b) NH₃ / NH₄⁺ ✓

(c) H₂O / HCl

(d) NaOH / NaCl

5. The substance that can act as both acid and base is called:

(a) Neutral

(b) Amphoteric ✓

(c) Basic

(d) Salt

6. In $\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^-$, water acts as:

(a) Base

(b) Acid ✓

(c) Salt

(d) Catalyst

7. Which indicator turns pink in basic solution?

(a) Litmus

(b) Methyl orange

(c) Phenolphthalein ✓

(d) Starch

8. Acids change blue litmus into:

(a) Green

(b) Red

(c) Yellow

(d) Pink

9. Neutralization reaction produces:

(a) Acid only

(b) Base only

(c) Salt and water

(d) Gas only

10. Which statement is TRUE about Bronsted-Lowry theory?

(a) Only works in water

(b) Limited to Arrhenius definition

(c) Based on proton transfer

(d) Only acids are defined

11. In the reaction $\text{HCl(g)} + \text{NH}_3\text{(g)} \rightarrow \text{NH}_4\text{Cl(s)}$, NH_3 acts as:

(a) Acid

(b) Base

(c) Salt

(d) Oxidizing agent

12. The conjugate acid of NH_3 is:

(a) NH_2^-

(b) NH_4^+

(c) NO_3^-

(d) N_2

13. Which of the following is amphoteric?

(a) HCl

(b) NaOH

(c) H_2O

(d) KCl

14. Arrhenius theory is limited because:

(a) It ignores salts

(b) It only applies in aqueous solutions

(c) It defines bases incorrectly

(d) It ignores indicators

15. A Bronsted base must have:

- (a) Lone pair of electrons ✓
- (b) Positive charge
- (c) Acidic nature
- (d) Salt formation ability

16. According to Lewis concept, an acid is a substance which:

- (a) Donates proton
- (b) Accepts electron pair ✓
- (c) Produces H^+
- (d) Neutralizes base



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17. A Lewis base is a species that:

- (a) Accepts proton
- (b) Donates electron pair ✓
- (c) Produces OH^-
- (d) Accepts H^+

18. A coordinate covalent bond is formed when:

- (a) Both atoms donate electrons
- (b) One atom donates electron pair ✓

(c) Both atoms accept electrons

(d) Electrons are transferred

19. In the reaction $\text{NH}_3 + \text{Ag}^+ \rightarrow [\text{Ag}(\text{NH}_3)_2]^+$, NH_3 acts as:

(a) Lewis acid

(b) Lewis base

(c) Proton donor

(d) Salt

20. In the same reaction, Ag^+ acts as:

(a) Lewis base

(b) Proton donor

(c) Lewis acid

(d) Neutral species

21. Which of the following is a Lewis acid?

(a) NH_3

(b) H_2O

(c) BF_3

(d) OH^-

22. BF_3 behaves as a Lewis acid because:

-
- (a) It has complete octet
- (b) It donates electrons
- (c) It has incomplete octet
- (d) It contains hydrogen

23. In $\text{BF}_3 + \text{F}^- \rightarrow \text{BF}_4^-$, F^- acts as:

- (a) Lewis acid
- (b) Lewis base
- (c) Proton donor
- (d) Neutral molecule

24. The charge on the product BF_4^- is:

- (a) +1
- (b) 0
- (c) -1
- (d) +2

25. Which statement is TRUE about Lewis acid-base reactions?

- (a) Always involve H^+ transfer
- (b) Only occur in water
- (c) Involve electron pair donation and acceptance

(d) Produce only salts

26. Pure water is a poor conductor of electricity because:

(a) It contains no ions

(b) Very few molecules ionize ✓

(c) It is non-polar

(d) It contains salts

27. The concentration of pure water is approximately:

(a) 1 mol dm^{-3}

(b) 10 mol dm^{-3}

(c) 55.5 mol dm^{-3} ✓

(d) 100 mol dm^{-3}



28. The ionic product of water is represented by:

(a) K_a

(b) K_b

(c) K_w ✓

(d) K_c

29. K_w is defined as:

(a) $[H^+]$

(b) $[\text{OH}^-]$

(c) $[\text{H}^+][\text{OH}^-]$ ✓

(d) $[\text{H}_2\text{O}]$

30. The value of K_w increases with:

(a) Decrease in temperature

(b) Increase in temperature ✓

(c) Pressure only

(d) Addition of acid

31. When an acid is added to water, K_w :

(a) Increases

(b) Decreases

(c) Remains constant ✓

(d) Becomes zero

32. In pure water, the concentration of H^+ and OH^- ions are:

(a) Unequal

(b) Equal ✓

(c) Zero

(d) Very high

33. The term pH was introduced by:

- (a) Arrhenius
- (b) Bronsted
- (c) Sorenson
- (d) Lewis

34. pH scale is used to express:

- (a) Temperature
- (b) Pressure
- (c) Acidity or basicity of solution
- (d) Concentration of salt

35. The concentration of H^+ and OH^- are difficult to express because they are:

- (a) Very high
- (b) Equal
- (c) Very low
- (d) Constant

36. The ionization constant (K_a) is a measure of:

- (a) Concentration

(b) Strength of acid ✓

(c) Temperature

(d) Pressure

37. For a weak acid HA, the equilibrium is:

(a) $HA \rightarrow H^+ + A^-$

(b) $HA + H_2O \rightleftharpoons H_3O^+ + A^-$ ✓

(c) $HA \rightleftharpoons H_2O + A^-$

(d) $HA \rightarrow H_2 + A^-$

38. An acid is considered weak if:

(a) $K_a > 1$

(b) $K_a = 1$

(c) $K_a < 10^{-3}$ ✓

(d) $K_a = 0$

39. A strong acid has K_a value:

(a) Less than 10^{-3}

(b) Equal to zero

(c) Greater than 1 ✓

(d) Equal to 10^{-7}

40. The concentration of H_3O^+ in weak acid solution is:

- (a) Equal to initial acid concentration
- (b) Greater than initial concentration
- (c) Less than initial concentration ✓
- (d) Zero

41. K_a expression for HA is:

- (a) $[\text{HA}]/[\text{H}^+][\text{A}^-]$
- (b) $[\text{H}^+][\text{A}^-]/[\text{HA}]$ ✓
- (c) $[\text{H}^+]/[\text{HA}]$
- (d) $[\text{A}^-]/[\text{H}^+]$



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42. The common ion effect refers to:

- (a) Increase in ionization
- (b) Suppression of ionization ✓
- (c) Formation of ions
- (d) Neutralization

