Acids, Bases and Salts — Complete Exercise Solution

Multiple Choice Questions (MCQs)

Q(i): According to Bronsted, Acids are:

Answer: a) Proton donor

Explanation: According to the Bronsted-Lowry theory, an acid is any substance that can donate a proton (H^+) to another substance. For example, HCl donates H^+ to water, so it is a Bronsted acid.

Q(ii): NH_3 is:

Answer: b) Base

Explanation: Ammonia (NH_3) has a lone pair of electrons on the nitrogen atom. It accepts a proton (H^+) from water to form NH_4^+, which is the behaviour of a base. According to both Bronsted-Lowry and Lewis theories, NH_3 is a base.

Q(iii): Neutral solution has a pH value:

Answer: c) 7

Explanation: The pH scale runs from 0 to 14. A pH of 7 means the concentration of H^+ ions and OH^- ions are exactly equal, making the solution perfectly neutral. Pure water at 25°C has a pH of 7.

Q(iv): Lower the pH value, ______ will be an acid.

Answer: b) Stronger

Explanation: pH is inversely related to the concentration of H^+ ions. The lower the pH, the higher the H^+ concentration, and therefore the stronger the acid. For example, HCl with pH = 1 is a much stronger acid than acetic acid with pH = 3.

Q(v): Salts are ______ compounds.

Answer: c) Neutral

Explanation: Salts are ionic compounds formed by the neutralisation reaction between an acid and a base. In their pure form, salts like NaCl are considered neutral compounds, though some salts can produce acidic or basic solutions when dissolved in water.

Q(vi): Those bases which give hydroxyl ions in water are called:

Answer: b) Alkalies

Explanation: Alkalies are a specific group of bases that dissolve in water and release hydroxide ions (OH^-). For example, NaOH dissolves in water to give Na^+ and OH^- ions. All alkalies are bases, but not all bases are alkalies.

Q(vii): $KHSO_4$ is a ______ salt:

Answer: b) Acidic

Explanation: $KHSO_4$ (Potassium hydrogen sulphate) still contains a replaceable hydrogen atom (H^+) in the HSO_4^- ion. Because it can still donate a proton, it behaves as an acidic salt. It is formed by the partial neutralisation of H_2SO_4 with KOH.

Short Answer Questions

Q(i): What are double salts?

Answer: Double salts are salts that are formed by the crystallisation of two different salts together in a fixed molar ratio. They contain two different cations or two different anions in their crystal structure. When dissolved in water, double salts dissociate completely into all their constituent ions and do not retain any special properties. A very common example is Mohr's Salt, which is $FeSO_4 \cdot (NH_4)_2SO_4 \cdot 6H_2O$. Another well-known example is Alum, which is $KAl(SO_4)_2 \cdot 12H_2O$. These salts exist only in the solid crystalline state and lose their identity completely upon dissolution in water.

Q(ii): What are amphoteric substances?

Answer: Amphoteric substances are those that can behave both as an acid and as a base depending on the nature of the substance they react with. In other words, they can donate a proton (acting as an acid) in one reaction and accept a proton (acting as a base) in another. Water (H_2O) is the most classic example of an amphoteric substance. When water reacts with a strong acid like HCl, it accepts a proton and acts as a base. When it reacts with a strong base like NH_3, it donates a proton and acts as an acid. Aluminium hydroxide $Al(OH)_3$ is another important example — it reacts with both acids and bases. Amino acids are also amphoteric in nature.

Q(iii): Differentiate between lone pair and bond pair of electrons.

Answer: A bond pair of electrons is a pair of electrons that is shared between two atoms to form a covalent bond. These electrons belong to both atoms simultaneously and are responsible for holding the atoms together. For example, in H_2O, each O–H bond contains one bond pair of electrons. A lone pair of electrons, on the other hand, is a pair of electrons that belongs to only one atom and is not involved in any bonding. These electrons are also called non-bonding pairs. For example, in the water molecule (H_2O), the oxygen atom has two lone pairs of electrons that are not shared with any other atom. Lone pairs play a very important role in Lewis acid-base theory, as they are donated by Lewis bases to Lewis acids.