43. When HCl is added to NaCl solution, NaCl:

- (a) Dissolves more
- (b) Remains unchanged

(c) Crystallizes out

(d) Forms gas

44. In common ion effect, addition of common ion causes:

(a) Increase in solubility

(b) Decrease in solubility

(c) No change

(d) Increase in temperature

45. The solubility of KClO_3 decreases when KCl is added due to:

(a) Neutralization

(b) Hydrolysis

(c) Common ion effect

(d) Oxidation



46. In H_2S solution, addition of HCl:

(a) Increases S^{2-} concentration

(b) Decreases S^{2-} concentration

(c) Has no effect

(d) Forms salt only

47. In NH_3 solution, addition of NH_4Cl causes:

-
- (a) Increase in OH^- concentration
 - (b) Decrease in OH^- concentration ✓
 - (c) No change
 - (d) Formation of acid

48. The common ion effect is useful in:

- (a) Organic synthesis
- (b) Qualitative analysis ✓
- (c) Nuclear reactions
- (d) Electrolysis

49. The ion suppressed in $\text{NH}_3 + \text{NH}_4\text{Cl}$ solution is:

- (a) H^+
- (b) NH_4^+
- (c) OH^- ✓
- (d) Cl^-

50. The common ion effect helps in formation of:

- (a) Gases
- (b) Precipitates ✓
- (c) Acids

(d) Bases

51. Buffer solutions are those which:

(a) Increase pH rapidly

(b) Resist change in pH

(c) Neutralize all acids

(d) Conduct electricity

52. A buffer solution maintains its pH even when:

(a) Heated

(b) Diluted

(c) Frozen

(d) Filtered



53. An acidic buffer is prepared by mixing:

(a) Strong acid + strong base

(b) Weak acid + its salt with strong base

(c) Weak base + strong base

(d) Strong acid + weak base

54. Which of the following is an acidic buffer?

(a) HCl + NaCl

(b) $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$ ✓

(c) $\text{NaOH} + \text{NaCl}$

(d) $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$

55. A basic buffer is prepared by mixing:

(a) Weak acid + salt

(b) Strong acid + salt

(c) Weak base + its salt with strong acid ✓

(d) Strong base + salt

56. Which of the following is a basic buffer?

(a) $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$

(b) $\text{HCl} + \text{NaCl}$

(c) $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$ ✓

(d) $\text{NaOH} + \text{NaCl}$

57. The pH of an acidic buffer is:

(a) Equal to 7

(b) Greater than 7

(c) Less than 7 ✓

(d) Zero

58. Buffer action is due to:

- (a) Neutralization
- (b) Common ion effect
- (c) Hydrolysis
- (d) Oxidation

59. When acid is added to an acidic buffer, it reacts with:

- (a) Acid
- (b) Salt (conjugate base)
- (c) Water
- (d) Base only



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60. When base is added to a buffer, it reacts with:

- (a) Salt
- (b) Water
- (c) Acid component
- (d) Indicator

61. The Henderson equation is:

- (a) $\text{pH} = \log[\text{H}^+]$
- (b) $\text{pH} = \text{pK}_a + \log\left(\frac{[\text{Salt}]}{[\text{Acid}]}\right)$

(c) $\text{pH} = \text{pKa} - \log[\text{H}^+]$

(d) $\text{pH} = \text{Kw}$

62. If $[\text{Salt}] = [\text{Acid}]$ in a buffer, then pH is equal to:

(a) 7

(b) pKa ✓

(c) 0

(d) 14

63. The pH of buffer depends on:

(a) Temperature only

(b) pKa and ratio of salt to acid ✓

(c) Pressure

(d) Volume only

64. Increasing salt concentration in acidic buffer will:

(a) Decrease pH

(b) Increase pH ✓

(c) No change

(d) Make solution neutral

65. Buffers are important in blood because they:

-
- (a) Increase acidity
 - (b) Maintain constant pH ✓
 - (c) Remove oxygen
 - (d) Increase pressure

66. The solubility product is represented by:

- (a) K_a
- (b) K_b
- (c) K_w
- (d) K_{sp} ✓

67. K_{sp} is defined as:

- (a) Product of ion concentrations at equilibrium ✓
- (b) Sum of ions
- (c) Difference of ions
- (d) Total concentration

68. For $PbCl_2$, the expression of K_{sp} is:

- (a) $[Pb^{2+}][Cl^-]$
- (b) $[Pb^{2+}][Cl^-]^2$ ✓
- (c) $[Pb^{2+}]^2[Cl^-]$

(d) $[\text{PbCl}_2]$

69. The value of K_{sp} is usually:

(a) Very large

(b) Zero

(c) Very small

(d) Equal to 1

70. K_{sp} depends on:

(a) Pressure

(b) Volume

(c) Temperature

(d) Colour

71. Solubility product is a measure of:

(a) Conductivity

(b) Dissociation of sparingly soluble salt

(c) Acidity

(d) Basicity

72. The solubility of a salt can be calculated from:

(a) K_a

(b) K_b

(c) K_w

(d) K_{sp} ✓

73. The presence of a common ion will:

(a) Increase solubility

(b) Decrease solubility ✓

(c) No effect

(d) Increase temperature

74. When Na_2CrO_4 is added to PbCrO_4 solution:

(a) Solubility increases

(b) Precipitation occurs ✓

(c) No change

(d) Gas forms

75. Precipitation occurs when:

(a) Ionic product $< K_{sp}$

(b) Ionic product $= K_{sp}$

(c) Ionic product $> K_{sp}$ ✓

(d) $K_{sp} = 0$

76. When a salt dissolves in water, hydrolysis means:

- (a) Neutralization
- (b) Reaction with water affecting pH ✓
- (c) Formation of gas
- (d) Oxidation

77. A salt of strong acid and strong base gives:

- (a) Acidic solution
- (b) Basic solution
- (c) Neutral solution ✓
- (d) Buffer solution



78. NaCl solution is:

- (a) Acidic
- (b) Basic
- (c) Neutral ✓
- (d) Buffer

79. NH₄Cl solution is:

- (a) Neutral
- (b) Basic

(c) Acidic

(d) Buffer

80. A salt of weak acid and strong base gives:

(a) Acidic solution

(b) Basic solution

(c) Neutral solution

(d) No solution

81. An acid-base indicator shows colour change due to:

(a) Temperature change

(b) Change in concentration

(c) Ionization equilibrium of indicator

(d) Formation of salt

82. In the equilibrium $\text{HIn} \rightleftharpoons \text{H}^+ + \text{In}^-$, the colour change occurs because:

(a) H^+ changes colour

(b) HIn and In^- have different colours

(c) In^- is colourless

(d) HIn is always colourless

83. In strongly acidic solution, the indicator will mainly exist as:

- (a) In^-
- (b) HIn ✓
- (c) OH^-
- (d) H_2O

84. In strongly basic solution, the indicator will mainly exist as:

- (a) HIn
- (b) In^- ✓
- (c) H^+
- (d) H_2O

85. The colour of litmus in basic medium is:

- (a) Red
- (b) Blue ✓
- (c) Yellow
- (d) Colourless

86. Which factor determines the colour of an indicator at a given pH?

- (a) Temperature

(b) Pressure

(c) Ratio of HIn to In^- ✓

(d) Volume

87. An indicator changes colour sharply when:

(a) pH is constant

(b) pH changes rapidly near equivalence point ✓

(c) Temperature increases

(d) Pressure decreases

88. The end point of titration is detected by:

(a) pH meter

(b) Indicator colour change ✓

(c) Temperature rise

(d) Gas evolution

89. The equivalence point is best defined as:

(a) Indicator change point

(b) Point of complete neutralization ✓

(c) Point of maximum pH

(d) Start of reaction

90. For accurate titration, the end point should be:

- (a) Far from equivalence point
- (b) Close to equivalence point
- (c) At neutral pH only
- (d) At high temperature

91. In strong acid–strong base titration, the pH at equivalence point is:

- (a) Less than 7
- (b) Greater than 7
- (c) Equal to 7
- (d) Zero

92. Which indicator is most suitable for strong acid–strong base titration?

- (a) Methyl orange
- (b) Phenolphthalein
- (c) Litmus
- (d) Starch

93. In titration of HCl with NaOH, pH increases because:

-
- (a) H^+ increases
 - (b) OH^- neutralizes H^+ ✓
 - (c) Salt forms
 - (d) Water evaporates

94. In strong acid-weak base titration, equivalence point pH is:

- (a) 7
- (b) Greater than 7
- (c) Less than 7 ✓
- (d) Equal to 14

95. The acidic nature at equivalence point (weak base titration) is due to:

- (a) Excess base
- (b) Strong acid
- (c) Conjugate acid formation ✓
- (d) Water formation

96. In $NH_3 + HCl$ titration, the buffer region exists:

- (a) After equivalence point
- (b) Before equivalence point ✓

(c) At neutralization

(d) At the end

97. Which indicator is best for strong acid–weak base titration?

(a) Phenolphthalein

(b) Methyl orange ✓

(c) Litmus

(d) Universal indicator

98. A titration curve shows:

(a) Temperature vs time

(b) pH vs volume added ✓

(c) Pressure vs volume

(d) Mass vs time



99. The steep vertical rise in titration curve represents:

(a) Start of reaction

(b) Buffer region

(c) Equivalence point region ✓

(d) End of solution

100. If wrong indicator is used in titration, the result will be:

-
- (a) Accurate
- (b) No change
- (c) Incorrect end point ✓
- (d) Neutral solution

❖ Important Short Questions:

1. Define Bronsted-Lowry acid.

Ans: A Bronsted-Lowry acid is a substance that donates a proton (H^+) in a chemical reaction. It releases H^+ ions into the solution. This definition is based on proton transfer concept. Acids can only act as acids when they lose H^+ .

Example: $HCl \rightarrow H^+ + Cl^-$

2. Define Bronsted-Lowry base.

Ans: A Bronsted-Lowry base is a substance that accepts a proton (H^+). It has lone pair of electrons to bond with H^+ ions. Bases remove H^+ from other substances. This is also called proton acceptor.

Example: $NH_3 + H^+ \rightarrow NH_4^+$

3. What is a conjugate acid-base pair? Give one example.

Ans: A conjugate acid-base pair consists of two species that differ by one proton (H^+). When an acid loses H^+ it forms its conjugate base.

When a base gains H^+ it forms its conjugate acid. These pairs exist in equilibrium reactions.

Example: NH_3 / NH_4^+

4. Why is water considered amphoteric?

Ans: Water is called amphoteric because it can act as both an acid and a base. It donates H^+ in some reactions and accepts H^+ in others. Its behavior depends on the substance it reacts with. This dual nature makes it very important in chemistry.

Example: $H_2O + HCl \rightarrow H_3O^+$

5. Give one reaction where water acts as a base.

Ans: Water acts as a base when it accepts a proton (H^+). In this case, it reacts with acids like HCl. Water gains H^+ and forms hydronium ion (H_3O^+). This shows its basic behavior in proton transfer reactions.

Example: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

6. Define Lewis acid.

Ans: A Lewis acid is a species that accepts a pair of electrons to form a coordinate bond. It does not necessarily contain hydrogen. It has empty orbitals to accept electrons. This concept is broader than Bronsted theory.

Example: $BF_3 + NH_3 \rightarrow BF_3 \leftarrow NH_3$

7. Define Lewis base.

Ans: A Lewis base is a species that donates a pair of electrons. It has lone pair of electrons available for bonding. It forms coordinate covalent bonds with Lewis acids. This explains many non-proton reactions.

Example: NH_3 donates electron pair to BF_3

8. Why is BF_3 considered a Lewis acid?

Ans: BF_3 is a Lewis acid because boron has an incomplete octet (only 6 electrons). It can accept an electron pair from another species. It has vacant orbitals for bonding. Therefore, it behaves as an electron pair acceptor.

Example: $\text{BF}_3 + \text{F}^- \rightarrow \text{BF}_4^-$

9. What is ionic product of water (K_w)?

Ans: Ionic product of water (K_w) is the product of concentrations of H^+ and OH^- ions in water. It represents the extent of ionization of water. Its value at 25°C is 1×10^{-14} . It remains constant in all aqueous solutions at constant temperature.

Formula: $K_w = [\text{H}^+][\text{OH}^-]$

10. How does temperature affect K_w ?

Ans: The value of K_w increases with increase in temperature. This is because ionization of water is an endothermic process. Higher

temperature produces more H^+ and OH^- ions. Therefore, equilibrium shifts forward.

11. What is the relationship between $[\text{H}^+]$ and $[\text{OH}^-]$ in pure water?

Ans: In pure water, concentrations of H^+ and OH^- ions are equal. This is because water ionizes equally into both ions. At 25°C both are $1 \times 10^{-7} \text{ mol dm}^{-3}$. This makes pure water neutral in nature.

Relation: $[\text{H}^+] = [\text{OH}^-]$

12. Define pH.

Ans: pH is a scale used to measure the acidity or basicity of a solution. It is defined as the negative logarithm of hydrogen ion concentration. It tells how acidic or basic a solution is. Lower pH means more acidic and higher pH means more basic solution.

Formula: $\text{pH} = -\log[\text{H}^+]$

13. What is pOH?

Ans: pOH is the measure of hydroxide ion concentration in a solution. It is defined as the negative logarithm of OH^- concentration. It is used to express basicity of a solution. pH and pOH are related to each other.

Formula: $\text{pOH} = -\log[\text{OH}^-]$

14. Why is pH scale used?

Ans: pH scale is used because H^+ and OH^- concentrations are very small and difficult to handle. It simplifies calculations by using logarithmic

values. It provides a convenient way to measure acidity and basicity. It ranges from 0 to 14.

15. What is Ka (ionization constant of acid)?

Ans: Ka is the equilibrium constant for the ionization of a weak acid in water. It shows how much an acid dissociates into ions. Higher Ka means more ionization. It is used to measure acid strength.

Formula: $K_a = \frac{[H^+][A^-]}{[HA]}$

16. How does Ka indicate strength of an acid?

Ans: Ka value directly shows acid strength. A large Ka means acid ionizes more and is strong. A small Ka means acid ionizes less and is weak. So, strength increases with increasing Ka value.