Q(iv): What will be the pH and pOH of 0.001M NaOH solution?

Answer:
Given: Concentration of NaOH = 0.001, M = 10^{-3} M, NaOH is a strong base, so it completely dissociates: $NaOH \rightarrow Na^+ + OH^-$ Therefore, $[OH^-] = 10^{-3}$ M

Step 1 — Calculate pOH:

$pOH = -\log[OH^-]$
$pOH = -\log(10^{-3})$
$pOH = 3$

Step 2 — Calculate pH:

$pH + pOH = 14$
$pH = 14 - pOH$
$pH = 14 - 3$
$pH = 11$

Result: The pH of 0.001M NaOH = **11** and the pOH = **3**. Since pH > 7, the solution is basic, which makes sense for NaOH.

Q(v): Calculate the pH and pOH of 0.05M HCl solution.

Answer:
Given:
Concentration of HCl = 0.05
M = 5×10^{-2} M, HCl is a strong acid, so it completely dissociates:
$HCl \rightarrow H^+ + Cl^-$ Therefore, $[H^+] = 0.05$ M = 5×10^{-2} M

Step 1 — Calculate pH:

$pH = -\log[H^+]$
$pH = -\log(5 \times 10^{-2})$
$pH = -(\log 5 + \log 10^{-2})$
$pH = -(0.699 - 2)$
$pH = -(-1.301)$
$pH = 1.30$

Step 2 — Calculate pOH:
$pH + pOH = 14$
$pOH = 14 - 1.30$
$pOH = 12.70$

Result: The pH of 0.05M HCl = **1.30** and the pOH = **12.70**. Since pH < 7, the solution is acidic, which is expected for HCl.

Long / Detailed Questions

Q(i): What are salts and explain their different types.

Answer:

Definition of Salts Salts are ionic compounds that are formed when an acid reacts with a base in a neutralisation reaction. In this reaction, the hydrogen ions (H^+) from the acid combine with the hydroxide ions (OH^-) from the base to form water, while the remaining ions combine to form a salt. The general equation is: Acid + Base → Salt + Water

For example: $HCl + NaOH \rightarrow NaCl + H_2O$ Salts are made up of a positive ion (cation) from the base and a negative ion (anion) from the acid. They are generally solid crystalline substances that are soluble in water.

Types of Salts

1. Normal Salts (Neutral Salts) These are salts formed by the complete neutralisation of an acid by a base, where all the replaceable hydrogen atoms of the acid are replaced by metal ions. They do not contain any replaceable hydrogen or hydroxyl group.

Examples include NaCl (table salt), K_2SO_4, and Na_2CO_3. In aqueous solution, normal salts formed from strong acids and strong bases give a neutral solution with pH = 7.

2. Acidic Salts These are salts that are formed by the partial neutralisation of a polyprotic acid (an acid with more than one replaceable hydrogen atom). They still contain one or more replaceable hydrogen atoms. Because of this, they can donate protons and produce an acidic solution in water.

Examples include $NaHCO_3$ (sodium hydrogen carbonate), $NaHSO_4$ (sodium hydrogen sulphate), and $KHSO_4$. Their pH is less than 7 in solution.

3. Basic Salts These are salts formed by the partial neutralisation of a base by an acid. They still contain one or more hydroxyl (OH^-) groups that have not been neutralised. They produce a basic solution in water.

Examples include Pb(OH)Cl (lead hydroxychloride) and Mg(OH)Cl. Their pH is greater than 7 in solution.

4. Double Salts These are salts formed when two different simple salts crystallise together in a definite molar ratio. They exist only in the solid state. In solution, they break apart into all their individual ions.

Examples include Mohr's Salt ($FeSO_4 \cdot (NH_4)_2SO_4 \cdot 6H_2O$) and Alum ($KAl(SO_4)_2 \cdot 12H_2O$).

5. Complex Salts These are salts that contain a complex ion, which is a central metal ion surrounded by ligands (molecules or ions that donate lone pairs). Unlike double salts, complex salts do not fully dissociate in water — the complex ion remains intact.