17. What is meant by weak acid?

Ans: A weak acid is an acid that does not completely ionize in water. It partially dissociates into ions in equilibrium. Most molecules remain undissociated. Example includes acetic acid.

Example: CH_3COOH

18. Define common ion effect.

Ans: Common ion effect is the suppression of ionization of a weak electrolyte due to addition of a strong electrolyte having a common ion. It shifts equilibrium according to Le Chatelier's principle. It reduces solubility or ionization.

19. How does common ion effect influence solubility?

Ans: Common ion effect decreases the solubility of a sparingly soluble salt. Addition of a common ion shifts equilibrium to the left. This reduces ion formation in solution. As a result, precipitation increases.

20. Give one example of common ion effect.

Ans: When HCl is added to NH_3 solution, NH_3 ionization is suppressed due to NH_4^+ ion. Similarly, adding NaCl to AgCl solution reduces solubility of AgCl. This shows common ion effect clearly.

Example: $\text{NH}_3 + \text{NH}_4\text{Cl}$ solution

21. What is a buffer solution?

Ans: A buffer solution is one that resists change in pH when small amounts of acid or base are added. It maintains almost constant pH. It is made of weak acid/base and its salt. Buffers are important in biological systems.

22. How is an acidic buffer prepared?

Ans: An acidic buffer is prepared by mixing a weak acid with its salt. The salt is formed from a strong base. This combination resists pH change in acidic range. Example is acetic acid and sodium acetate.

Example: $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$

23. State Henderson equation for buffer pH.

Ans: Henderson equation is used to calculate pH of buffer solutions. It relates pH with pKa and ratio of salt to acid. It helps in buffer preparation. It is widely used in analytical chemistry.

Formula: $\text{pH} = \text{pKa} + \log\left(\frac{[\text{Salt}]}{[\text{Acid}]}\right)$

24. What is solubility product (Ksp)?

Ans: Ksp is the product of the concentrations of ions of a sparingly soluble salt at equilibrium. Each concentration is raised to its stoichiometric coefficient. It is constant at a given temperature. It determines solubility.

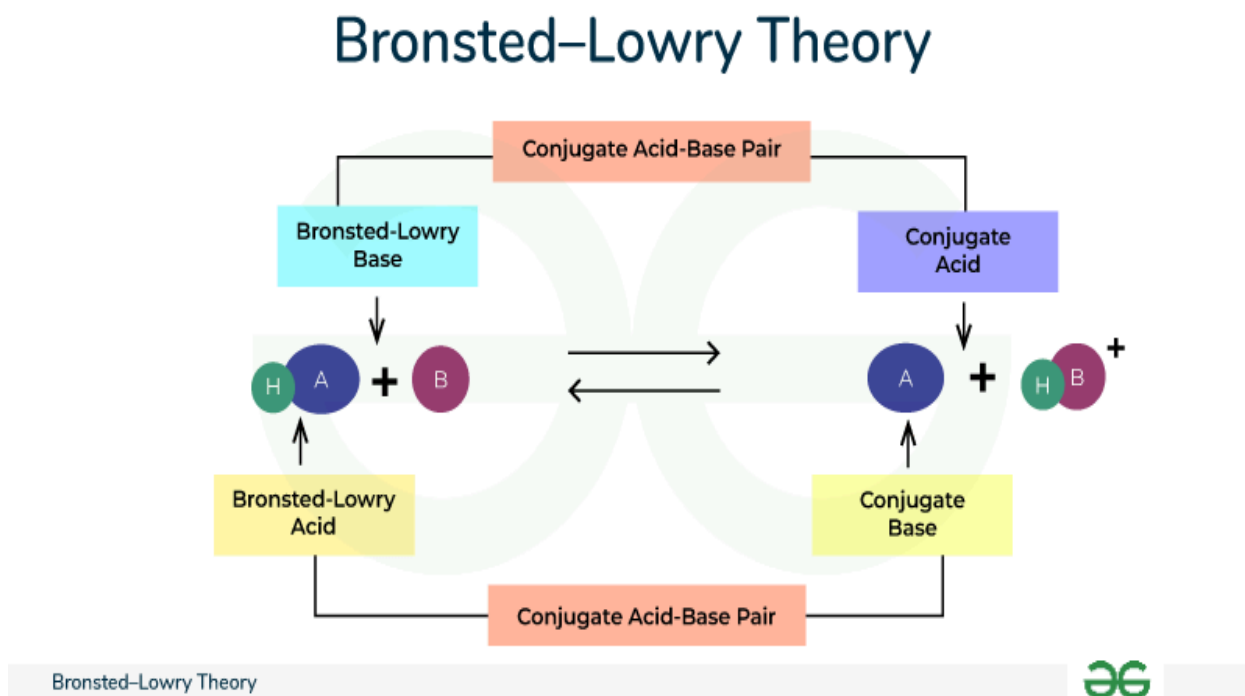
Example: $\text{PbCl}_2 \rightarrow [\text{Pb}^{2+}][\text{Cl}^-]^2$

25. What is salt hydrolysis?

Ans: Salt hydrolysis is the reaction of salt ions with water to produce acidic or basic solutions. It depends on strength of acid and base forming the salt. It affects pH of solution. Some salts give neutral, acidic or basic solutions.

❖ Important Long Questions:

☀ Q1. Explain Bronsted-Lowry concept of acids and bases. Also give suitable examples to support your answer.



❖ Answer:

The Bronsted-Lowry concept defines acids and bases in terms of proton (H^+) transfer reactions. According to this theory, an acid is a substance that donates a proton (H^+) in a chemical reaction, while a base is a substance that accepts a proton (H^+). This concept was introduced to overcome the limitations of the Arrhenius theory and can be applied in both aqueous and non-aqueous systems.

When an acid donates a proton, it forms its conjugate base, and when a base accepts a proton, it forms its conjugate acid. Therefore, acid-base reactions always involve two conjugate acid-base pairs.

For example, when hydrochloric acid dissolves in water: $\text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$

In this reaction, HCl donates a proton, so it acts as a Bronsted-Lowry acid, while water accepts the proton, so it acts as a Bronsted-Lowry base.

Another example is the reaction between ammonia and water: $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$

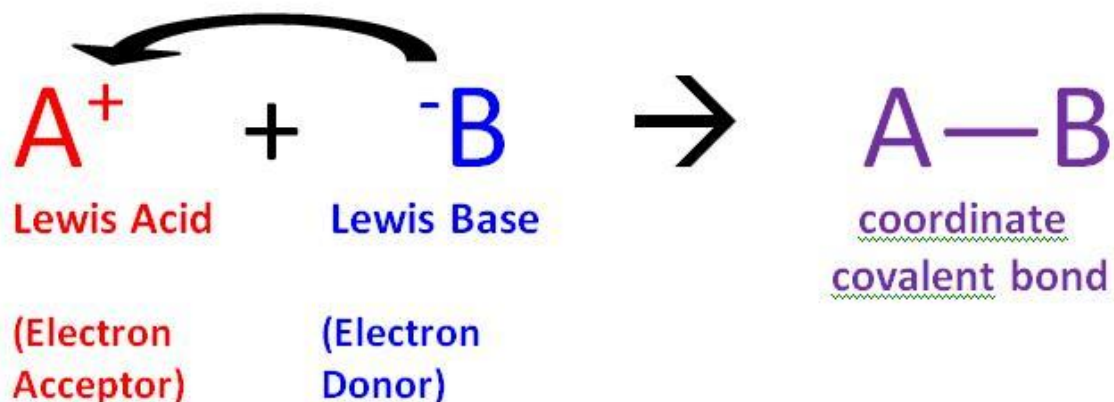
Here, NH_3 accepts a proton and acts as a base, while water donates a proton and acts as an acid.

This theory is very important because it explains acid-base behavior in different solvents and even in gas-phase reactions.

✓ Key Exam Points

- Acid = proton donor
- Base = proton acceptor
- Always forms conjugate acid-base pairs
- Applicable in aqueous and non-aqueous systems

★ **Q2. Describe Lewis concept of acids and bases. Explain with examples how coordinate covalent bond is formed.**

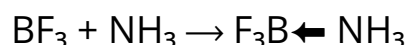


❖ **Answer:**

The Lewis concept of acids and bases explains acid-base behavior in terms of electron pair transfer instead of proton transfer. According to Lewis theory, a Lewis acid is a species (atom, ion, or molecule) that can accept a pair of electrons, while a Lewis base is a species that can donate a pair of electrons. This concept is more general than the Bronsted-Lowry theory because it does not require the presence of hydrogen ions.

A coordinate covalent bond is formed when both electrons in a covalent bond are donated by the same atom (the Lewis base) to an electron-deficient species (the Lewis acid).

For example, in the reaction between ammonia and boron trifluoride:



BF_3 acts as a Lewis acid because boron has an incomplete octet and can accept an electron pair. NH_3 acts as a Lewis base because nitrogen has

a lone pair of electrons that it donates to BF_3 , forming a coordinate covalent bond.

Another important example is: $\text{Ag}^+ + 2\text{NH}_3 \rightarrow [\text{Ag}(\text{NH}_3)_2]^+$

Here, Ag^+ is a Lewis acid and NH_3 is a Lewis base. Two coordinate bonds are formed between silver ions and ammonia molecules.

This concept is widely used in complex formation reactions and coordination chemistry.

✔ Key Exam Points

Lewis acid = electron pair acceptor

Lewis base = electron pair donor

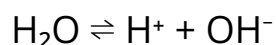
Coordinate bond = both electrons from one atom

Important in complex ion formation

☀ **Q3. Define ionic product of water (K_w). Discuss its importance and explain how temperature affects its value.**

❖ **Answer:**

The ionic product of water (K_w) is defined as the product of the concentrations of hydrogen ions (H^+) and hydroxide ions (OH^-) present in water at equilibrium. In pure water, a very small number of water molecules ionize to form these ions according to the equation:



Therefore, the ionic product of water is expressed as:

$$K_w = [H^+][OH^-]$$

At 25°C, the value of K_w is $1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$. In pure water, the concentrations of H^+ and OH^- are equal, each being $1.0 \times 10^{-7} \text{ mol dm}^{-3}$, which makes the solution neutral.

Importance of K_w

K_w is very important in chemistry because it helps in:

- Calculating pH and pOH of solutions
- Determining the acidic or basic nature of a solution
- Understanding ionization of water in different conditions
- Relating H^+ and OH^- concentrations in all aqueous solutions

Effect of Temperature on K_w

The value of K_w increases with increase in temperature. This is because the ionization of water is an endothermic process, meaning it absorbs heat. When temperature increases, more water molecules ionize, producing more H^+ and OH^- ions. As a result, K_w increases.

For example:

- At 0°C → K_w is smaller
- At 25°C → $K_w = 1.0 \times 10^{-14}$
- At 100°C → K_w becomes much larger

Even though K_w increases with temperature, pure water remains neutral because $[H^+]$ is always equal to $[OH^-]$.

✓ **Key Exam Points**

- $K_w = [H^+][OH^-]$
- Constant at fixed temperature
- Increases with temperature
- Water remains neutral because $[H^+] = [OH^-]$

★ **Q4. Explain pH and pOH scales. Why are these scales important in chemistry? Describe their applications.**

❖ **Definition of pH**

The pH scale is a logarithmic scale used to measure the acidity or basicity of a solution. It is defined as the negative logarithm of hydrogen ion concentration:

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

A solution is acidic when $pH < 7$, neutral when $pH = 7$, and basic when $pH > 7$. Lower pH indicates higher acidity.

Definition of pOH

The pOH scale measures the hydroxide ion concentration in a solution. It is defined as:

$$pOH = -\log[OH^-]$$

It is used to express the basic nature of a solution.

Relationship between pH and pOH

At 25°C, pH and pOH are related by the equation:

$$\text{pH} + \text{pOH} = 14$$

If pH increases, pOH decreases and vice versa.

Importance of pH and pOH Scales

pH and pOH scales are important because the concentrations of H^+ and OH^- ions are very small and difficult to handle in normal form. These scales simplify calculations using logarithms. They are widely used to determine whether a solution is acidic, basic, or neutral.

Applications of pH and pOH

- In agriculture to check soil acidity for better crop growth
- In medicine to maintain blood pH around 7.4
- In industry for chemical production and quality control
- In water treatment for safe drinking water
- In food industry for taste, preservation, and safety

☀ Q5. Buffer Solutions: Definition, Types, Action and Henderson Equation

❖ Definition of Buffer Solutions

Buffer solutions are solutions that resist changes in pH when small amounts of acid or base are added to them. They maintain a nearly constant pH even on dilution or upon standing for a long time. Buffers play an important role in chemical and biological systems.

Types of Buffer Solutions

1. Acidic Buffer

An acidic buffer is formed by mixing a weak acid and its salt with a strong base. It has a pH less than 7.

Example: Acetic acid (CH_3COOH) + Sodium acetate (CH_3COONa)

2. Basic Buffer

A basic buffer is formed by mixing a weak base and its salt with a strong acid. It has a pH greater than 7.

Example: Ammonium hydroxide (NH_4OH) + Ammonium chloride (NH_4Cl)

How Buffers Resist Change in pH

Buffer solutions resist pH change due to the presence of a weak acid/base and its conjugate salt.

When acid is added:

The extra H^+ ions are neutralized by the basic component of the buffer, so pH does not change significantly.

When base is added:

The OH⁻ ions are neutralized by the acidic component of the buffer, preventing large changes in pH.

Thus, the buffer system maintains equilibrium and keeps pH almost constant.