Examples include $K_4[Fe(CN)_6]$ (potassium ferrocyanide) and $[Cu(NH_3)_4]SO_4$ (tetraamminecopper(II) sulphate).

Q(ii): Write a detailed note on pH.

Answer: Definition of pH The pH of a solution is defined as the negative logarithm (base 10) of the molar concentration of hydrogen ions (H^+) in that solution. It was introduced by the Danish chemist Søren Peder Lauritz Sørensen in 1909. The "p" stands for the German word "Potenz" meaning power, and "H" stands for hydrogen. Mathematically: **$pH = -\log[H^+]$**

The pH Scale The pH scale ranges from 0 to 14 and is used to express how acidic or basic a solution is. It is a logarithmic scale, which means each unit change in pH represents a tenfold change in H^+ concentration. If pH is less than 7, the solution is acidic. If pH is equal to 7, the solution is neutral. If pH is greater than 7, the solution is basic (alkaline).

Relationship Between pH and pOH At 25°C, water undergoes a slight self-ionisation: $H_2O \rightleftharpoons H^+ + OH^-$. The product of $[H^+]$ and $[OH^-]$ is always 10^{-14} at 25°C. This gives us the important relationship:

$pH + pOH = 14$ where $pOH = -\log[OH^-]$

pH of Common Substances Gastric acid (stomach) has a pH of about 1–2. Lemon juice has a pH of about 2–3. Pure water is 7. Blood is slightly basic at 7.4. Bleach has a pH of about 12–13.

The importance of pH is extremely important in everyday life and science. In medicine, the pH of blood must be maintained between 7.35 and 7.45 for normal body function. In agriculture, soil pH determines which crops can grow. In industry, pH control is important in manufacturing food, medicines, and chemicals. In environmental science, the pH of rainwater and rivers indicates pollution levels.

Calculation Example
For a 0.01 M HCl solution: $[H^+] = 0.01 = 10^{-2}$ M, so $pH = -\log(10^{-2}) = \mathbf{2}$

Q(iii): Define common ion effect. Explain with example how it affects chemical reactions.

Answer: Definition of Common Ion Effect The common ion effect states that when a strong electrolyte is added to a weak electrolyte solution, and both share a common ion, the degree of ionisation (dissociation) of the weak electrolyte is suppressed or decreased. In other words, the addition of a common ion shifts the equilibrium of the weak electrolyte in the backward direction, reducing the concentration of its ions in the solution. This is a direct application of Le Chatelier's Principle — when stress is applied to a system at equilibrium (by adding more of a common ion), the equilibrium shifts to oppose that stress, which means it shifts backwards.

Explanation with Example Consider the ionisation of acetic acid (CH_3COOH), which is a weak acid. It partially ionises in water as follows: $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$ At equilibrium, only a small percentage of acetic acid molecules are ionised. Now, if we add sodium acetate (CH_3COONa) to this solution, it is a strong electrolyte and dissociates completely: $CH_3COONa \rightarrow CH_3COO^- + Na^+$ This addition significantly increases the concentration of the acetate ion (CH_3COO^-), which is the common ion shared between acetic acid and sodium acetate. According to Le Chatelier's Principle, the equilibrium of acetic acid shifts to the left (backward direction) to reduce the excess CH_3COO^-. As a result, more CH_3COOH molecules are reformed and the degree of ionisation of acetic acid decreases considerably.

Effect on Solubility The common ion effect also reduces the solubility of sparingly soluble salts. For example, when NaCl (common source of Cl^- ion) is added to a saturated solution of AgCl, the increased concentration of Cl^- ions causes AgCl to precipitate further, reducing its solubility. $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$ Adding NaCl increases $[Cl^-]$, so the equilibrium shifts left, and more AgCl precipitates.

Applications of Common Ion Effect It is used in qualitative analysis to selectively precipitate ions. It is used in buffer solutions to maintain a steady pH. It is also used in industries to purify salts by precipitation.