Henderson Equation

The pH of a buffer solution is calculated using the Henderson equation:

$$\text{pH} = \text{pKa} + \log\left(\frac{[\text{Salt}]}{[\text{Acid}]}\right)$$

This equation shows that buffer pH depends on:

- pKa of weak acid
- Ratio of salt to acid concentration

When [Salt] = [Acid], then pH = pKa.

◆ Summary:

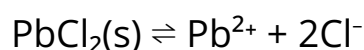
Buffer solutions are special solutions that resist changes in pH. They are of two types: acidic buffers and basic buffers. Their pH stability is due to neutralization of added acids or bases by buffer components. The Henderson equation is used to calculate buffer pH and is very important in chemistry and biological systems such as blood regulation.

☀ Q6. Solubility Product (K_{sp}), Common Ion Effect and its Applications

❖ Definition of Solubility Product (K_{sp})

The solubility product (K_{sp}) is the product of the molar concentrations of ions of a sparingly soluble salt in a saturated solution at equilibrium. Each ion concentration is raised to the power of its coefficient in the balanced chemical equation. It is a constant at a given temperature.

For example, for PbCl₂:



$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{Cl}^-]^2$$

K_{sp} represents the extent of dissociation of a sparingly soluble salt.

Common Ion Effect

The common ion effect is the suppression of ionization or solubility of a weak electrolyte or sparingly soluble salt due to the addition of a strong electrolyte containing a common ion.

According to Le Chatelier's principle, the equilibrium shifts in the backward direction to reduce the added ion concentration.

Example of Common Ion Effect

For PbCrO₄:



If Na_2CrO_4 is added, CrO_4^{2-} (common ion) increases, so equilibrium shifts left and more PbCrO_4 precipitates.

Applications of Common Ion Effect

1. Precipitation of Salts

Common ion effect is used to decrease solubility and cause precipitation of sparingly soluble salts. When ion concentration increases beyond K_{sp} , precipitation occurs.

2. Qualitative Analysis

It is widely used in salt analysis to separate different groups of ions. **For example:**

H_2S is used to precipitate Group II metal ions by controlling S^{2-} concentration using common ion effect.

3. Purification of Salts

Pure crystals of salts can be obtained by reducing solubility using a common ion.

4. Buffer and Industrial Processes

It is also useful in buffer solutions and industrial chemical processes where control of ion concentration is required.

Condition for Precipitation

Precipitation occurs when:

Ionic product > K_{sp}

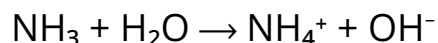
◆ **Summary:**

Solubility product (K_{sp}) is a constant that represents the equilibrium between a sparingly soluble salt and its ions. The common ion effect reduces solubility by shifting equilibrium and is widely used in precipitation reactions and qualitative analysis of ions.

 **Exercise**

Q.1 Four choices are given for each question. Select the correct choice

I. Given the following reaction:



- (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

II. The pH of $10^{-3} \text{ mol dm}^{-3}$ of an aqueous solution of H_2SO_4 is:

- (a) 3.0
- (b) 2.7

(c) 2.0

(d) 1.5

III. The solubility product of AgCl is $2.0 \times 10^{-10} \text{ mol}^2 \text{ dm}^{-6}$. The maximum concentration of Ag^+ ions in the solution is:

(a) $2 \times 10^{-10} \text{ mol dm}^{-3}$

(b) $1.41 \times 10^{-5} \text{ mol dm}^{-3}$

(c) $1.0 \times 10^{-10} \text{ mol dm}^{-3}$

(d) $4 \times 10^{-20} \text{ mol dm}^{-3}$

IV. Which indicator is typically used for titrations involving strong acids and strong base?

(a) Methyl red

(b) Phenolphthalein

(c) Bromothymol blue

(d) Litmus solution

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

(a) $\text{OH}^- (\text{aq})$

(b) $\text{H}^+ (\text{aq})$

(c) $\text{H}_2\text{O} (\text{l})$

(d) H_3O^+ (aq)

VI. 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

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

(a) pH of the solution is 7.0

(b) The indicator changes colour

(c) Equal volumes of acid and base have been added

(d) Acid and base are completely neutralized ✓

VIII. If the concentration of Cl^- ion in a solution is increased, the solubility of AgCl will:

(a) Decrease ✓

(b) Increase

(c) Remain unchanged

(d) Become zero

IX. 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

X. If the pH of a solution is 11, what is the $[\text{OH}^-]$ concentration in the solution?

(a) $1 \times 10^{-3} \text{ M}$

(b) $1 \times 10^{-11} \text{ M}$

(c) $1 \times 10^{-2} \text{ M}$

(d) $1 \times 10^{-14} \text{ M}$

XI. 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

Q.2 Attempt the following short-answer questions:

a. Define the following with example

i) Ionization constant (K_a)

Ionization constant (K_a) is the equilibrium constant that shows the extent of ionization of an acid in water. It indicates the strength of an acid. Higher K_a means stronger acid.

Example: For CH₃COOH

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

ii) Solubility product (K_{sp})

Solubility product is the product of concentrations of ions of a sparingly soluble salt at equilibrium. Each ion concentration is raised to its coefficient. It shows solubility of a salt.

Example: $AgCl \rightleftharpoons Ag^+ + Cl^-$

$$K_{sp} = [Ag^+][Cl^-]$$

iii) Common ion effect

It is the suppression of ionization of a weak electrolyte when a strong electrolyte with a common ion is added. It shifts equilibrium backward.

Example: Addition of NaCl decreases solubility of AgCl.

iv) Acid-base indicator

An acid-base indicator is a substance that changes colour depending on pH of solution. It is used to detect end point in titration.

Example: Phenolphthalein (colourless → pink)

b. Differentiate between

i) Hydrolysis vs Dissolution

Hydrolysis is the reaction of salt ions with water to produce acidic or basic solution. It changes the pH of solution.

Dissolution is simply the process of dissolving a substance in water without chemical reaction. It does not necessarily change pH.

ii) Acidic vs Basic buffer solutions

An **acidic** buffer consists of a weak acid and its salt with a strong base and has $\text{pH} < 7$.

Basic buffer consists of a weak base and its salt with strong acid and has $\text{pH} > 7$.

Example: $\text{CH}_3\text{COOH}/\text{CH}_3\text{COONa}$ (acidic), $\text{NH}_4\text{OH}/\text{NH}_4\text{Cl}$ (basic)

c. Conjugate acid-base pairs

A conjugate acid-base pair consists of two species that differ by one proton (H^+). When an acid donates a proton, it forms its conjugate base. When a base accepts a proton, it forms its conjugate acid.

Example: $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$

Here $\text{NH}_3/\text{NH}_4^+$ is a conjugate pair.

d. Relationship between acid and conjugate base strength

The strength of an **acid** is inversely related to the strength of its **conjugate** base. A strong **acid** has a weak **conjugate** base because it easily donates H^+ . A weak acid has a strong conjugate base.

Example: HCl (strong acid) \rightarrow Cl^- (weak base)

e. Identify types of bases (Arrhenius, Bronsted-Lowry, Lewis)

i) $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$

$NaOH$ produces OH^- ions in aqueous solution, so it is an Arrhenius base.

The OH^- ion can accept a proton (H^+), so it is also a Bronsted-Lowry base.

Additionally, OH^- has a lone pair of electrons which it can donate, making it a Lewis base.

✔ **Conclusion:** $NaOH$ is an Arrhenius, Bronsted-Lowry, and Lewis base

ii) $HF + H_2O \rightleftharpoons F^- + H_3O^+$

In this reaction, H_2O accepts a proton from HF , so it acts as a Bronsted-Lowry base.

H_2O also has lone pairs of electrons, so it can donate an electron pair, making it a Lewis base.

It is not an Arrhenius base because it does not produce OH^- ions directly.

✓ **Conclusion:** H₂O is a Bronsted-Lowry and Lewis base



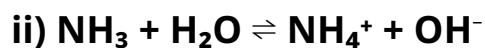
- NH₃ accepts a proton (H⁺), so it is a Bronsted-Lowry base.
- It donates a lone pair of electrons to form a bond with H⁺, so it is also a Lewis base.
- It is not an Arrhenius base because it does not produce OH⁻ ions.

✓ **Conclusion:** NH₃ is a Bronsted-Lowry and Lewis base

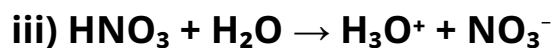
f. Identify whether water acts as an acid or base



Water accepts a proton from HCl, so it acts as a base.



Water donates a proton to NH₃, so it acts as an acid.



Water accepts a proton from HNO₃, so it acts as a base.



Water accepts a proton from acetic acid, so it acts as a base.

g) Which salt dissolves more readily in acidic solution: BaCO₃ or CuS? Explain.

We compare their solubility using K_{sp} values:

-
- K_{sp} of $\text{BaCO}_3 = 1.1 \times 10^{-10}$
 - K_{sp} of $\text{CuS} = 8 \times 10^{-34}$

Effect of acid on solubility

In acidic solution, H^+ ions react with anions:

For carbonate: $\text{CO}_3^{2-} + 2\text{H}^+ \rightarrow \text{H}_2\text{CO}_3 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$

For sulfide: $\text{S}^{2-} + 2\text{H}^+ \rightarrow \text{H}_2\text{S}$

This removes anions from solution and shifts equilibrium forward, increasing solubility.

Comparison

- BaCO_3 has higher K_{sp} , so it is relatively more soluble.
- CO_3^{2-} is easily removed by acid (forms CO_2 gas).
- CuS has extremely low K_{sp} (very insoluble).
- S^{2-} is not removed as effectively as carbonate under normal conditions.

Final Answer

BaCO_3 dissolves more readily in acidic solution than CuS .

Conclusion

- Acid increases solubility by removing anions
- $\text{BaCO}_3 \rightarrow$ soluble in acid (CO_2 gas forms)
- $\text{CuS} \rightarrow$ extremely insoluble due to very low K_{sp}

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

The common ion effect decreases solubility because the addition of a common ion shifts the equilibrium towards the undissolved solid. According to Le Chatelier's principle, the system tries to oppose the increase in ion concentration. As a result, ionization of the salt is suppressed and its solubility decreases.

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

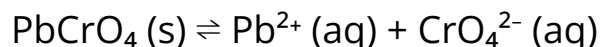
The solubility product (K_{sp}) is the product of the molar concentrations of ions of a sparingly soluble salt in a saturated solution at equilibrium, each raised to their respective powers. It remains constant at a given temperature.

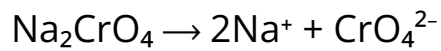
Factors affecting solubility product:

- Temperature (most important factor)
- Nature of the salt
- Nature of solvent

(Note: Common ion affects solubility, not K_{sp} value)

j. What happens when Na_2CrO_4 is added to saturated solution of PbCrO_4 ? (Explain with equations)





Addition of Na_2CrO_4 increases the concentration of CrO_4^{2-} ions (common ion). According to Le Chatelier's principle, equilibrium shifts to the left. This causes more PbCrO_4 to precipitate and its solubility decreases.

k. According to Lewis acid-base concept, boron trifluoride (BF_3) acts as an acid.

According to Lewis theory, an acid is an electron pair acceptor. In BF_3 , boron has an incomplete octet with only six electrons in its valence shell. Therefore, it can accept a pair of electrons from a donor species, making BF_3 a Lewis acid.



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

pH is calculated using the formula:

$$\text{pH} = -\log[\text{H}^+]$$

Given: $[\text{H}^+] = 1 \times 10^{-5}$ M

Now substitute:

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

$$\text{pH} = 5$$

Final Answer:

pH = 5

❖ **DESCRIPTIVE QUESTIONS**

☀ **Q3. Describe the Brønsted–Lowry theory of acids and bases. Provide examples of conjugate acid–base pairs and explain clearly their relationship.**

❖ **Answer:**

The Brønsted–Lowry theory defines acids and bases in terms of proton (H^+) transfer reactions. It was proposed by J.N. Brønsted and T.M. Lowry in 1923. According to this theory, acid-base reactions are not limited to aqueous solutions but can occur in any medium.

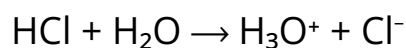
According to this theory:

- An acid is a substance that donates a proton (H^+).
- A base is a substance that accepts a proton (H^+).

This means every acid-base reaction involves the movement of a proton from one species to another.

Examples of Brønsted–Lowry acid-base reactions

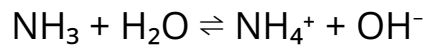
1. Hydrochloric acid in water



In this reaction:

-
- HCl donates a proton → acts as an acid
 - H₂O accepts a proton → acts as a base

2. Ammonia in water



In this reaction:

- NH₃ accepts a proton → acts as a base
- H₂O donates a proton → acts as an acid

Conjugate Acid–Base Pairs

A conjugate acid–base pair consists of two species that differ by only one proton (H⁺).

- When an acid loses a proton, it forms its conjugate base.
- When a base gains a proton, it forms its conjugate acid.

Examples of conjugate pairs:

- HCl / Cl⁻
- NH₃ / NH₄⁺
- H₂O / OH⁻
- H₂O / H₃O⁺

Relationship between conjugate acid–base pairs

The strength of a conjugate pair is always inversely related:

- A strong acid forms a very weak conjugate base.

-
- A weak acid forms a relatively stronger conjugate base.
 - **Similarly**, a strong base forms a weak conjugate acid.

This relationship ensures that proton transfer reactions can proceed in both forward and reverse directions.

Conclusion

The Brønsted–Lowry theory explains acids and bases as proton donors and acceptors. It introduces the concept of conjugate acid–base pairs, which differ by one proton and have opposite strengths, making acid-base chemistry more general and widely applicable.

✨ Q4. Define the Lewis theory of acids and bases. How does this theory differ from the Brønsted–Lowry theory? Give examples of Lewis acids and bases that do not involve proton transfer.