Q(iv): Write different concepts of acids.

Answer: Over the years, chemists have proposed several theories to define what an acid is. Each theory is broader than the one before it.

1. Arrhenius Concept of Acids (1884) According to Svante Arrhenius, an acid is a substance that dissolves in water and produces hydrogen ions (H^+) in the solution. This was the earliest concept of acids and is limited to aqueous (water) solutions only. For example, HCl dissolved in water produces H^+ and Cl^- ions. H_2SO_4 produces $2H^+$ and SO_4^{2-} ions. The main limitation of this theory is that it only applies to water-based reactions and cannot explain acid-base behaviour in non-aqueous solvents.

2. Bronsted-Lowry Concept of Acids (1923) According to Johannes Bronsted and Thomas Lowry, an acid is any substance that can donate a proton (H^+) to another substance. This concept is broader than Arrhenius because it applies to non-aqueous systems as well. The acid that donates the proton becomes a conjugate base after losing it. For example, in the reaction $HCl + H_2O \rightarrow H_3O^+ + Cl^-$, HCl is the acid (proton donor) and H_2O is the base (proton acceptor). The conjugate base of HCl is Cl^-. A key feature of this theory is the concept of conjugate acid-base pairs. The limitation is that it still requires a proton (H^+), so it cannot explain acids that have no hydrogen atoms.

3. Lewis Concept of Acids (1923) According to Gilbert Newton Lewis, an acid is any substance that can accept a lone pair of electrons from another substance. This is the broadest concept of all and does not require hydrogen at all. A Lewis acid is an electron pair acceptor and a Lewis base is an electron pair donor. For example, BF_3 (boron trifluoride) is a Lewis acid because boron has an empty orbital and can accept a lone pair of electrons from NH_3 (a Lewis base). The reaction forms a coordinate covalent bond: $BF_3 + :NH_3 \rightarrow F_3B \leftarrow NH_3$. This concept explains the acid-base behaviour of substances like metal ions, $AlCl_3$, and $FeCl_3$ that have no hydrogen atoms at all.

Q(v): Define Lewis acids and bases, giving examples.

Answer: Lewis Acids A Lewis acid is defined as any atom, ion, or molecule that can accept a lone pair of electrons from another species to form a coordinate covalent bond. Lewis acids are electron pair acceptors. They are typically electron-deficient species, meaning they have incomplete octets, empty orbitals, or positive charges that attract electron pairs. Common examples of Lewis acids are BF_3 (boron trifluoride), which has only 6 electrons around boron and readily accepts a lone pair. $AlCl_3$ (aluminium chloride) is another example, as aluminium has an empty p-orbital. Metal cations such as Fe^{3+}, Cu^{2+}, and Zn^{2+} are also Lewis acids because they have empty d-orbitals that can accept electron pairs. Even H^+ (proton) is a Lewis acid because it has an empty 1s orbital.

Lewis Bases A Lewis base is defined as any atom, ion, or molecule that can donate a lone pair of electrons to another species to form a coordinate covalent bond. Lewis bases are electron pair donors. They must have at least one lone pair of electrons available for donation. Common examples of Lewis bases are NH_3 (ammonia), which donates its lone pair on the nitrogen atom. H_2O (water) is a Lewis base as it has two lone pairs on the oxygen atom. F^-, Cl^-, and OH^- ions are Lewis bases. Organic amines like CH_3NH_2 are also Lewis bases.

Lewis Acid-Base Reaction Example When BF_3 reacts with NH_3, BF_3 acts as the Lewis acid (electron pair acceptor) and NH_3 acts as the Lewis base (electron pair donor). The nitrogen atom in NH_3 donates its lone pair to the empty orbital of boron in BF_3, forming a new coordinate covalent bond and producing the adduct $F_3B \leftarrow NH_3$.

Significance of Lewis Concept The Lewis concept is the most general and powerful acid-base theory. It does not require the presence of water or even hydrogen, making it applicable to a vast range of chemical reactions including those occurring in organic chemistry, coordination chemistry, and industrial catalysis.