❖ Answer:

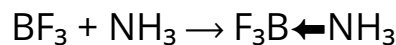
The Lewis theory of acids and bases was proposed by G.N. Lewis in 1923. According to this theory, acids and bases are defined on the basis of electron pair transfer rather than proton transfer.

- A Lewis acid is a species that accepts a pair of electrons.
- A Lewis base is a species that donates a pair of electrons.

When a Lewis acid reacts with a Lewis base, they form a coordinate covalent (dative) bond, in which both electrons are supplied by the base.

Examples of Lewis acid-base reactions (without proton transfer)

1. BF_3 and NH_3 reaction

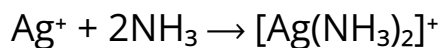


In this reaction:

- BF_3 has an incomplete octet and accepts an electron pair \rightarrow Lewis acid
- NH_3 has a lone pair and donates it \rightarrow Lewis base

A coordinate bond is formed between B and N atoms.

2. Silver ion and ammonia complex formation



In this reaction:

- Ag^+ accepts electron pairs \rightarrow Lewis acid
- NH_3 donates electron pairs \rightarrow Lewis base

This forms a complex ion without involving any proton transfer.

Difference between Lewis and Brønsted-Lowry theories

The Brønsted-Lowry theory defines acids and bases in terms of proton (H^+) transfer:

- Acid = proton donor
- Base = proton acceptor

The Lewis theory defines acids and bases in terms of electron pair transfer:

- Acid = electron pair acceptor
- Base = electron pair donor

Main difference

- Brønsted–Lowry theory always involves H^+ (proton transfer).
- Lewis theory does not require protons and is based only on electron pairs, so it explains a wider range of reactions.

Conclusion

The Lewis theory is a more general concept than Brønsted–Lowry theory because it explains acid-base reactions based on electron pair donation and acceptance, including many reactions where no proton transfer occurs.

🌟 **Q5. Discuss applications and implications of the common ion effect in various fields.**

❖ **Answer:**

The common ion effect is the suppression of the ionization of a weak electrolyte or the decrease in solubility of a sparingly soluble salt due to the addition of a strong electrolyte containing a common ion.

This effect has many important applications in chemistry and industry.

1. Purification of salts

The common ion effect is used to purify salts by reducing their solubility.

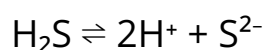
For example, pure sodium chloride (NaCl) is obtained by passing HCl gas through a saturated solution of NaCl (brine). The common ion (Cl⁻) decreases solubility, and NaCl crystallizes out.

2. Qualitative analysis in chemistry

It is widely used in salt analysis to control the concentration of ions.

Example:

In the presence of HCl, the concentration of S²⁻ ions is reduced, which helps in selective precipitation of metal sulphides in group analysis.



Addition of H⁺ (from HCl) shifts equilibrium to the left, reducing S²⁻ concentration.

3. Separation of ions by selective precipitation

Different ions can be separated based on their different solubility products (K_{sp}).

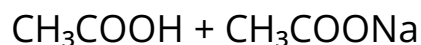
Example:

Ag⁺, Pb²⁺, and Hg₂²⁺ ions are separated by controlling Cl⁻ concentration, where only the least soluble salts precipitate first.

4. Buffer solutions

The common ion effect is the basic principle behind buffer solutions.

Example:



The common ion CH_3COO^- suppresses ionization of acetic acid and maintains stable pH.

5. Control of pH in biological systems

It helps maintain constant pH in biological fluids.

Example:

The blood buffer system ($\text{H}_2\text{CO}_3 / \text{HCO}_3^-$) uses the common ion effect to resist pH changes in the body.

6. Reduction of solubility in industrial processes

In industrial chemistry, solubility of salts is controlled using the common ion effect to obtain desired precipitation and product separation.

Conclusion

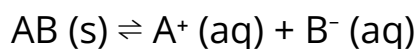
The common ion effect is a key principle in chemistry that controls solubility and ionization. It is widely used in salt purification, qualitative analysis, buffer systems, and industrial processes, making it very important in both laboratory and real-life applications.

☀ **Q6. What is the solubility product for sparingly soluble salts? Give its two applications.**

❖ **Answer:**

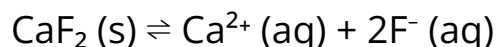
The solubility product (K_{sp}) is the equilibrium constant for a sparingly soluble salt in a saturated solution. It represents the product of the molar concentrations of ions of the salt, each raised to the power of their coefficients in the balanced equation.

For a general salt:



$$K_{sp} = [A^+][B^-]$$

For a salt like CaF_2 :



$$K_{sp} = [Ca^{2+}][F^-]^2$$

The value of K_{sp} is constant at a given temperature and indicates how soluble a salt is in water.

Applications of solubility product

1. To predict precipitation

- K_{sp} helps to predict whether a precipitate will form or not.
- If the ionic product (Q_{sp}) is greater than K_{sp} , precipitation occurs.
- If Q_{sp} is less than K_{sp} , no precipitation occurs.

Example:

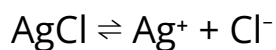
If $[Ag^+][Cl^-] > K_{sp}$ of AgCl, then AgCl precipitates.

2. To calculate solubility of salts

K_{sp} is used to calculate the molar solubility of sparingly soluble salts in water.

Example:

For AgCl,



If solubility = S , then

$$K_{sp} = S \times S = S^2$$

$$\text{So, } S = \sqrt{K_{sp}}$$

Conclusion

The solubility product is a constant that defines the equilibrium between a sparingly soluble salt and its ions in solution. It is mainly used to predict precipitation and calculate solubility in chemical systems.

✨ **Q7. 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?**

❖ **Answer:**

A titration curve for a strong acid–strong base reaction is a graph between pH of the solution and the volume of strong base added.

At the start, the solution contains only a strong acid (e.g., HCl). Since it completely ionizes in water, the concentration of H⁺ ions is high, so the pH is very low (about 1–3).

As a strong base (e.g., NaOH) is added, OH⁻ ions neutralize H⁺ ions: H⁺ + OH⁻ → H₂O

Due to this neutralization, the pH increases slowly at first. However, near the equivalence point, a very sharp and sudden increase in pH occurs because almost all H⁺ ions have been neutralized.

After the equivalence point, excess OH⁻ ions remain in solution, so the pH becomes basic and increases gradually.

Shape of the curve (description)

- Starts at low pH (strong acid region)
- Gradual rise in pH
- Very steep vertical jump near equivalence point
- Ends at high pH (strong base region)

Identification of equivalence point

The equivalence point is the point where the number of moles of acid equals the number of moles of base (complete neutralization).

On the graph, it is identified by:

-
- The middle of the steep vertical rise in pH
 - At this point, the solution is neutral (pH \approx 7) for strong acid–strong base titration
 - It is the point of maximum slope on the curve

Conclusion

The titration curve of a strong acid with a strong base shows a sharp pH change at the equivalence point, where complete neutralization occurs and the pH is approximately 7.

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